THE LABORATORY STUDY

OF

CHEMISTRY

HERBERT R. SMITH, M.A.
LAKE VIEW HIGH SCHOOL, CHICAGO

AND

HARRY M. MESS, B.A.
NICHOLAS SENN HIGH SCHOOL, CHICAGO

NEW YORK
HENRY HOLT AND COMPANY

ANCIENT LABORATORIES

An Alchemist at work. Boston Museum of Fine Arts.

An Ancient Laboratory and Symbols.

(Frontispiece)

MODERN LABORATORIES

A Student Laboratory. The Organic Laboratory, University of Illinois.

A Commercial Laboratory. Dearborn Chemical Co., Chicago. (*Frontispiece*)

THE LABORATORY STUDY
OF
CHEMISTRY

HERBERT R. SMITH, M.A.
LAKE VIEW HIGH SCHOOL, CHICAGO

AND

HARRY M. MESS, B.A.
NICHOLAS SENN HIGH SCHOOL, CHICAGO

NEW YORK
HENRY HOLT AND COMPANY

[1918]

Norwood Press
J. S. Cushing Co. — Berwick & Smith Co.
Norwood, Mass., U.S.A.

PREFACE

LABORATORY practice is now considered a prime requisite in the study of the sciences. Indeed, the "laboratory method" is being extended to various other subjects, even including mathematics. It is in harmony with the natural method by which man has collected the store of knowledge that forms the basis of modern civilization, as well as the method which the student must use after he quits school for active life outside. Chemistry offers a fruitful field for training in observation and reasoning, and for stimulating interest in science through direct contact with the facts of nature; but it will have its maximum result only when its study is initiated in the laboratory.

The study of chemistry too often has become largely the study of a textbook, with laboratory work relegated to the subordinate position of merely illustrating the text. With such a method the student in the laboratory merely watches for results already known, and seeks to answer questions that arise by consulting the text rather than by observing the experiment itself. Such study is highly artificial and affords little training. The proper function of the textbook is to furnish information that is not accessible to the student by experiment.

It is recognized, however, that the student should have some preliminary knowledge in order to perform an experiment intelligently and without loss of valuable time. This book supplies such information in the form of introductory discussions having the following aims: to reveal the purpose and general plan of the experiment; to supply interesting information, such as recent and historical facts pertaining to the principles involved; and to give an insight into the chemical actions that are to be observed. Information contained in the average textbook has, in general, been consciously avoided since that will be available in its proper place. An exception has been made, however, of the molecular, atomic, and ionic theories. In most textbooks, the treatment of these theories is scattered through the book so that it is difficult for the student to get a clear conception of them. This book treats each one completely, yet briefly, basing most of the statements on experimental data. The evidence of the existence of atoms has been stated as clearly as possible. The simple relationship between volumes and weights, as pointed out by Mr. C. E. Linebarger and also by Dr. H. N. McCoy, has been used.[1] It makes this otherwise difficult subject easy.

Any textbook can be used in connection with this book as a manual. For the convenience of the student references to numerous textbooks and other works on chemistry have been provided. The experiments dealing with fundamental facts are marked in the table of contents with (*) and they constitute a minimum year's work. Optional experiments are included in order to give the apter students enough work. The apparatus required for the essential experiments is simple and, as a general rule, inexpensive.

We make grateful acknowledgments for critical reading of the manuscript to Mr. C. E. Linebarger of Lake View High School, Chicago, Dr. John C. Hessler of James Millikin University, and Dr. H. I. Schlesinger of the University of Chicago; and for illustrative material to Assistant Director Morris Carter of the Boston Museum of Fine Arts, Dr. W. A. Noyes of the University of Illinois, and Mr. W. A. Converse of the Dearborn Chemical Supply Company.

---

[1] C. E. Linebarger, *Elementary Chemistry*. H. N. McCoy, *School Science and Mathematics*, June, 1908.

CONTENTS

QUALITATIVE ANALYSIS

APPENDIX

TABLE OF CONTENTS

(To be filled out by the student)

Number	Title of Experiment	Page

Number	Title of Experiment	Page

Number	Title of Experiment	Page

Number	Title of Experiment	Page

THE LABORATORY STUDY OF CHEMISTRY

THE USE OF GENERAL LABORATORY APPARATUS

HEATING

Bunsen Burner. This burner is in general use in laboratories wherever gas can be obtained as a fuel. In its simplest form the burner consists of a base having the tube (*a*) through which the gas enters, and the jet (*b*) from which it issues into the upright tube (*c*) that is screwed to the base. A short movable tube (*d*) with openings to coincide with those of the upright tube serves to regulate the quantity of air admitted through the openings and mixed with the gas before it is burned. If the apertures are closed so that no air can mingle with the gas, the flame of the burner is luminous and will deposit soot (carbon) on objects held in it. If the apertures are open, the carbon is completely burned; then the flame is non-luminous and deposits no carbon. The jet should have an opening large enough to allow all the gas to pass out that can be completely burned in a non-luminous flame when the air apertures are wide open. If much less gas is used with the apertures wide open, the flame is short and readily "strikes back," and the gas burns at the jet. In this condition the burner gets hot and gives off a poisonous gas, so the gas should be turned off at once. In any case the "striking back" is the result of too little quantity or pressure of the gas. To prevent its recurrence the jet should be cleaned out with a wooden splinter, and the gas cock opened wide when lighting the gas at the mouth of the burner. After lighting the gas, the size of the flame should be regulated by the gas cock. If corrosive liquids are spilled into the burner, the upright tube should be unscrewed, and the jet and threads of the screw cleaned and oiled.

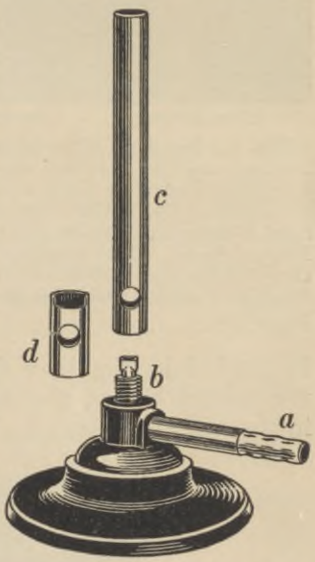


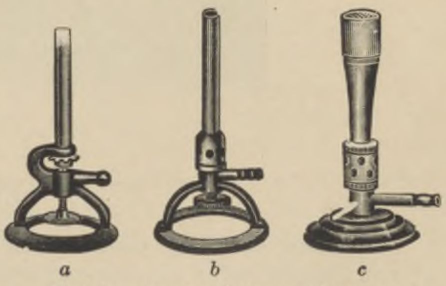
FIG. 1. Bunsen burner.

FIG. 2. Bunsen burner, low form.

The jet of the low form of burner (Fig. 2) is not so readily clogged by materials accidentally dropped into the top of the burner, so it keeps in better condition, and the burner is conveniently low.

The Boyce and Tirrell forms (Fig. 3, *a* and *b*) are very efficient in heating power, and have a screw to regulate the gas supply.

The Meker form (Fig. 3, *c*), having a grid at the top, obtains more complete combustion by better mixing of air and gas. Hence it has a more uniform flame without the cool cone found in other forms. See Figure 4.

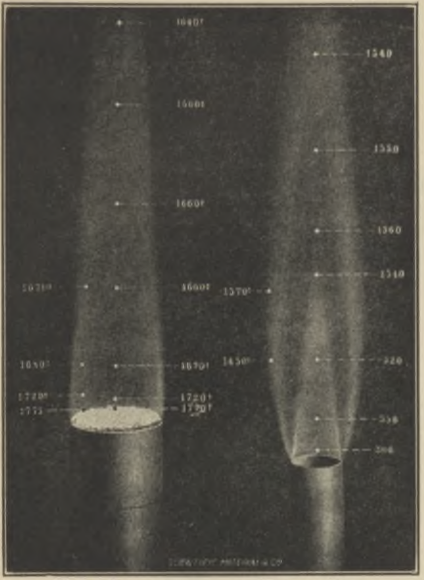


a *b* *c*

FIG. 3. Improved types of burners.

Dry Heating. If substances are heated dry, the test-tube should be held in a horizontal position (Fig. 5, *a*) so that any moisture condensing on the cool part of the tube will not run back on the hot glass and break it.

Heating Liquids. When liquids are heated in a test-tube, the tube should be held at an angle in the flame (Fig. 5, *b*) to prevent, as much as possible, the hot liquids from being thrown out of the tube by the force of the generated vapors. For further safety the mouth of the tube should be pointed, **not in the direction of the persons at the adjoining table**, but toward the sink, where any hot liquid thrown out can do no harm. If the tube is kept in slow motion while in the flame, the tube is not so likely to be broken. Liquids in flasks and beakers are usually heated over a wire gauze, which spreads the heat evenly over the surface of the glass and thereby lessens the chance of breakage. If glassware "boils dry" over a flame, breakage is almost certain. Such accidents can be prevented by using a water bath. The bath (Fig. 6) is filled two-thirds full of water and the burner placed underneath. The concentric rings on the top permit fitting the opening to the size of glass vessel used. When the water boils, turn down the gas enough to keep the water just boiling, and replenish the water in the bath when it becomes low.



Meker burner Bunsen burner

FIG. 4. The flame of the Meker burner is homogeneous and does not show any cold cone, its temperature being nearly uniform throughout. The coolest part of a Meker flame is hotter than the hottest portion of a Bunsen flame. The temperatures in the above photograph were measured with a Le Chatelier thermo-electrical couple of 0.2 mm. diameter.


FIG. 6. Water bath.

Heating Porcelain. Liquids can be boiled out entirely from porcelain dishes without breakage, provided the flame does not touch the dish when it becomes dry. They may be heated in the bare flame with the full power of the burner without breaking, provided the temperature of the dish is raised **slowly** and **evenly** to this highest point and then lowered in the same manner.

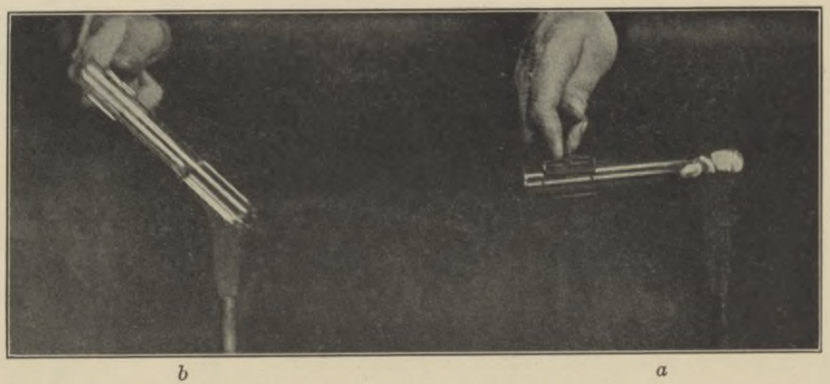


FIG. 5. *a*, Heating solids. *b*, Heating liquids.

SHAPING GLASSWARE

Breaking (Cutting) Glass. The small sizes of tubes and rods may be broken squarely across at any desired point by making a small notch with the corner of a file at this point. Then holding the tube with both hands, having the thumbs back to back and the thumb nails against the tube on the side opposite the notch (Fig. 7), push with the thumbs and pull with the fingers. If the tube does not break **easily**, the directions have not been followed. This method is not successful for tubes of a half-inch diameter or more, so other methods must be used. Heat red hot the end of an iron rod (one-fourth inch diameter) that has been cut squarely across, and after filing a notch in the glass at the desired point, lay the hot end of the rod upon the notch (Fig. 8). If the rod is hot enough, the glass will crack for a short distance. By drawing the rod slowly ahead of the crack the break may be led in any desired direction around the tube. Similarly, it is possible to cut, or rather break, any piece of glass along any desired line as exactly as it is possible for the operator to draw the hot rod. Figure 9 shows such a rod conveniently bent and mounted in a handle.


FIG. 7. Breaking glass tubing.

Bending Glass. When dry glass is heated slowly enough to red heat, it does not break, but becomes soft and can be bent or fashioned to any desired shape. If only a centimeter's length of a tube is heated where a bend is to be made, the tube wall on the inner side of the bend buckles and partly closes the tube. At the same time the outer wall becomes thin (Fig. 10, *A*). Such a tube is weak and breaks easily at the bend. If several centimeters of the tube's length are heated, the bend can be made more uniform (*B*) without the objectional features of the bend in tube (*A*). If a wing top is put on the burner tube, a flat flame is produced, which enables one to heat the glass tube for a length of several centimeters. It is easier to make a uniform bend with this device. If the top of this flat flame is not higher in the center, the lips of the wing top should be spread farther apart in the center. To bend a tube place it lengthwise in the flat flame with the point for the bend in the center of the flame. Rotate the tube to secure uniform heating. When the tube begins to bend, stop rotating and allow the tube to bend mainly by its own weight, assisting by a little pressure in the desired direction with the fingers. After bending the tube, close the air openings in the burner and hold the bent part of the tube in the luminous smoky flame until it is covered with soot. Allow the tube to cool slowly, then wipe off the soot. The coating of soot prevents rapid cooling. The slow cooling of glass is called **annealing**. If the annealing is not done, the glass may break the next time the previously heated part is put into a flame.


FIG. 8. Breaking glass with red hot iron.


FIG. 9. Rod for breaking glass.

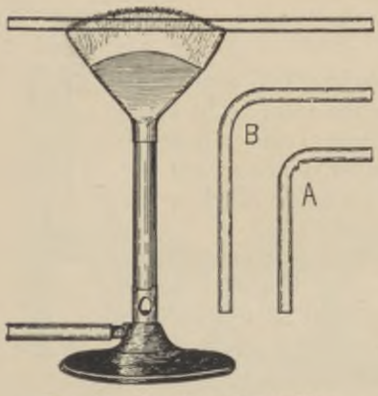


FIG. 10. Bending glass tubing.

Fire Polishing. The edges of broken tubes are very sharp and injure rubber tubing or stoppers in which the sharp ends are inserted. If these ends are held slantingly downward in a flame (Fig. 11) for 10–15 seconds while the tube is slowly rotated, the sharp edges become round. The ends of all tubes should be so treated.

Sealing End of Tube. Holding the tube by the ends, heat it as near one end as possible without burning the fingers. Rotate the tube slowly until it becomes quite soft. Keeping the tube in the flame, pull it apart where heated. The two pieces of tube will be closed at the ends where pulled apart. To shape the closed end operate it as for fire polishing, only longer. When the glass is soft, quickly put open end of tube in the mouth and blow carefully. Repeat the process until the desired shape is obtained, whether a rounded closed end or a bulb. Remember to heat the tube where it is to be blown out, and to rotate the tube to secure uniform symmetrical shape. A tube can be closed at

FIG. 11. Fire polishing.

the end without loss of length by heating the end until the glass is soft. Then by touching the soft glass with a piece of waste tubing the end may be pulled out and sealed. The waste piece of tubing is used as a handle. This plan is commonly used in sealing the broken ends of test-tubes.

Jet Tube. Proceed as for "sealing end of tube" until the glass is soft. Then remove the tube from flame and quickly pull it out or apart. Break the pulled-out part where it will give the desired opening for the jet.

Drilling Holes in Glass. Holes can be drilled through glass of any thickness without breaking, provided the glass is supported sufficiently on the side exactly opposite where the drilling is done. A drill can be shaped from a broken file. While drilling, keep the glass wet with a solution of the following parts by weight: 10 parts of ether, 26 parts of gum camphor, and 32 parts of turpentine.

FILTERING

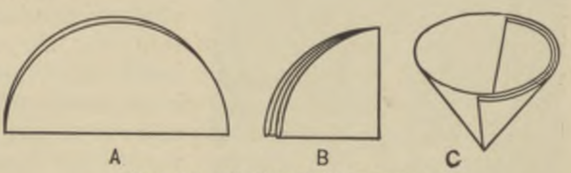


FIG. 12. Folding a filter paper.

FIG. 13. Pouring into a filter.

Folding a Filter. Fold the circular sheet to form a semicircular one (Fig. 12, *A*), then again to make a quarter of a circle (*B*). Separate one of the quarters from the other three (*C*). Put the conical filter in a funnel, moisten with water, and fit it smoothly to the funnel with the fingers. Support the funnel by a ring or a filter stand. The end of the funnel stem is ground off slantingly so that the liquid can be directed against the side of the receiving vessel. No drops form then to cause splashing. The liquid to be filtered should be poured on the side of the filter, being directed by the stirring rod, and not into the apex of the filter. Fill a filter only three-fourths full, never to the edge. Make no attempt to stir the liquid while on the filter or the filter may be punctured. The liquid passing through the filter is called the **filtrate**, and any solid remaining on the filter is called the **residue.**

MEASURING

FIG. 14. Graduate.

By Volume. Liquids are usually measured in cylindrical vessels called graduates (Fig. 14) because of the graduated scale of cubic centimeters on the side. The scale divisions are numbered from the bottom up for the convenient reading of the volume of a liquid poured into the graduate, and the divisions are also numbered downward for the convenient reading of the volume of a liquid poured out of a graduate filled to the zero mark.

The surface of liquids in narrow tubes always presents a curved shape resembling a new moon, so it is called a **meniscus.** It is always customary to read the position of the mid part of the curve (Fig. 15, *A*) with the eye at the same level as the surface of the liquid. If the eye of the observer is not at the same level as the surface of the liquid, the volume will be read incorrectly; for example, along the line (*B*) or (*C*). Such an error is called an error of parallax.

After the capacities of test-tubes and beakers are learned, the volumes of liquids can be estimated with sufficient accuracy for all ordinary cases. When **exact** volumes are specified, use some vessel that is designed for the accurate measurement of volumes.

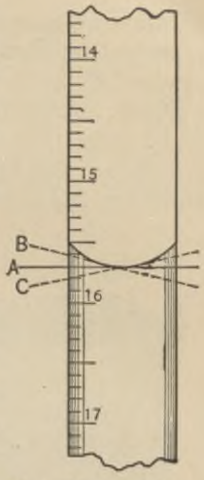


Fig. 15. Reading volumes.

Fig. 16. Pipette.

The pipette (Fig. 16) is a means of measuring a definite quantity of a liquid with great accuracy. The jet end of the pipette is thrust deeply into the liquid to be measured, and the other end placed in the mouth. By drawing the air out of the pipette the atmospheric pressure forces the liquid up into it. Take care to keep the end of the pipette well below the surface of the liquid or the entering air will throw the liquid into the mouth. Do not try to measure poisonous liquids with this instrument until you are skilled in its use. When the surface of the liquid is above the volumetric mark, remove the pipette from the mouth and quickly place the forefinger over the upper end. By moving the finger, the liquid is allowed to run down until the surface is exactly even with the mark. Avoid parallax! The pipette may now be removed from the liquid. The atmospheric pressure holds the liquid in the pipette if the finger is held closely over the upper end. Removal of finger will discharge the liquid. Finally, hold the tip of the pipette against the side of vessel measured into for 15 seconds to remove any liquid in the jet end. After using the pipette, it should be cleaned well, rinsed, and stood in a rack to drain dry. It may be used wet, provided it is first rinsed twice with the solution to be measured, discarding the rinsings.

Burettes are narrow, graduated tubes having some form of cock at one end for controlling the quantity of liquid discharged from them (Fig. 17). They are graduated in fifths or tenths of cubic centimeters. Note whether there are five or ten divisions to the cubic centimeter so that correct readings will be taken. In any case a burette should be rinsed at least twice with the liquid to be measured before it is filled for use. The burette should be filled some distance above the zero mark, then with the stopcock opened wide for a moment the tip below the cock will be filled with the liquid. Next the liquid is adjusted to the zero mark, the desired quantity used, and the exact quantity used determined by reading the position of the surface of liquid in the burette. Avoid parallax!

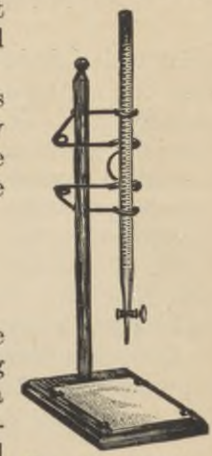


Fig. 17. Burette.

Since a change of temperature changes the volume of liquids, this effect must be considered in very careful measurement of volumes. By general agreement all volumetric apparatus is graduated to measure correct volumes at 15° C., so the liquid must be warmed or cooled, as the case may be, to this temperature when measured.

WEIGHING MATERIALS

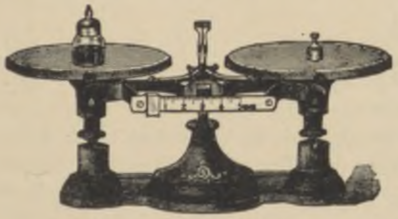


Fig. 18. Trip balance.

A trip balance (Fig. 18) is commonly used to measure the approximate quantities of chemicals to be used. A chemical balance for accurate weighing is shown in Figure 19. The materials are **always** placed on a paper or in a suitable dish on the left scale pan. In accurate weighing the container must be weighed first, then additional weights placed on the right scale pan to make up the weight of the material wanted. The material is then put on the left pan until the pans are balanced. Any object to be weighed is placed on the left pan and the heaviest weight, or one thought to be heavier than the object, is placed on the right pan. Whenever an added weight makes the right pan heavier than the other pan containing the object, this weight is removed and the next smaller one used. If the added weight does not make the right pan heavier, this weight is kept on the pan and the next smaller one is used. This process is continued until the smallest weight has been used. The separate weights are first counted up from the vacant spaces in the weight box, after which the figures representing them are checked when the weights are replaced in the box. If the balance has a graduated beam with a sliding weight, inconvenience and loss of small separate weights may be avoided. An automatic scale (Fig. 20) having no weights to lose or count up is still more convenient and a great time saver.

Use of the Hydrometer. For the rapid determination of density (specific gravity) the hydrometer (Fig. 21) is commonly used. The scale division marked 1.000 is the depth to which the instrument will sink in pure water. In lighter liquids it sinks deeper, so densities of less than **one** are graduated above this mark, while those greater than **one** are found lower on the scale. The figures expressing the density are read directly from the scale. Specially graduated hydrometers for determining the percentage of alcohol in a liquid are called alcoholometers. Those for measuring the density of milk are called lactometers, of which there are two scales in present use. The Quevenne scale is the most convenient, being graduated from 15° to 40° to correspond to the densities 1.015 to 1.040. The scale reading must be preceded by 1.0 to give the figures for the density. The New York Board of Health scale has its zero equal to the density of water (1.0) and its 100° mark equal to the density 1.029. These scale readings may be converted to Quevenne degrees, so far as they correspond, by multiplying by 0.29. Milk should be at 60° F. (approximately 15° C.) when the density is to be determined.

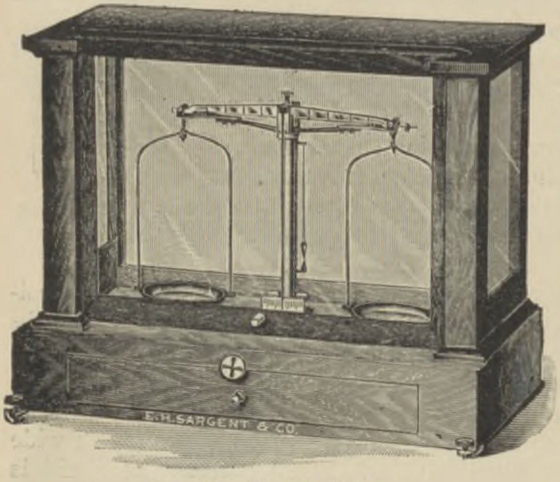


Fig. 19. Chemical balance.

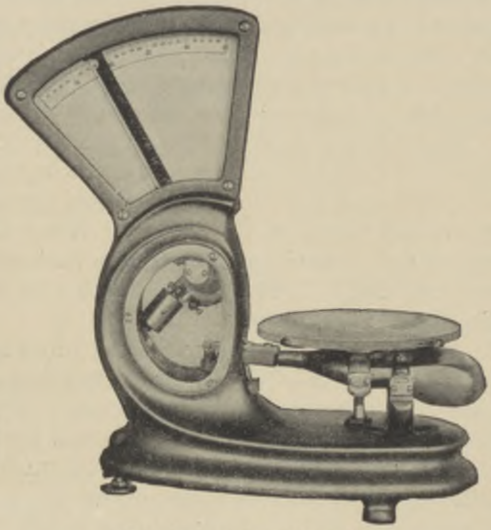


Fig. 20. Automatic scale.

Fig. 21. Hydrometer.

LABORATORY DIRECTIONS

In order that the laboratory work may be as pleasant and profitable as possible, it is essential that all members of the class give careful heed to the following general regulations.

Conduct. No student can do **excellent** work nor even **good** work and, at the same time, talk much to his neighbor. Careful observation will verify this statement, "The most noisy students are among the poorest in grade." Set an excellent standard in this matter and try earnestly to live up to it. This is the **first step** toward doing the best work. The instructor will appreciate such efforts very much. Perfect work is not expected from any student, for every one makes mistakes. If his **attitude** is **right**, he will not make the same mistake twice. If he is unable to reach a clear understanding of work after a faithful attempt, he should not fail to consult the instructor.

System. Every piece of apparatus that is used should have a definite place for its keeping in the laboratory. The instructor has located the apparatus where it will be most accessible. The "saving of time" is the watchword in all modern industrial labor. Begin the practice of it here by making every movement count toward the performance of an experiment. **Return articles to place after use.** The following points are frequently neglected:

1. Care in handling chemicals. The instructor will illustrate how this is best done, calling attention to all points to be observed.
2. If chemicals are spilled on the table or floor, clean up at once.
3. If the student must lay a stopper down, let it be put top down on the table. Always replace it after material is obtained, even if another student is waiting for the bottle.
4. Put back **none** of the contents of a reagent bottle except at the direction of the instructor.
5. Report any breakage of general apparatus at once to the instructor. Immediate repairs may save further loss.
6. Take no school material from the laboratory without permission from the instructor.
7. Undertake no experiments of your own devising without first consulting the instructor.

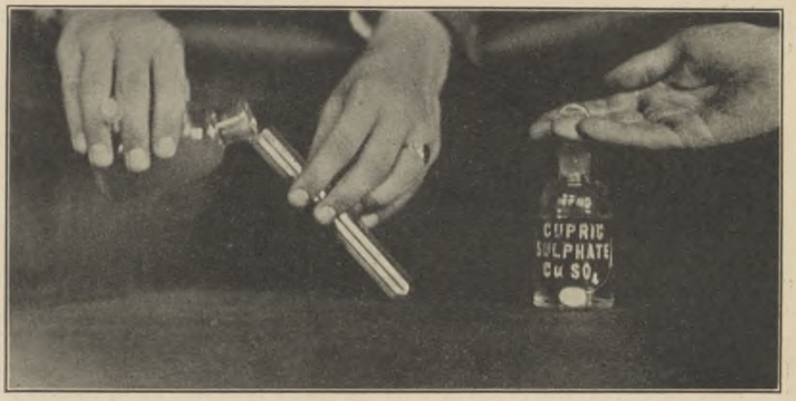


Fig. 22. Pouring from a bottle.

Failure to comply with any one of the last three of above points will be sufficient cause for dismissing a student from class.

Disposal of Waste. **All solids** must be put in the receptacles for waste at the tables. If they are thrown into the sinks, the pipes may become clogged. If so, the students at that table will be assessed the cost of clearing them. **Liquids** may be emptied into the sinks. If much acid or other reagent is emptied, let the water run for a few minutes to flush it out into the sewer where it cannot corrode the sink or pipes. **Gases** that are ill-smelling or poisonous should be generated only in the hoods. Whenever any material in the process of an experiment begins to smoke or fume, it should be carried at once to the hood.

Cleanliness. All glassware used must be thoroughly clean. After use it should be cleaned before it is put away. Then it is always ready for use. Each student should have a clean, soft towel, soap, and a long, rubber apron to protect the clothing from possible injury. At the close of a period of laboratory work each student should sponge off the table top with a wet sponge or cloth. Then

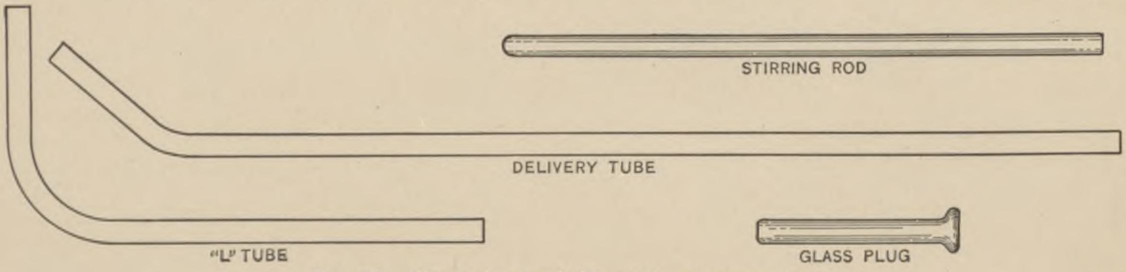


Fig. 23. Glass tubes needed in setting up apparatus.

the table should be wiped off with a dry sponge or cloth. Keep the reagent bottles in order on shelf. Wipe them with a damp cloth each week.

Danger !! Many reagents are very active. If they are handled carelessly, there is serious danger. Whenever there is any possibility of danger the student will be warned by directions in the course of the experiment. If the student **follows instructions** and uses reasonable care in other things, he is in no more danger in the laboratory than in any other room in the building. Taste no substance except when directed to do so, and then do not swallow the material tasted. Smell (**do not inhale**) the fumes or vapors of a substance **cautiously**, until it is certain that the odor is not violent.

Preparation for Laboratory Work. The entire experiment assigned should be read carefully to determine the purpose of it. Then it should be studied to know the method of procedure. The heading of the record sheet including the purpose should be written previous to the laboratory period so that the experiment can be started at the beginning of the period. After performing the experiment, the introduction should be studied with as many references as possible for following recitation.

Laboratory Record. A record, preferably in ink, must be made of observations and results of experiment as it is performed. The heading, consisting of subject, experiment number, date, name, and purpose, should be neatly arranged as the instructor indicates. Begin each new experiment on a new page. Never use abbreviations. Leave a margin at the top and sides, and use the same division headings in the record as in the printed directions. Whenever apparatus is " set up," make a cross-sectional diagram of it and label all important parts. Be sure to answer all questions. The sign (?) in the directions indicates that a record of observations or results is expected. Leave no parts unrecorded. If in difficulty, consult the instructor. Use blank spaces in printed directions whenever they are designed for record. Otherwise use a blank sheet with proper heading. Use a tabulated form of record whenever possible. Express fractions decimally, being careful about the location of the decimal point. If an error is made in the record, correct it as neatly as possible, making no erasures whatever.

Independence. Please remember that what you copy from your neighbor does not educate you, and is of little value. Preserve your independence. After an honest effort ask the instructor for assistance. When performing an experiment, observe with **your own eyes**. When writing a record, explain with **your own words**. When answering questions, think with **your own mind**.

ABBREVIATIONS OF TITLES OF REFERENCE TEXTS

B. & W. — Blanchard & Wade, *Foundations of Chemistry*. American Book Co.
Brownlee. — Brownlee, Fuller, Hancock, Sohon & Whitsit, *First Principles of Chemistry*. Allyn & Bacon.
H. & S. — Hessler & Smith, *Essentials of Chemistry*, Revised Edition. Benj. H. Sanborn & Co.
McP. & H. — McPherson & Henderson, *First Course in Chemistry*. Ginn & Co.
M. & L. — Morgan & Lyman, *Elementary Chemistry*. The Macmillan Co.
Newell. — Lyman C. Newell, *General Chemistry*. D. C. Heath & Co.
Smith. — Alexander Smith, *Elementary Chemistry*. The Century Co.

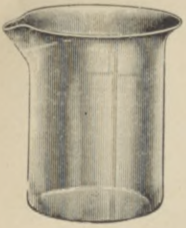


Fig. 24.

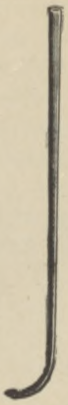


Fig. 25.

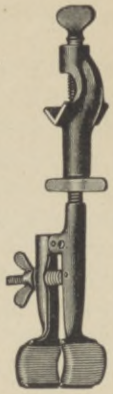


Fig. 26.

Fig. 27.

Fig. 28.

Fig. 29.

Fig. 30.

Fig. 31.

Fig. 32.

Fig. 33.

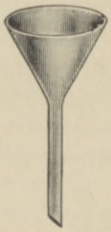


Fig. 34.

Fig. 35.

Fig. 36.

Fig. 37.

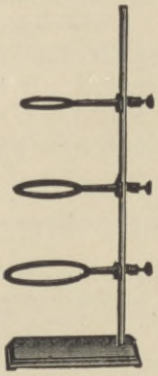


Fig. 38.

Fig. 39.

Fig. 40.

Fig. 41.

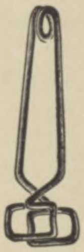


Fig. 42.

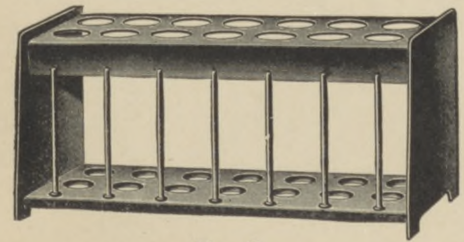


Fig. 43.

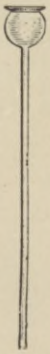


Fig. 44.

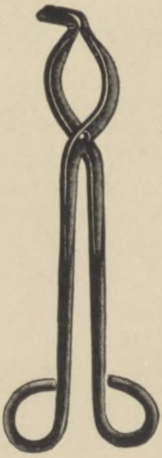


Fig. 45.

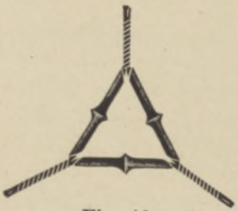


Fig. 46.

Fig. 47.

Fig. 24. Beaker. Fig. 25. Blow pipe. Fig. 26. Burette clamp. Fig. 27. Calcium chlorid tube. Fig. 28. Combustion spoon. Fig. 29. Erlenmeyer flask. Fig. 30. Evaporating dish. Fig. 31. File. Fig. 32. Florence flask. Fig. 33. Forceps. Fig. 34. Funnel. Fig. 35. Gas generating bottle. Fig. 36. Mohr's pinch clamp. Fig. 37. Mortar and pestle. Fig. 38. Ring stand and rings. Fig. 39. Rubber stopper. Fig. 40. Test-tube. Fig. 41. Test-tube brush. Fig. 42. Test-tube clamp. Fig. 43. Test-tube rack. Fig. 44. Thistle tube. Fig. 45. Tongs. Fig. 46. Triangle. Fig. 47. Watch glasses.

PROPERTIES OF MATTER

Chemistry deals chiefly with the composition of matter. Chemists are continually discovering and making new substances. Over 200,000 substances are known at present, all differing from one another by certain characteristic qualities called **properties**. By means of these properties we know of the existence of matter and can recognize a substance as one previously studied. So characteristic are some properties that we may frequently distinguish a substance from all others by observing only one or two properties. Because of this we frequently hear certain similes: "heavy as lead," "white as snow," "sweet as sugar." When a new substance is made or discovered, the chemist describes its properties and classifies it with other substances of its kind. After being made known to the world in the literature of chemistry the new substance awaits a possible use.

In primitive times men used substances as Nature provided them. Since then the needs of men have become more exacting; so much so that few substances are now used as Nature composed them. Raw materials must now be converted into substances having properties that fit them exactly for certain uses. As civilization has developed the need of the services of the chemist has grown, and as it continues to develop his services will be still more indispensable. It is also apparent that the study and knowledge of properties must be of great importance to each person in every occupation of the world.

It is not possible for one to know the properties of all substances, because of their great number. This is not necessary since each person in his chosen vocation deals with relatively few of the total number of substances actually known. We cannot be sure that we know all the properties of any one substance. Researches in science reveal new properties that have been hidden until the proper refinements of method are at hand to discover them to human minds. But the practical things to know are the common properties of the substances that we use every day, since the efficient use of these substances depends on our knowledge of their properties. Such properties are: color, odor, taste, feeling, sound, specific gravity, melting and boiling points, crystal form, luster, ductility, malleability, tenacity, elasticity, hardness, brittleness, conductivity of heat and electricity, solubility, porosity, and transparency.

Other important properties of a different kind are those which appear when a substance is heated, acted on by electricity, or brought into contact with other substances. In these instances the substance is usually transformed into a new one having different properties. For example, if iron is kept in moist air, it slowly changes into a brown substance (rust) having entirely different properties. In this form iron is not so useful. It is of the greatest practical importance to prevent such changes, and to transform useless substances into useful ones. Such is the work of the chemist. In dealing with the changes in matter he recognizes two kinds: **physical changes**, in which the composition of substances is unchanged, as is shown by **no alteration** of the characteristic properties; and **chemical changes**, in which the composition is changed and new substances are formed having new properties.

In addition to the actual matter in substances, there is a definite quantity of energy contained in each. Those substances which react readily usually contain much energy. When they change chemically, energy is liberated in some readily noticeable form as heat, light, or electricity. Energy so contained in substances is called internal or chemical energy (energy that is available only by chemical change).

Study Topics: Duncan, *New Knowledge:* "The Three Entities," pp. 1–7. *Some Chemical Problems of To-day:* "The Prizes of Chemistry," pp. 1–18; "On the Trend of Chemical Invention," pp. 107–127.

Bird, *Modern Science Reader:* "Unsolved Problems of Chemistry," pp. 305–322; "The Old and New Alchemy," pp. 245–270.

Martin, *Triumphs and Wonders of Modern Chemistry:* "The Wonders of Chemical Change," pp. 55–81.

References: Smith, pp. 1–16; M. & L., pp. 3–7; Newell, pp. 1–8; B. & W., pp. 9–13; H. & S., pp. 1–3; McP. & H., pp. 8–12, 19; Brownlee, pp. 1–7.

EXPERIMENT 1

Purpose: ....

....

Apparatus: Sheet iron square, 10 × 10 cm., or iron spoon.

Materials: 2 cc. bulk of each: lump charcoal (carbon), salt (sodium chlorid), sugar (sucrose), moth balls (naphthalene), sulfur, paraffin, tin, 5 cm. of magnesium ribbon, rock candy.

I. **Physical Properties.** Get a few at a time of the above materials on separate labeled papers, observe carefully the common properties, and record results in the tabulated form below. Reserve a 3-mm. piece (equivalent to a 3 mm. cube) of the first six substances, and all of the last two for use in Part II after use in Part I. If the work is unfinished at the end of the period, save the materials in papers until the next period. Put the large portion of each substance in separate test-tubes one-fourth filled with water. If a substance floats, its specific gravity is less than one. Record as (1−). Be sure that it really floats and is not merely sticking to the side of the tube. If it sinks, its specific gravity is greater than one. That is, it is heavier than water. Record as (1+). Heat each substance while in the water by holding the tube in the burner flame. Shake gently to hasten solubility. Do not heat the tube above the surface of the water. It may break if you do. If the substance is soluble, it will mix with the water and cannot be distinguished from it. If the substance is insoluble, note whether it melts or not. Allow substances to cool in the tubes while continuing work and observe the charcoal, paraffin, and naphthalene again when cool.

SUBSTANCE	COLOR	ODOR	TASTE	FEELING	SOUND	SP. GR.	SOL. OR INSOL.	MOST CHARACTERISTIC PROPERTY
Charcoal........								
Paraffin.........								
Naphthalene.....								
Salt.............								
Sugar...........								
Tin.............								
Magnesium......								

Questions: 1. What special sense gives you the most information? The least?

....

2. Of what value is it to have a knowledge of properties?

....

....

II. **Chemical Properties.** Use the reserved portion (3 mm.) of each substance from Part I. Do not use more than this specified amount, except the whole portions of tin and magnesium. Heat the charcoal as it rests on a corner of the wire gauze. Hold the magnesium ribbon in the flame with the crucible tongs. Heat the remaining substances on the sheet iron square or iron spoon, supported at a convenient height above the burner by means of the iron ring clamped to the stand. Remove any residue from previous substance before heating the next. Note any flame, smoke, odor, change of appearance, and the relative amount of residue. Record observations as tabulated below.

SUBSTANCE	BURNS?	FLAME? COLOR	FUMES? SMOKE?	ODOR OF FUMES	AMOUNT OF RESIDUE	APPEARANCE OF RESIDUE
Charcoal...						
Paraffin....						
Naphthalene						
Salt.......						
Sugar......						
Tin.......						
Magnesium.						

3. Are your senses of sight, smell, taste, touch, and hearing equally well developed? Give reasons for your answer.

....

....

4. What difficulties would every one encounter if substances had no properties?

....

....

5. Of what importance in our daily activities is the fact, "Under the same conditions the properties of a substance are always the same" (Law of Properties)?

....

....

....

III. **Changes in Matter.** (*a*) Examine a piece of rock candy, record its properties, especially its taste...

Pulverize a centimeter piece in a clean mortar, and taste it again............................
Put the powdered mass in the evaporating dish, fill it half full of water, place it on an iron ring, and heat the mixture while stirring it. When the powdered candy has dissolved, taste the solution.

.... Heat the solution vigorously until the liquid thickens, then heat **slowly** until the syrup just begins to turn brown. Remove the syrup with a clean spoon, cool, and taste it..

(*b*) Heat the remainder of the syrup in the spoon until it becomes black and **scales loose** from the spoon. Remove some and identify it by its properties...............................

....

Clean and put away the apparatus, and leave the table in good condition.

6. Of what is rock candy made?..

7. Does this substance persist through all the changes in Part (*a*)?........................

Part (*b*)? ....

8. Are physical or chemical changes brought about by the acts of pulverizing, dissolving, and evaporating?....................................How do you know?...................

9. What kind of a change takes place in Part (*b*)?....................................

How do you know?..

10. What is the residue in Part (*b*)?...

11. Give your method of distinguishing between physical and chemical changes..............

....

....

Optional Work: (Record on a separate sheet.) Heat a platinum wire, a copper wire, a piece of wood, and a piece of glass in the burner flame. Classify the changes as physical or chemical and state reasons for such classification.

Classify the changes noted in Parts I and II of this experiment. State the evidence which guides your judgment in each case.

VOLUMES, WEIGHTS, AND DENSITIES

Chemistry is one of the " exact sciences " and no great progress can be made by the student until he learns the use of instruments and apparatus for the accurate measurement of quantities. The most important are those for the measurement of volumes and weights. When the words **about** or **approximately** are used with a quantity, no great accuracy in its measurement is required, and an experienced experimenter can usually estimate the amounts closely enough. In other cases, however, the success of the experiment will often depend upon the care used in measuring the quantities exactly. The use of quantities of materials largely exceeding those called for must be particularly guarded against in order to avoid a wholly unnecessary waste of material and time in performing the experiment.

Quantity of matter may be measured by volume, the amount of space occupied, or by weight, the measurement of the earth's attraction for substances on it. Volume units probably came into use first. When balances and scales came into use as means of measuring the quantity of matter more accurately by weight, it was natural that the usual weight of a substance found in the unit of volume should become the unit weight of that substance. So a bushel of potatoes or wheat became 60 lb., and a barrel of flour 196 lb. These statements are the relations between weight and volume units. When scientists considered this relation of mass to volume for all substances, they introduced the word **density** and expressed it as **the weight of a substance in unit volume.** They also used the term **specific gravity** when they compared the weight of a substance with the weight of an equal volume of water as a standard substance. This is only heaviness in terms of water. Density and specific gravity are both expressions of the quantity relation between mass and volume.

In the English system the units of mass (weight) and volume were established without reference to a simple convenient relation to the standard substance water. 1 cu. ft. of water weighs 62.4 lb. In the metric system the unit of mass, the gram, was chosen as the amount of water, the standard substance, found in unit volume, the cubic centimeter. Thus 1 cc. of water weighs 1 g. The number of grams of any substance in 1 cc. also indicates how many times that substance is heavier than an equal volume of water. Thus in the metric system, the numerical values of density and specific gravity are the same.

The density of a substance can always be found by dividing the weight of a certain portion of it by its volume. If D = density, W = weight, and V = volume, we may write the following equations:

$$(1)\quad D=\frac{W}{V}. \qquad (2)\quad W=DV. \qquad (3)\quad V=\frac{W}{D}.$$

When a certain volume of a solid occurring in irregular pieces is required, the density is looked up in some reference work or determined by experiment, and then the weight of the required amount of the substance is computed by means of equation (2). It is quite easy to get a certain weight of a solid. With liquids, however, it is easier to measure out a certain volume than a weight, so when a certain weight of liquid is required, the corresponding volume may be computed from the density by means of equation (3). For the convenience of workers in certain industries, tables have been made showing the relation between the densities and the percentage compositions of solutions of certain solutes. Thus the per cent of alcohol in a water solution may be easily found by determining the density of the solution and referring to the proper tables (see p. 246).

When extreme accuracy is not required, the density of a liquid is usually obtained by floating a hydrometer (Fig. 21) in some of the liquid in a tall cylinder. The reading of the bottom of the meniscus on the stem (avoid parallax) gives the density in grams per cubic centimeter.

References: M. & L., p. 6; H. & S., pp. 4–5; B. & W., p. 42; Brownlee, pp. 95–96; Smith, pp. 6, 71.

VOLUMES, WEIGHTS, AND DENSITIES

EXPERIMENT

Purpose: ....

Apparatus: Pipette, trip scale, hydrometer.

Materials: Roll sulfur, sodium chlorid (saturated solution).

I. **Capacities.** Fill the graduate with water to the highest mark (see Fig. 15) and then fill the 10-cm. test-tube from the graduate. The volume of the water poured out is the capacity of the test-tube. Record it below and repeat with the 15-cm. test-tube:

Capacity of the 10-cm. test-tube .... cc.
Capacity of the 15-cm. test-tube .... cc.

II. **Weights and Volumes.** (*a*) Weigh a clean dry evaporating dish as accurately as the balance permits and record the weight. Place in the dish about 10 g. of roll sulfur, using no powder, but pieces small enough to slip into the graduate easily. Weigh again, carefully, and record. Fill the graduate with water exactly to the mark indicating 10 cc. less than its graduated capacity, and let the sulfur slide into the water gently so as not to splash. Loosen air bubbles by means of a fine wire or by gentle shaking; lose no water. Read and record the volumes before and after adding the sulfur. Compute the weight and volume of the sulfur. Put the wet sulfur into a separate dish, not into the stock bottle.

Weight of the dish and sulfur .... g.
Weight of the dish .... g.
Weight of the sulfur .... g.
Volume of the water and sulfur .... cc.
Volume of the water .... cc.
Volume of the sulfur .... cc.

(*b*) Clean and dry the evaporating dish used above. Fill a clean pipette (Fig. 16) to the mark with a saturated solution of salt (sodium chlorid) and run the solution slowly into the dish. Weigh the dish and solution. Then pour the solution into the sink.

Weight of the dish and solution .... g.
Weight of the dish .... g.
Weight of the solution .... g.
Volume of the solution .... cc.

III. **Density.** (*a*) From the results of II (*a*) compute the weight of 1 cc. of sulfur.

Density of sulfur .... g. per cc.

(*b*) From the results of II (*b*) compute the weight of 1 cc. of the salt solution.

Density of the salt solution .... g. per cc.

(*c*) Pour some of the same salt solution into a clean dry cylinder and float a hydrometer (Fig. 21) in it. A hydrometer for liquids heavier than water must be used. Read and record the density of the solution. Compare this reading with the density as determined just above, and if the difference is more than 0.02 g. per cc. repeat one or both determinations.

Reading on the hydrometer .... g. per cc.
Difference between density determinations ....

IV. **Optional.** Find the capacities of the flasks and beakers or other apparatus in your desk, and try to estimate various amounts of water in different vessels, measuring the amounts afterwards in your graduate as a check.

CLASSIFICATION OF MATTER

Science is often defined as " classified knowledge," and an important part of the scientific method consists in bringing together various facts so that relationships between them may be perceived, understood, and remembered, and laws and theories based upon these relationships. The development of the present science of chemistry from the " black art " of alchemy in the Middle Ages was due to the introduction of various schemes of classification. One of the most important of these classifications is that which arranges the 200,000 different kinds of matter into a comparatively small number of groups.

As a first step, all the different substances are divided into two main classes. Most substances can be separated into two or more substances, undergo **decomposition**; or they can be made by the simple union of two or more substances into one, a variety of chemical action known as **combination** or **synthesis**. Such substances are called **compounds**. The remaining substances, about eighty in number, have never been made by the combination of other substances, and have never — by human agency, at least — been decomposed into two or more substances. Accordingly, they are regarded as being simpler forms of matter than the compounds, and are called **elements**. (See list of elements on p. 236.)

Very few of the materials found in nature are single substances in a pure state, but consist of large or small particles of two or more substances more or less uniformly distributed through the material and loosely or firmly held together as the case may be. Such a material is called a **mixture**, and its ingredients are called **components**. These components are not united chemically, and each retains its identity and its own set of physical and chemical properties. That the properties of a mixture are intermediate between the properties of its components, is due to the fact that **under the same conditions a substance always has the same properties**. This is called the **Law of Properties**.

It is not always easy to distinguish between mixtures, compounds, and elements. On account of the difficulty in finding the proper methods of analysis and synthesis, some substances now known to be compounds were long regarded as elements. If a substance can be analyzed or synthesized, it is not an element, but in some cases it is difficult to determine whether it is a mixture or a compound. Three distinguishing characteristics of mixtures and compounds are brought out in this experiment; two others are: (1) compounds are perfectly homogeneous — all parts exactly alike — while mixtures are usually not; (2) all samples of a given compound are always found to contain the same elements in the same proportion, or, more briefly, each compound has a definite percentage composition, while mixtures may be made with any proportions of the components.

For the sake of brevity, chemists have devised a way of summarizing a chemical change by means of signs and symbols. These symbols are a very important part of chemical language. (See list of the symbols of the elements on p. 236.) For practice preliminary to their use, the full names of the elements and compounds will be used at present in writing " **word equations** " of chemical changes. The following examples of such word equations may be used to express some of the results of the first experiment.

$$\text{carbon} \underset{\text{(with)}}{+} \text{oxygen} \underset{\text{(forms)}}{\longrightarrow} \text{carbon dioxid.}$$

$$\text{sulfur} + \text{oxygen} \longrightarrow \text{sulfur dioxid.}$$

$$\text{magnesium} + \text{oxygen} \longrightarrow \text{magnesium oxid.}$$

As will be noticed in the equations above, compounds of two elements, binary compounds, are named by putting together the names of both elements, changing the second to end in " **id**." If one of the elements is a metal, its name comes first. When both are non-metals, the proper or usual order must be learned from the instructor or the text.

References: B. & W., pp. 15–19; Smith, pp. 5, 290; M. & L., pp. 12–14; H. & S., pp. 7–9; Newell, pp. 5–6.

MIXTURES AND COMPOUNDS

EXPERIMENT

Purpose: State the purpose on record sheet after the proper heading.
Apparatus: Magnifying glass or microscope, square of sheet iron, tin can cover or old spoon.
Materials: Powdered sulfur, zinc dust, carbon disulfid, filter papers, 4-cm.

I. **Mixtures.** (*a*) Examine some powdered sulfur and record its properties on another sheet, copying the form below. Look at a little of it through a magnifying glass or microscope (?). Drop a pinch of the sulfur into about 3 cc. of concentrated hydrochloric acid in a test-tube and mix by gentle shaking. Note any effervescence or odor other than that of the acid used (?). Add a pinch of sulfur to about 3 cc. of carbon disulfid in a clean, dry test-tube and shake for two minutes. Note whether any of the sulfur disappears (?). Pour the liquid through a small, dry filter (see Fig. 13) into a watch glass and allow the liquid to evaporate in the hood. The appearance of a residue of sulfur in the watch glass **proves** that sulfur is soluble in the liquid.

(*b*) Examine some zinc dust in the same manner and record your observations in the same form.

(*c*) Stir together 4 g. zinc dust and 2 g. powdered sulfur. Examine the resulting powder exactly as you did the sulfur in (*a*), noting particularly whether the zinc and sulfur can be distinguished by the magnifying glass or microscope (?). Also examine carefully and try to identify any residue on the paper or in the test-tube as well as that on the watch glass when trying the solubility in carbon disulfid (?).

RECORD OF OBSERVATIONS

(Use a similar tabular form.)

SUBSTANCE EXAMINED	COLOR	UNDER MICROSCOPE	ACTION ON HYDROCHLORIC ACID		BEHAVIOR IN CARBON DISULFID	
			Effervescence	Odor	Residue in Tube	Residue on Glass
Sulfur.........						
Zinc............						
Sulfur and Zinc						
Compound of sulfur and zinc						

Questions: 1. Do the properties of the zinc persist in the mixture? Of the sulfur?

2. What kind of change (physical or chemical) is involved in making the mixture? In separating its ingredients?

II. **Compounds**: (*a*) Place the rest of the mixture on the square of sheet iron and support the latter on the stand. Heat the powder by placing the lighted burner beneath it, removing the burner as soon as vigorous action begins. Describe the action (?).

(*b*) When the product is cool, examine it carefully as you did the sulfur in I (*a*) and record all observations.

Questions: 3. Do the properties of the zinc persist in the compound? Of the sulfur?

4. What kind of change occurs in the making of the compound?

5. Write a word equation as a brief statement of the making of the compound.

6. Can you separate the zinc and sulfur from the compound by the same sort of change as that used in separating them from the mixture? Why?

7. Divide the rest of the page into two columns, head one Mixtures and the other Compounds, and write in them contrasting statements of the **five** differences between mixtures and compounds.

THE THEORY OF MOLECULES

One of two hypotheses regarding the constitution of matter must be true. Either it must be continuous with no divisions and intervening spaces, or it must be composed of mutually cohering particles having intervening spaces. This hypothesis of separate particles was first advanced by an ancient Greek philosopher, Democritus, over 2300 years ago. His followers and their opponents argued very vigorously in support of their respective hypotheses, but no experiments were performed to test the validity of either one. But during the last 200 years much experimental work has been done, and all the results lead to the acceptance of the hypothesis of separate particles.

This hypothesis assumes that matter is not indefinitely divisible. A grain of sugar can be crushed into very many small particles, each one having all the properties of sugar, and each of these could be again and again divided into still smaller pieces — each one still sugar — if means and instruments for holding and dividing such small objects were available. But it is assumed that this continued division of the sugar particles would at last produce such a small particle that further division would separate the compound substance, sugar, into its constituent elements. This **smallest particle of a substance which possesses all the properties of that substance is called the molecule**. The results of the chemical analysis of sugar show that it is composed of carbon, hydrogen, and oxygen. This fact makes it necessary to assume that there are particles of matter smaller than the molecules. They are called **atoms**.

The Evidences of the Existence of Molecules. All substances are compressible to some degree, solids and liquids slightly and gases greatly compressible. It is more reasonable to explain this by the assumption of separate or discrete particles than by the continuous matter hypothesis, for the molecules can be brought closer together, and the given quantity of matter occupies less space. As solids and liquids are but slightly compressible, the molecules in these states of matter must be quite close together, while the molecules of gases must be assumed to be much farther apart in order to account for the great compressibility of gases. The great attraction between the closely packed molecules of solids accounts for their tendency to keep their shape. In liquids the molecules are not so close, so the attractions between the molecules are less, and the liquid is mobile. In gases the attractions between the molecules are negligible because of the relatively great distances between them, and the gases tend to spread out indefinitely into all the space available.

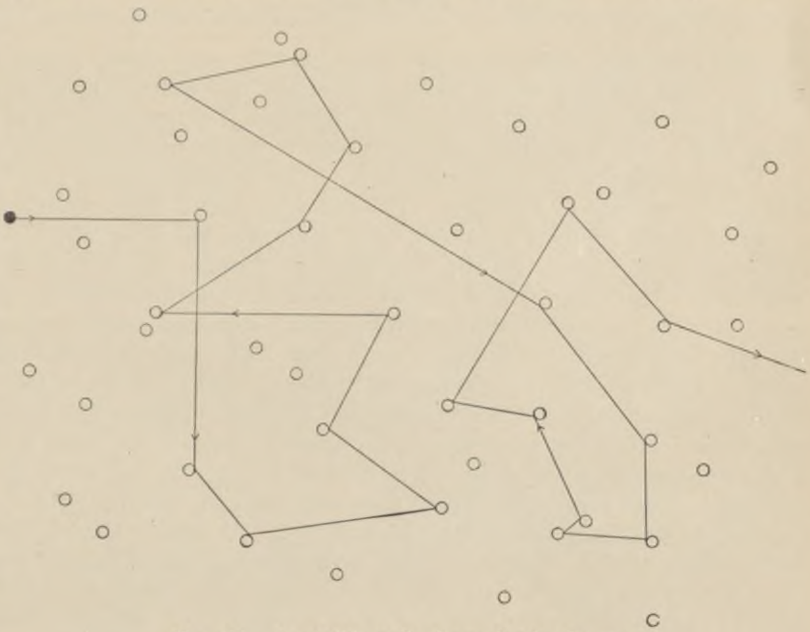


Fig. 48. A possible path of a molecule of a gas.

When alcohol is mixed with water, the final volume of the mixture is less than the sum of the original volumes of the alcohol and water. A reasonable explanation of this fact is that the molecules of different kinds fit more closely together and hence occupy less space.

Facts Which Indicate That Molecules Have Definite Properties. Motion. If some liquid of characteristic odor, like ammonia or ether, is poured out in one place in a quiet room, in a short time the odor can be noticed in all parts of the room. Some of the molecules must have broken away from the attractions of the other molecules. It is assumed from this fact that molecules must be in rapid motion. This mingling of molecules due to motion is called **diffusion**. In the evaporation of a liquid some molecules must have a greater velocity than others because they escape while some remain much longer in the liquid state. If the molecule can acquire sufficient velocity to overcome the attraction of its fellows for it, it leaves the liquid not to return (Fig. 49). Many others having less velocity leave the surface of the liquid but are pulled back by the attraction of the many molecules beneath it. If a stone were thrown upward with a velocity of 50,000 ft. per second, it would leave the earth and be lost in interstellar space. Such an event is highly improbable.

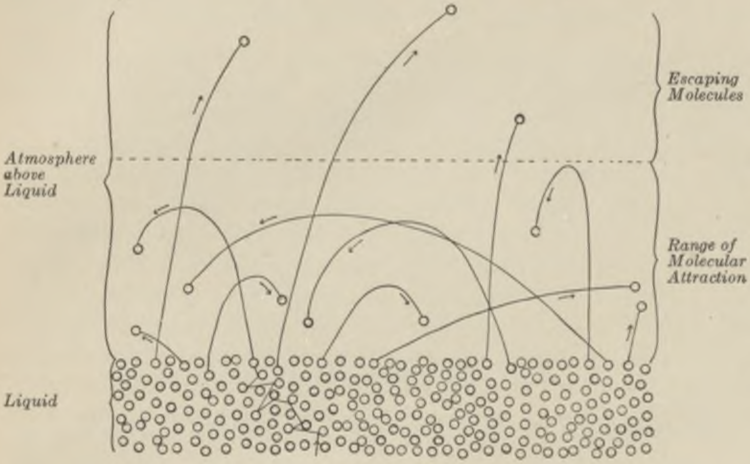


Fig. 49. Evaporation.

If some liquid bromin is put in a tall covered vessel so that the air currents cannot disturb it, the heavy bromin vapor from the liquid can be seen slowly rising through the lighter air in the vessel. This shows that the movement of the molecules is greater in force than gravity.

If now the bottom of the vessel is warmed, the vapor is seen to rise faster than before. This shows that a rise of temperature increases the motion of the molecules.

The average velocity of molecules of one substance seems to be different from that of another. This may be shown by surrounding a porous cup, sealed to a funnel, with hydrogen gas (Fig. 50). At first some of the air in the cup bubbles out through the tube in the water into which the tube is placed. Soon the bubbles cease to appear. If now the surrounding hydrogen gas is removed, water rises in the tube for some distance and after a time sinks back into the dish below. An explanation consistent with these facts is that the more rapidly moving hydrogen molecules enter the cup through the pores faster than the slower molecules of the gases of the air can move out. This increases the number of the molecules in the cup and hence the pressure. Then some of the gases pass out through the tube until the pressure is equalized inside the cup and out. At length the cup becomes filled with a mixture of the hydrogen and air of exactly the same proportions as that in the surrounding vessel, and the movements of gases in and out are the same in quantity. Finally when the outside supply of hydrogen is removed, the more rapidly moving hydrogen molecules inside the cup move out faster than the air molecules move in to take their places. This causes less pressure inside the cup so the water is forced up the tube by the greater atmospheric pressure. As the hydrogen escapes from the cup the water in the tube slowly falls as the air enters through the pores of the cup.

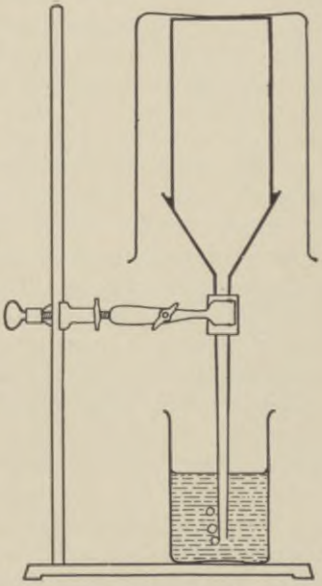


Fig. 50. Diffusion of gases.

Size and Weight of Molecules. Graham passed hydrogen gas successively through a number of very fine porous partitions. The final gas was identical with that which he started with. Had the molecules been of different sizes, the larger would have been left behind and the final gas would have had different properties from that started with. Graham concluded that all molecules of the same gas have the same size. From the similarities in the properties of different gases it is believed that the molecules of all gases do not differ greatly in size.

The chemist, Stas, very carefully compared the weights of equal numbers of molecules of the same substance taken from different sources, and in every case he found the weights to be the same. But when he compared the weights of equal numbers of molecules of different substances, he found the weights to be very different. His conclusion was that molecules of the same substance always had the same weight, but that molecules of different substances had different weights. Molecular velocity seems to be closely related to molecular weight, the lighter molecules having the greater velocity.

The Pressure of Gaseous Molecules. The pressure of solids and liquids on the walls of containing vessels is due mainly to gravity. But gases exert an upward pressure as well as downward on an inclosing wall. To understand how the molecules are responsible for this pressure, think of a swarm of tireless angry bees shut within a glass vessel. As the bees flew furiously about they would strike blows against the walls of the containing vessel. If their number were legion instead of a few hundred, and if they flew one-fourth mile per second instead of a few feet, the number of blows against the walls of the containing vessel would be enormously greater; so great in fact that there would be no perception of separate blows but a steady continuous pressure, which is the fact that is observed. If the molecules are shut in half the former space, this half volume would contain as many molecules as the whole volume. Being half as far apart and so twice as many of them in the same space, the pressure would be doubled. This is exactly the case when a volume of gas is compressed to one-half its former volume; the pressure of the gas is doubled if the temperature is kept constant. This inverse proportionality between the volume and the pressure of a gas is known as **Boyle's Law**.

If some water is tightly inclosed in a vessel and is heated strongly, it is possible to burst a very strong vessel by the expanding steam. Heating the molecules increases their velocity, as was noted before, making them strike more blows per second until the resulting pressure bursts the vessel. Yet when properly controlled the multitudinous blows of individual molecules perform the daily tasks of the civilized world. This is the force utilized by the steam engine, which turns the wheels of industries and commerce. All this tremendous amount of energy annually used must pass through the form of molecular motion. So regularly does this change of velocity, and hence the change of pressure, follow the temperature that we have a sure means of controlling this great force for the use of man.

A volume of a gas measured at 0° C., is found to gain or lose $\frac{1}{273}$ of that volume for each degree that it is warmed or cooled. At 273° C. its volume is doubled, and at 273° below zero its volume, by calculation, should be zero. But all gases escape this annihilation by liquefying before the temperature has fallen to this point. To have zero volume all molecular motion in a gas would have to cease, so that it would not tend to spread out and occupy space. If there were no molecular motion at this low temperature, there could be no gas pressure and no heat. From these considerations the temperature, − 273° C., is called the **absolute zero** of temperature, and the temperatures reckoned from this starting point are called absolute temperatures. They are always 273° greater than the corresponding centigrade temperatures. Since, then, 0° C. = 273° A., it may be seen that both the volume and absolute temperatures increase or decrease $\frac{1}{273}$ of their value at 0° C. for each degree rise or fall of temperature. Therefore the volume of a gas is directly proportional to its absolute temperature, the pressure being kept constant. This relation is known as **Charles' Law**. Temperature is nothing more than molecular velocity or vibration.

States or Conditions of Matter. The greater the velocity of molecules the less will be the force with which they cling to each other, or cohere. In solids cohesion is greater than in liquids, for considerable force is necessary to separate the molecules from one another, while in liquids it is very easy to pull part of the liquid from the remainder of it. When we think of the fact that heat is all that is needed to change a solid — ice for example — into the liquid form, we can understand why. The molecular velocity has been accelerated until cohesive force seems less by comparison. In gases the cohesion of molecules is negligible because of the greater molecular velocity and the distances between the molecules. It may be seen then that states of matter — solid, liquid, and gas — are conditioned by temperature and hence a function of it.

Composition of Matter. The molecular theory makes clear the difference between the elements, mixtures, and compounds. In an element the molecules are all alike, and the molecules contain only one kind of atoms (Fig. 51). The molecules of a compound are also all alike, but differ from the molecules of an element by being made up of two or more different kinds of atoms. In a mixture the molecules are not all alike. There are molecules of two or more different substances present, and they may not be uniformly mixed.

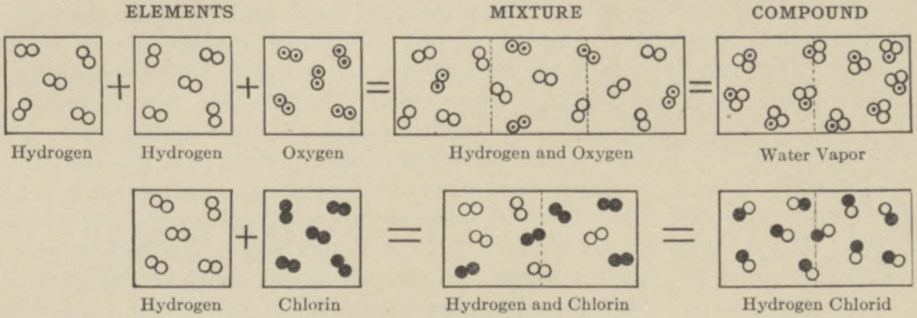


Fig. 51. Constitution of matter.

OUTLINE

The fundamental assumptions regarding the constitution of matter.

Facts and Inferences

1. Divisibility.
2. Compressibility.
3. Loss of Volume by Mixtures.
4. Evaporation.
5. Diffusion with the Effect of Heating.
6. Size.
7. Weight.
8. Pressure. Boyle's Law.
9. Temperature. Charles' Law.
10. Cohesion and States of Matter.
11. Composition of Elements, Mixtures, and Compounds.

Theory

The theory is the explanation and relation of the facts and inferences stated in the opposite column.

Write the theory in your own words.

CONDITIONS THAT INDUCE CHEMICAL ACTION

Since very few substances are put to use by man in the form in which they are found, it is necessary that many substances should be changed and refined by chemical means. In the great chaos of the formation of the earth the chemically active elements were united with others to form compounds. As a rule most naturally occurring compounds are stable and not easily changed, so refining methods are expensive in time and effort. Frequently, enormous urging is necessary to carry out a desired chemical change. In most cases some form of energy is applied to the substance to be changed, although other means are sometimes possible. The use of **heat** is very common. **Electricity** is taking its place in many cases for two reasons. The ease with which it can be transmitted, stored, and turned off or on when wanted makes it most convenient; and it is possible to carry along a chemical action easily with it, that either could not otherwise be accomplished or only with a very great intensity of heat.

The molecular theory explains how the application of heat starts a chemical action. If a substance reacts readily, with another, its molecules contain much energy and are generally unstable — easy to decompose. If action, ever so slight, can start by simple contact, the energy liberated by the chemical action will hasten the action sometimes to the degree of an explosion. Simple contact is insufficient, however, for the majority of cases where chemical changes are desired. While heat is being applied, the increasing molecular motion causes more violent collisions with adjoining molecules. At some stage of temperature the bond of attraction holding the constituent parts of the molecule is broken by the violence of the collision, leaving the parts more or less separated. If the energy set free by the decomposition is in the form of heat, the chemical action may now continue without exterior aid at a rate that depends on the quantity of heat liberated. All such reactions are called **exothermic** because of the energy given off. These reactions furnish most of the energy used in the world to-day, in fact all except that which is derived from water wheels and windmills. The compounds formed by exothermic reactions are stable, mainly because of the energy lost in their making. As a rule it may be said that this energy must be restored to obtain the reacting substances in their original form.

In some cases the energy lost is too slight to be noticeable. In other cases instead of liberating energy a chemical change may cause it to be absorbed. In this case energy must be applied continually to keep the action going. Reactions of this type are called **endothermic.** The compounds formed are generally unstable.

The use of **light energy** by man to cause chemical changes is limited mainly to photographic work with silver compounds. Nature, however, uses light lavishly. All green plants are dependent on sunlight for their growth. The use of **mechanical motion** for causing chemical action is confined to " setting off " explosives. Other means besides the application of energy are effective in causing chemical changes. Substances called **catalytic agents** have this power. While they cause other substances to change by some mysterious influence, they are themselves unchanged. In **solution** a substance reacts more readily because of the finely divided form when dissolved.

Study Topics: Duncan, *The Chemistry of Commerce:* "Catalysis," pp. 15–41; "High Temperatures and Modern Industry," pp. 88–108.

Bird, *Modern Science Reader:* "The Chemical Regulation of the Processes of the Body," pp. 181–194; "Modern Explosives," pp. 17–28; "Marvels of a Plant's Growth and the Chemistry of Decay," pp. 103–114.

References: M. & L., pp. 14–16; Smith, pp. 155, 165–167, 238; B. & W., pp. 413–424; McP. & H., pp. 59–61; H. & S., pp. 22, 289–292.

CONDITIONS THAT INDUCE CHEMICAL ACTION

EXPERIMENT

Purpose: (After reading carefully the entire experiment including the previous page, write the purpose of this experiment on the record sheet following the proper heading.)

Apparatus: Tile, battery of 4 dry cells with platinum or carbon plates, hammer.

Materials: Potassium permanganate, glycerin, strontium nitrate, shellac, potassium chlorate, silver nitrate solution, copper sulfate, sulfur, cream of tartar, sodium bicarbonate, manganese dioxid, splinter.

I. **Chemical Action by Simple Contact**. Powder about 10 g. of potassium permanganate by grinding it in a mortar. Make a cone-shaped pile of it on a tile in the hood. With the end of a pencil make a deep depression in the center of the cone, and fill it with glycerin. Observe for a minute and record the evidences of a chemical change (?).

II. **Heat**. Pulverize separately 1.5 g. of potassium chlorate, 9 g. of strontium nitrate, and 3 g. of shellac. Grind the first substance **gently** in the mortar; do not pound it. All materials should be dry. Mix the substances thoroughly on paper, then make a cone-shaped pile of the mixture on a tile in the hood, and heat by directing a burner flame on it (?). Note and record all evidences of a chemical change. The mixture is used to make red fire. If barium nitrate is substituted for the strontium nitrate in the same quantity, it makes green fire.

III. **Electricity**. Dip two plates (electrodes) connected to battery by copper wires into a solution of copper sulfate. Watch the surfaces of the electrodes for evidences of a chemical change (?). After a few minutes remove electrodes from the solution and examine them (?). To restore them to original condition dip them into nitric acid.

IV. **Light**. To about 5 cc. of silver nitrate solution in a test-tube add about 5 cc. of hydrochloric acid. Pour contents of tube on a filter fitted to the funnel (Introduction, p. 4). Throw away the filtrate. Remove the filter with the silver chlorid carefully, and spread it out in the sunlight for a few minutes (?).

V. **Mechanical Motion (Percussion)**. Powder about 2 g. of potassium chlorate, as above, with care. **Using the fingers** mix with it on paper about 1 g. of sulfur powder (flowers). Examine for chemical action (?). Put **only** a pinch of the mixture on a brick or stone and strike it a vigorous blow with a hammer (?).

VI. **Solution**. Mix about 4 g. of cream of tartar (potassium bitartrate) with about 2 g. of sodium bicarbonate. Examine for chemical action (?). Pour water on the mixture (?).

VII. **Catalysis**. Put about 3 g. of potassium chlorate in a 15-cm. tube and support it in a vertical position. Heat with a flame **not touching** the tube until the crystals melt and bubbles rise (effervescence) in the liquid. The bubbles show that a gas (oxygen) is escaping. Test for it by holding a long splinter with a glowing end **just above** the liquid, not in it (?). Stop heating and wait until effervescence has ceased. When the action has stopped and the substance is still liquid, drop a pinch of powdered manganese dioxid into it (?). Test for oxygen again (?). The manganese dioxid is not changed chemically in this action. Remove the residue from the tube with water.

Questions: 1. What is a spontaneous chemical change?

2. Give uses of mixtures such as were made in Part II.

3. Give uses of electricity as shown in Part III.

4. Give uses of silver salts as shown in Part IV.

5. What substances are "set off" (exploded) by jarring or detonation?

6. The materials in VI are the chief ingredients of baking powder. Explain its use as shown by the chemical action.

7. Name all the forms of energy whose use is shown in this experiment.

VIII. **Optional**: Sun–**light** and **heat**–plants–**chemical energy**–wood and coal–**heat**–steam engine–**mechanical motion**–dynamo–**electricity**–incandescent lamp–**light**.

Write a short account of the methods of transforming energy as suggested above.

OXYGEN

Oxygen is the most abundant element. Since life depends on its chemical activity, and the greater part of the energy used by the industrial world at the present time is generated by its chemical action with combustible substances, it is also the most important element. Although most of the properties of air are due to the free oxygen it contains (one-fifth by volume), the inactive gas, nitrogen, which comprises most of the remaining portion of the air, has such a great retarding influence upon the activity of the oxygen that the latter must be obtained in a pure state when its properties are to be studied.

Pure oxygen for commercial purposes is now most cheaply made by liquefying air and then letting the nitrogen escape by evaporation; but when small quantities are needed, it is cheaper to decompose one of its compounds. In the latter case, four points must be taken into consideration: (1) the cost of the compound used, (2) the quantity of oxygen which may be obtained from it, (3) the amount of energy required to liberate the oxygen, and (4) the possibility of using the residual substance in some other process, or of changing it back again to the original substance to be used over again. Sand (silicon dioxid, 53.3 % oxygen) is very cheap, but requires so much energy for its decomposition that it cannot be used. Hydrogen peroxid contains 94 % oxygen, and gives up half of it at ordinary temperatures along with considerable quantities of heat; but its cost is too great to allow its use.

When chemical action occurs between oxygen and other elements, compounds called **oxids** are formed. They are distinguished by prefixing the name of the other element, and often a syllable is prefixed to the word **oxid** to give some indication of the proportion of oxygen present; *e.g.* carbon dioxid, sulfur dioxid, phosphorus pentoxid, iron oxid, magnesium oxid, etc. When compounds burn in oxygen, the oxids of the elements in the compound are formed just as if the elements were burning separately.

Pure oxygen is used in pulmotors to revive those asphyxiated by gas or water; to maintain life in those who are very weak; to keep the air in submarines in a respirable condition; and to replace air as a supporter of combustion when very high temperatures are required, as in the oxy-hydrogen and oxy-acetylene flames.

Study Topics: Ramsay, *The Gases of the Atmosphere.*
Thorpe, *Essays in Historical Chemistry*: "Joseph Priestley."
Bird, *Modern Science Reader*: "The Commercial Production of Oxygen"; "The Ox-hydric Process of Cutting Metals."
Bartow, "Purification of Sewage by Aëration in the Presence of Activated Sludge," *The Journal of Industrial and Engineering Chemistry*, Vol. 7 (1915), p. 318; and Vol. 8 (1916), pp. 15, 17.

References: H. & S., pp. 12–27; B. & W., pp. 25–48; M. & L., pp. 22–33; Smith, pp. 23–31; McP. & H., pp. 14–23; Brownlee, pp. 22–29; Newell, pp. 11–20.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Apparatus: 15-cm. test-tube fitted with 1-hole rubber stopper, glass and rubber tubing as shown in Figure 52, burette clamp, stand, slotted disk, water pan, five 250-cc. jars (wide-mouth bottles), 5 glass covers, combustion spoon.

Materials: Potassium chlorate, manganese dioxid, iron wire, lump charcoal, sulfur, red phosphorus, iron picture wire, asbestos paper, cardboard.

I. **Preparation.** (Two students working together.) Mix thoroughly on paper about 12 g. of potassium chlorate and about one-third the bulk of manganese dioxid. Slide it into a 15-cm. test-tube, hold the tube horizontally and tap it with the finger so as to distribute the mixture along the side of the lower half of the tube, leaving space above the mixture to the end of the tube. Attach a burette clamp (with care not to crush the tube) near the mouth of the tube, and adjust the clamp so that the mouth of the tube is slightly below its closed end. Insert the rubber stopper carrying the delivery tube as in Figure 52. Into one of the jars insert a circular disk of cardboard which is just slightly smaller than the inside diameter of the bottom of the jar (bend the disk but do not crease it in inserting it). Fill all the jars completely with water, cover them with the glass plates (holding the glass plates firmly against the mouths of the jars during the operation), and place them inverted in a water pan containing about 5 cm. of water. Remove the glass plates without letting air into the jars, and set one of them on the slotted metal disk which rests on the bottom of the water pan. Place the end of the delivery tube in the slot of the disk so that the gas may rise into the jar. Let one student heat the mixture with a moving flame, regulating the efflux of gas by the intensity of the heating. **Heat slowly** and remove the flame when the gas comes too fast. The other student changes the jars as they are filled, holding the glass plates to the mouths before raising them from the water and setting them mouth-down on the table till used. Fill 5 jars, but leave an inch of water in the jar containing the cardboard and use this jar of gas in II (*e*). Remove the delivery tube from the water **before** the heating is discontinued. Describe the action in the test-tube and make a diagram of the apparatus.

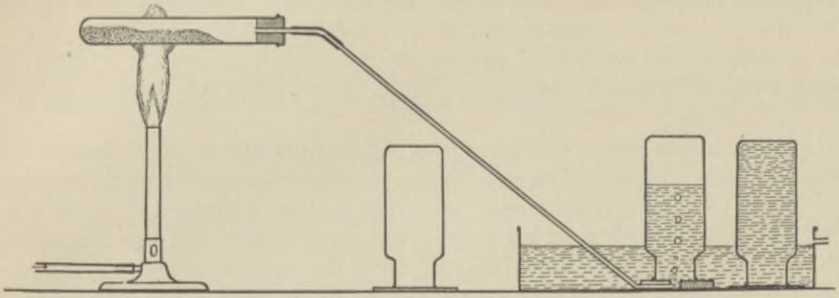


FIG. 52. Decomposition of potassium chlorate.

II. **Properties.** (*a*) Record the physical properties: color (first jar), odor (any noted is due to impurities), solubility (infer from the manner of collection).

(*b*) Wind one end of a 30-cm. length of wire around a small piece of charcoal, heat the charcoal to burning, and lower it into a jar of oxygen (?). Compare the burning in oxygen with the burning in air. Quench the charcoal in water and drop it into the waste jar.

(*c*) Line a combustion spoon with asbestos paper and put in it 0.5 cc. of sulfur. Ignite the sulfur with the burner flame turned down on it, and lower it into a second jar of oxygen (?). Compare the burning in oxygen and in air (?).

(*d*) Clean and reline the spoon, place in it a bit of red phosphorus (size of small pea), ignite it as before, and insert it in another jar of oxygen (?). Compare the burning as before (?).

Note: **Always burn all the phosphorus from the spoon before you put it away.**

(*e*) Select the jar containing the cardboard. Heat an untwisted end of a piece of picture wire, dip it at once into some powdered sulfur, ignite the sulfur, and quickly thrust the wire into the jar of oxygen holding it directly over the cardboard. If the sulfur is extinguished, try again. The result is very spectacular. Describe it (?).

Clean all the jars and set them away inverted to dry. Remove the residue in the test-tube with water. Clean the table top and put all apparatus away.

Questions: 1. What is the reason for the difference between the combustions in oxygen and in air?

2. Write word equations for the preparation and the combustions.

3. What would happen if the atmosphere should suddenly become 100 % oxygen?

HYDROGEN

The earth and the lower part of the atmosphere contain very little of this element in the free or uncombined form. Since 1900 it has been definitely learned that at a height of 60 miles the atmosphere is chiefly hydrogen.[1] It is the lightest known substance, so there is good reason for its position in the atmosphere. Under standard conditions of temperature and pressure, the hydrogen molecules move at the rate of 5571 feet per second. This is over a mile per second, and exceeds the velocity of all other molecules. This is further reason why gravity is unable to pull hydrogen closer to the earth.

In 1867 Graham analyzed a meteorite and found that it contained nearly three times its volume of hydrogen gas. While this white-hot metal was flying through the upper atmosphere, the hydrogen was probably absorbed. The spectroscope has shown that all stars contain this gas and the hottest ones are composed chiefly of it. Our own sun has immense quantities of it. Astronomers have observed the hydrogen flames towering aloft to a height of 100,000 miles.

In combined form hydrogen is quite plentiful as an important constituent of water, acids, petroleum, and plants. Since the hydrogen of the atmosphere is inaccessible, this element must be extracted from its compounds for the uses in the arts and industries. Certain active metals like sodium are active enough to react with water, setting free some of its hydrogen. But these metals are costly, so this method of preparing hydrogen on a commercial scale is impractical. If a cheaper and less active metal is used, the action on water is too slow, so an acid is used instead of water. Hydrochloric and sulfuric acids are the ones commonly used. In actual practice a mixture of acid and water is used. While the water itself takes no direct part in the chemical action of acids, it renders the acid more active. In a water solution some of the constituent hydrogen seems to be separated from the acid molecule. The majority of metals have the power to drive this form of hydrogen out of the solution as a free gas, while the metal takes its place in the solution. When the water is evaporated, the metal combines with the remainder of the acid.

$$\text{Iron} + \text{hydrogen sulfate} \longrightarrow \text{hydrogen} + \text{iron sulfate.}$$

As the cost of generating electricity has been steadily decreased, the method of separating hydrogen from water by electrolysis is being more used. In the next ten years this method may take the place of all others.

Study Topics: Thorpe, *Essays in Historical Chemistry:* "Henry Cavendish," pp. 60–79.
Ramsay, *The Gases of the Atmosphere:* "Phlogisticated Air," pp. 121–179.
Martin, *Triumphs and Wonders of Modern Chemistry:* "The Element Hydrogen," pp. 82–107.
References: B. & W., pp. 111–119; H. & S., pp. 38–44; M. & L., pp. 45–58; Smith, pp. 38–46; McP. & H., pp. 27–32; Brownlee, pp. 33–37; Newell, pp. 22–28.

EXPERIMENT

Purpose: State the purpose on record sheet after the proper heading.

Apparatus: 125-cc. wide-mouth bottle with 2-hole rubber stopper to fit, slotted disk, 5 glass gas-collecting jars with plate covers.

Material: Mossy zinc, hydrogen sulfate (sulfuric acid), pine splinter, copper sulfate solution, alkaline potassium permanganate solution.

I. **Preparation**. Fill the small generating bottle one-fourth full of mossy zinc. Rub a little soap on the wetted thistle tube stem and thrust it with a twisting motion through the stopper of the generator far enough so it shall extend into the layer of zinc when the stopper is in place. Push one arm of L-tube just through the other hole of the stopper, and connect the other end with the delivery

---

[1] Dewar, *Address to the British Association*, p. 20 (1902); Erdmann, *Lehrbuch der Anorganischen Chemie*, p. 347 (1906).

tube by means of a short piece of rubber tube. (See Fig. 53 for the arrangement of the apparatus.) Hold the generator securely upright by a clamp. Arrange the water pan to collect the gas over water as was done with oxygen. Have the jars completely filled with water and no air bubbles. Fill the smallest jar first. **Caution! Do not have a burner lighted or any flame at your table until the hydrogen is collected in the jars and the stopper removed from the generator.**

Pour dilute hydrogen sulfate (sulfuric acid) into the generator through the thistle tube until the zinc is well covered. If the action does not proceed rapidly enough in two minutes, add 2–3 cc. of copper sulfate solution. Collect five jars of the gas, remembering the order in which they were filled. Keep the jars inverted until you are ready to use the gas. If the action slows up add more acid. When sufficient gas has been collected, remove stopper from the generator and allow it to stand until after the gas is studied. Then filter the liquid from the generator (Introduction, p. 4) into the evaporating dish until it is three-fourths full. Put dish away securely in locker until the next laboratory period. If crystals have not formed by this time, evaporate one-third of the liquid and set aside again to crystallize. Wash the zinc remaining in the generator and return it to the stock bottle. Put jars away **inverted** to dry. Wipe off the table top and leave all apparatus in order.

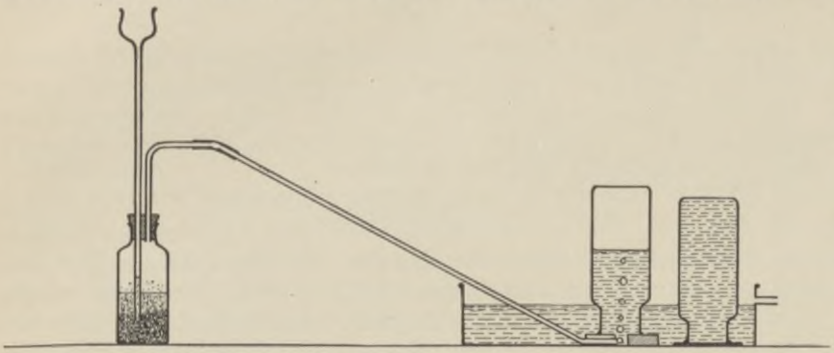


Fig. 53. Replacement of hydrogen by zinc.

II. **Properties.** (*a*) Set the last jar filled on the table in inverted position. Raise it from the table without cover and still inverted, then bring a blazing splinter to its mouth (?). If the hydrogen is pure, there will be little or no noise. Look carefully for a flame at mouth of the jar (?). Thrust the blazing splinter up into the gas (?).

Questions: 1. Does hydrogen burn?

2. Does hydrogen support combustion?

(*b*) Test with a flame the first jar filled and explain the different behavior of the burning gas (?).

3. Under what conditions is hydrogen dangerous when ignited?

(*c*) Examine the hydrogen for any color (?). Try its odor before and after shaking about 30 cc. of alkaline potassium permanganate in a jar of it to remove impurities (?). From the manner of collecting the gas consider whether it is soluble or not (?). Place a jar of the gas in an upright position and another in inverted position on the table. Lift the inverted jar from the table without cover, keeping it so, and at the same time remove the cover from the upright jar. Wait 2 minutes, and test each jar with a blazing splinter for hydrogen. State result in each case (?).

4. Compare the density (specific gravity) of hydrogen and air.

5. Complete the following word equation.

$$\text{Zinc} + \text{hydrogen sulfate} \longrightarrow \; ? \; + \; ?$$

6. Since zinc is an element, from what did the hydrogen come?

7. What became of the zinc used up by action of acid?

8. What is the name of the substance to be crystallized?

9. Why is this chemical action classed as a " substitution " or " replacement "? (See equation in 5.)

10. State two chemical and four physical properties of hydrogen.

EQUIVALENT WEIGHTS

In the preparation of hydrogen by the action of zinc on sulfuric acid, some of the zinc in the generator was used up in the chemical action. The zinc displaced hydrogen from the acid and was formed into a new substance, zinc sulfate, entirely different in its properties from either the zinc or the sulfuric acid. When the quantity of zinc taken is entirely used up, no more hydrogen is set free, even if there is an abundance of acid. This fact indicates that there is a **definite relation** between the weight of zinc used and the weight of the hydrogen displaced.

There are a number of other acids which may be used instead of sulfuric acid as sources of hydrogen. Some of them react with the zinc more vigorously than the sulfuric acid does, while others act very slowly. However, experiment shows that if the same weight of zinc is used with each acid, it will in each case displace the same weight of hydrogen. Thus 32.7 grams of zinc will always displace 1 gram of hydrogen. This weight of zinc is said to be **chemically equivalent to 1 gram** (**1.008 g.**) **of hydrogen**, or more briefly, the **equivalent weight of zinc is 32.7 grams.**

When a certain weight or volume of hydrogen is to be generated, the weight of zinc needed can be very easily calculated by means of the equivalent weight. Thus 5 grams of hydrogen will require 5 × 32.7 or 163.5 grams of zinc. When a certain volume of hydrogen is required, use is made of the fact that 1 gram of hydrogen occupies 11.2 L. at standard temperature and pressure (S. T. P.). The volume of hydrogen required is first reduced to the volume of dry gas at S. T. P., and then a proportion is written and solved as follows:

$$\frac{x \text{ (weight of zinc required)}}{32.7} = \frac{\text{volume needed (S. T. P.)}}{11.2}.$$

Other metals may be used in place of the zinc, some being more active than the zinc and some slower. Each metal has a certain equivalent weight, the values ranging from 7 grams for lithium to 204.1 grams for thallium.

The term "equivalent weight" is extended to include the non-metallic elements and even compounds. It is **that weight of an element or compound which will displace, liberate, combine with, or enter into chemical action with 1 gram** (**1.008 g.**) **of hydrogen, or with the equivalent weight of any other element.** The great importance of these equivalent weights lies in the fact that they can be used to represent the weight ratios in which substances enter into any kind of chemical action. The equivalent weight of oxygen is 8 grams, since that weight will combine with 1 gram of hydrogen. When oxygen combines with the metals, it is in the ratio of 8 grams of oxygen to one equivalent of the metal. When the metals are deposited from their solutions by electricity, as in nickel-plating, it is found that equal quantities of electricity in equal times will deposit weights of the various metals in proportion to their equivalent weights (Faraday's Law). Whenever chemical actions take place, **the weights of the substances reacting are always in the same ratio as their equivalents.**

If there is a definite ratio between the weights of combining elements, there must be a definite ratio between the weights of the constituent elements of the compound. This is the **Law of Definite Proportions**, and may be more fully stated as follows: **Every sample of a given compound is composed of the same elements in the same proportion by weight.**

THE CHEMICAL EQUIVALENT OF MAGNESIUM

EXPERIMENT

Purpose:..

Apparatus: 100-cc. graduated gas-measuring tube, deep glass jar, thermometer, barometer.

Materials: Magnesium ribbon (must be clean and bright), thread.

Magnesium ribbon is so nearly uniform in size that its weight, for this experiment, may be considered proportional to its length. Five meters of clean, bright ribbon should be measured off exactly and weighed carefully. From this data the weight of 1 cm., and then the weight of the length used can be computed. If greater accuracy is desired, the magnesium may be weighed.

FIG. 54. Replacement of hydrogen by magnesium.

Measure off exactly 14 cm. of ribbon, roll it into a small coil and tie a 20-cm. thread to it. Pour about 5 cc. of conc. hydrochloric acid into the gas-measuring tube. Hold the tube in an inclined position and pour water on the acid until the tube is full. Note: All water used in this experiment should be previously warmed to room temperature. Place the coil of ribbon in the mouth of the tube, letting the thread hang over the edge. Close the tube with the thumb or a small glass plate, invert the tube, and lower the mouth under water in the evaporating dish placed in the water pan. Fasten the tube upright in a clamp, allowing the mouth of the tube to rest upon the bottom of the dish (Fig. 54).

The heavier acid may be seen settling down through the water to the ribbon, which it attacks, liberating hydrogen. If small bits of ribbon break loose from the coil, tap the tube frequently with the finger so as to prevent their sticking to the walls of the tube and so escaping solution. When the action has ceased and the ribbon has been consumed, cover the end of the tube tightly and remove the tube without letting out any water. Carry the tube to a deeper vessel containing water and adjust it in a clamp so that the levels of water inside the tube and out are the same. Let the tube be untouched for 5 minutes while taking the barometer reading and the temperature of the water in the vessel. Record these in the table below, then read carefully, and record the volume of the hydrogen in the tube. If in doubt about any reading, ask the instructor to verify your readings. The tube may now be put away.

Finish all computations, handing in the paper showing all the work. Subtract the water vapor pressure from the barometer reading, and then compute the volume of the hydrogen under standard conditions. Compute the weight of this hydrogen if 1 L. of it weighs 0.09 gram. Now, if this weight of hydrogen was displaced by the weight of the magnesium used, how many grams of magnesium are needed to liberate **one gram** of hydrogen? Compute it. The result will be the chemical equivalent of magnesium.

RECORD OF RESULTS

Weight of 5 m. of magnesium ribbon .... g.
Weight (computed) of 1 cm. magnesium ribbon .... g.
Weight (computed) of 14 cm. magnesium ribbon .... g.
Volume of hydrogen obtained .... cc.
Temperature observed .... ° C.
Barometric pressure observed .... mm.
Water vapor pressure (see p. 233) .... mm.
Gas pressure of hydrogen in the tube .... mm.
Volume of hydrogen at 0° C. and 760 mm. (write formula) ...... cc.

Weight of hydrogen .... g.
Chemical equivalent of magnesium .... g.
Accepted value of the equivalent .... 12.16 g.
Error .... g.
% error (error × 100 ÷ 12.16) .... %

WATER

In a previous experiment hydrogen gas was burned in the air. The product formed was kept in an invisible gaseous form because of the great heat of the hydrogen flame. In this experiment special means will be used to collect and cool the product for observation. Many combustible substances are partly composed of hydrogen. When they are burned, this same product is formed. You may have noted the film of moisture that collects on the cold chimney of a lamp when it is first lighted; or the drops of liquid on the bottom of a vessel filled with cold water when put over a gas flame. A similar effect may be had by holding a cold dry bottle inverted with the mouth over a candle flame.

When Priestley burned some hydrogen in a glass vessel, he noted that "the inside of the glass, although clear and dry before, immediately became dewy." Cavendish thought this fact "well worth examining more closely." He burned hydrogen with air in a long glass tube and collected a liquid, which he identified as water.

The ancient philosophers believed water to be an element. Their methods were quite inadequate for the task of decomposing so stable a compound as water. Lavoisier is said to be the first to accomplish this by passing steam through a red-hot gun barrel. Then water has such entirely different properties from those of its constituents that no one but a chemist would learn the facts which would enable him to discover the composition of water as Cavendish did.

After identifying the product of this experiment, its composition is to be studied by decomposing it with the aid of the electric current. The results of this work will show what elements of the air have combined with the hydrogen as well as the probable proportions by volume of the gases combined. Since it is difficult to compare the weight of the liquid decomposed with that of the gases formed from it, such comparison will not be attempted. But to be conclusive, an analysis of a substance must show that the sum of the parts found makes up 100 % of the quantity of the substance analyzed. It must be certain that no constituent of the liquid has been overlooked in the analysis. This can be accomplished by taking portions of gases in the proportions as found by analysis and combining them to see if the same substance, water, is formed. This direct combination of elements is called **synthesis**.

In general the composition of a substance is established by an analysis for the different elements present as constituents. This is called **qualitative** analysis. Next, the quantity of each element present is determined by volume or by weight. This is called **quantitative** analysis. These data enable a chemist to know whether a substance is a mixture or a compound, and if the latter, what its formula is. The quantitative analysis of a substance by weight is termed the **gravimetric** analysis. For finding the composition of water this method is more direct, but it requires more skill and time.

Study Topics: Martin, *Triumphs and Wonders of Modern Chemistry*, pp. 82–106.
Thorpe, *Essays in Historical Chemistry*: "Discovery of the Composition of Water," pp. 90–97, 106–122, 123–148.
References: H. & S., pp. 45–55; Newell, pp. 23, 49–53; B. & W., pp. 100–107; McP. & H., pp. 28, 29, 46–51; M. & L., pp. 45, 49, 61; Smith, pp. 47, 48; Brownlee, pp. 37–44.

EXPERIMENT

Purpose: Write the purpose of this experiment on a separate sheet.

Apparatus: I. Hydrogen generator with jet tube and drying tube, a bottomless liter flask, Liebig condenser, stand with large clamp. II. Hofmann's or equivalent apparatus for electrolysis of water, 4 storage or dry cells, splinter. III. Eudiometer tube, spark coil and battery, apparatus used in II.

Materials: Zinc, mercury, anhydrous calcium chlorid.

I. **Collecting the Product of Burning Hydrogen**. Set up the condenser as a reflux condenser (Fig. 55). The hydrogen gas is obtained from a generator as in a previous experiment for the preparation of hydrogen. This gas passes through a drying tube containing calcium chlorid, which absorbs any moisture that may be carried over from the generator by the gas. Ignite the gas at the jet after the action has continued for 5 minutes and it is certain that all air has been swept out of the generator. Introduce the jet, after lighting, through the opening in the bottom of the flask attached to the condenser, and collect the condensed liquid in any convenient vessel. Observe and record the properties of the liquid (?). Test it with cobalt chlorid paper (?). Compare the effect with that of water on this paper (?). Identify the liquid (?).

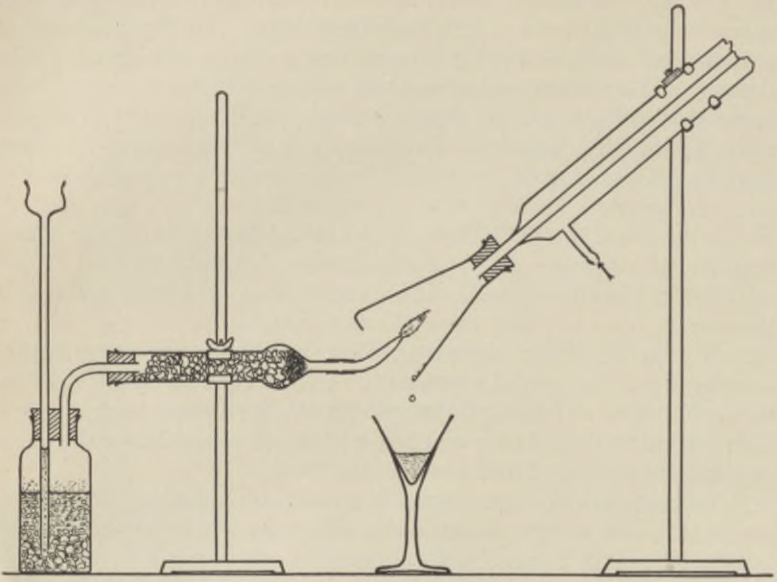


FIG. 55. The product of burning hydrogen.

II. **Analysis of Product by Electrolysis**. Fill the Hofmann apparatus with water containing about 5 % of sulfuric acid. The acid makes a solution which conducts electricity so that the water can be decomposed by the current which passes through it. At the conclusion of the procedure the solution will still contain the same quantity of the acid as at first, so it need not be further considered at this time. The battery of cells is connected to the electrodes of the apparatus. The metal plate at which the larger quantity of gas collects is called the **cathode**, and the other the **anode**. Describe the action (?).

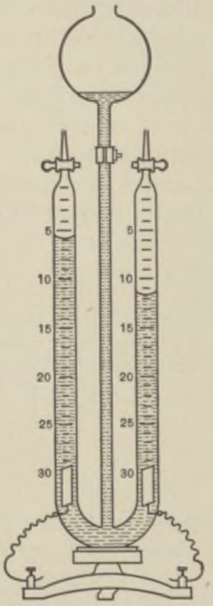


FIG. 56. Hofmann apparatus.

Questions: 1. Where do the bubbles of gas appear?

2. Read and record volume of gas at cathode (?). At anode (?).

3. Test the larger volume with a flame (?). What gas is it?

4. Test the remaining volume of gas with a glowing splinter (?). What gas is it?

III. **Synthesis of Product from Measured Volumes of Gases**. (Optional.) Fill the eudiometer with mercury. Introduce 10 cc. of oxygen from the Hofmann apparatus and adjust the surfaces of the mercury until they are the same. Read the volume of the oxygen carefully (?). In a similar manner introduce about 30 cc. of hydrogen; level and read volume as before (?). An excess of one gas is used to reduce the force of the explosion. A spark, obtained from the induction coil, is made to pass from one wire to the other inside the tube, igniting the mixture of the gases. After the explosion level the liquid surfaces again and read the volume of the residual gas. The oxygen is entirely used up.

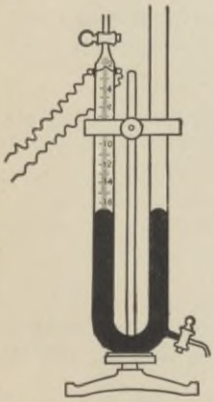


FIG. 57. Ure's eudiometer.

5. Test the residual gas with a flame (?). What is it?

6. Give the ratio of the volumes of gases combining, hydrogen to oxygen (?).

7. Compare this ratio with the one found by electrolysis in Part II (?).

8. State briefly the facts of this experiment which show the composition of water.

THE THEORY OF ATOMS

It has been found by the study of the composition of water that two volumes of hydrogen combine with one volume of oxygen to make two volumes of water vapor. The volume relations of other combined gaseous elements, as found by experimental analysis. are summarized below for further study, with the volumes found by the analysis of water.

VOLUME RELATIONS OF COMBINED GASEOUS ELEMENTS

2 vols. hydrogen + 1 vol. oxygen ⟶ 2 vols. water vapor.
1 vol. hydrogen + 1 vol. chlorin ⟶ 2 vols. hydrogen chlorid.
3 vols. hydrogen + 1 vol. nitrogen ⟶ 2 vols. ammonia.
1 vol. oxygen + 2 vols. nitrogen ⟶ 2 vols. nitrous oxid.
1 vol. oxygen + 1 vol. nitrogen ⟶ 2 vols. nitric oxid.

The above numbers expressing volumes are seen to be **small whole numbers**. This fact is significant, for it shows that there is a **unit of volume** used in Nature's plan of combining gases. If this were not so, the relative volumes of gases combined would be fractional if expressed in small numbers, or if not fractional they would be large whole numbers if they expressed the relative volumes exactly. In 1808 Gay-Lussac stated the facts of volume relations of gases. This law now bears his name. "**Gases combine in volumes which bear a simple numerical ratio to each other and to the volume of their gaseous product.**"

Since these numbers express only the relation of the volumes of the gases combined, they are equally true for any chosen volume, whether it is a liter, quart, or any other. Just what the size of this unit volume shall be is best decided by considering what volume will furnish the most convenient numbers when it is desired to express the proportions of these combined gases by **weight.**

In scientific work it is customary to use the liter as the standard of volume. If the liter is used as the standard gas volume, and this volume of the gaseous compounds of oxygen at standard temperature and pressure (S. T. P.) is analyzed, 0.715 gram of oxygen or some integral multiple of this weight is always found (see table below). Similarly, 0.045 gram of hydrogen or an integral multiple of this weight is always found in 1 L. of any gaseous hydrogen compound; and 0.625 gram of nitrogen or a multiple of this weight is found in a liter of any gaseous compound of nitrogen.

WEIGHTS OF ELEMENTS FOUND IN 1 L. OF DIFFERENT GASES

Gas	Weight of 1 L.	Weight of Oxygen	Weight of Hydrogen	Weight of Nitrogen
Water vapor	0.805 g.	0.715 g.	2 × 0.045 g.	
Carbon dioxid	1.966 g.	2 × 0.715 g.		
Oxygen	1.430 g.	2 × 0.715 g.		
Hydrogen chlorid	1.628 g.		0.045 g.	
Ammonia	0.760 g.		3 × 0.045 g.	0.625 g.
Hydrogen	0.090 g.		2 × 0.045 g.	
Nitrous oxid	1.965 g.	0.715 g.		2 × 0.625 g.
Nitrogen	1.250 g.			2 × 0.625 g.

The figures representing the weights of the elements found in one liter of any volatile substance are inconvenient decimals, and it is evident that the liter is not an ideal unit volume. If 22.4 L. is taken as the unit volume, all the weights in the above table will be multiplied by 22.4. Using 22.4 L. as the unit volume, the weight of oxygen in any volatile substance is always found to be 16 grams or an integral multiple of it, and the weight of hydrogen, the lightest element, is always a little more than one gram (1.008 g.) or an integral multiple thereof. Twenty other elements have whole numbers to represent their minimum occurrences in the new unit volume, and twenty others are nearly whole numbers.

WEIGHTS OF ELEMENTS FOUND IN 22.4 L. OF DIFFERENT GASES

Gas	Weight of 22.4 L.	Weight of Oxygen	Weight of Hydrogen	Weight of Nitrogen
Water vapor	18.016 g.	16 g.	2 × 1.008 g.	
Carbon dioxid	44.00 g.	2 × 16 g.		
Oxygen	32.00 g.	2 × 16 g.		
Hydrogen chlorid	36.468 g.		1.008 g.	
Ammonia	17.034 g.		3 × 1.008 g.	14.01 g.
Hydrogen	2.016 g.		2 × 1.008 g.	
Nitrous oxid	44.02 g.	16 g.		2 × 14.01 g.
Nitrogen	28.02 g.			2 × 14.01 g.

This table may be extended to include all gaseous compounds and elements. In every case it will be found that all the weights of any one element in its various compounds are represented by the same number or by whole multiples of this number.

In the molecular theory, all matter was considered to be composed of units which were called molecules. These units were assumed to be the same in any one substance, but different in different substances. The facts which warrant this belief were given. The uniform physical behavior of gases led Avogadro to believe that **equal volumes of gases under the same conditions of temperature and pressure contain equal numbers of molecules.** This theory cannot be proved by direct experimental evidence, so it has not been termed a law. Yet it has been tested indirectly by hundreds of experiments. In all these experiments it has never been found to be inconsistent with the facts, so it has been considered practically as a law. Assuming it to be true, then the weights of 22.4 L. of each of the gases in the above table have the same relative values as the weights of single molecules. These weights may then represent the weights of single molecules. Modern researches indicate that there are approximately 630,000,000,000,000,000,000,000 molecules in 22.4 L. of a gas under standard conditions, so that the molecular weights in our system really represent the weight of this enormous number of molecules instead of only one.

The constant minimum weights of the elements found in 22.4 L. of these gases cannot be the weights of molecules, since they are but fractional parts of the weights representing the molecules. Evidently they are units of a lower order that make up the molecule. Chemists call them **atoms.** Since there are always whole numbers of them — each having definite weight — in the molecule, the law of definite proportions may be stated : "**Any given chemical compound always has the same elements in the same proportions by weight.**"

In any compound which contains more than the smallest weight (the weight of one atom) it will be noticed that the larger weight is an integral multiple of the smallest weight. The multiple shows the number of atoms of that element present in the molecule. From these facts Dalton's law of multiple proportions can be understood. **If an element unites with a fixed amount of another element in more than one proportion by weight, the different weights of this element may be expressed by a ratio of small whole numbers.**

The above table also indicates the composition of elementary gases. Since the weight of the Unit Volume (22.4 L.) representing the weight of the molecule is **twice** the minimum weight of that element (atomic weight), each molecule of oxygen, hydrogen, and nitrogen must be composed of two atoms.

The composition of non-volatile substances can also be expressed by means of these atomic weights. Thus hæmatite contains iron and oxygen. Their atomic weights — as determined from their volatile compounds—are 55.84 and 16, respectively. Analysis of hæmatite shows that 111.68 (2 × 55.84) weights of iron are combined with 48 (3 × 16) weights of oxygen.

There are other methods of finding molecular weights which depend upon the properties of substances in solution. Another method of determining atomic weights — and a more accurate one than the method outlined in this discussion — is given in the next experiment.

Summary: 1. Simple ratios of reacting gas volumes show that Nature measures by **units of volume**.

2. Unit volume is best determined by the **simplicity** of relationship between weight and volume.

3. The weight of an element in 1 L. (S. T. P.) of any gaseous compound is a **constant minimum quantity**, or a **multiple** of it.

4. These constant minimum weights of the elements in **1 L.** are inconvenient decimal numbers.

5. The constant minimum weight of the lightest element is 1 (approximately) and many of the others whole numbers, if **22.4 L.** is the **unit volume**.

6. The uniform physical behavior of gases indicates that **equal volumes of gases** (S. T. P.) contain the **same number of molecules**.

7. Therefore the weights of unit volume are **proportional** to the weights of **single molecules** and may be used as such.

8. The constant minimum weights in unit volume are not molecular weights, but are **component parts** of the molecular weights.

9. They are the relative weights of **smaller units** of matter than the molecules. They are called **atoms**.

10. Most gaseous elements have 2 constant minimum weights (atomic weights) in the weight of unit volume (molecular weight). Thus the formulas of the molecules of hydrogen, nitrogen, oxygen, chlorin, and bromin, if each symbol represents **one atom**, must be H_2, N_2, O_2, Cl_2, Br_2.

MOLECULAR WEIGHT OF OXYGEN

EXPERIMENT

Purpose:……………………………………………………………………………………

Apparatus: 15-cm. test-tube (hard glass) fitted with 1-hole stopper and delivery tube, drying tube, 2-liter bottle, thermometer, barometer, balance, 500-cc. graduate.

Materials: Powdered potassium chlorate, powdered manganese dioxid (heated in a porcelain dish with full flame for 5 minutes and stirred with an iron wire), anhydrous calcium chlorid.

Mix about 6 g. of manganese dioxid and 7 g. of potassium chlorate on a paper, and slide the mixture into the hard-glass test-tube. Place a plug of glass wool or well-ignited asbestos fiber in the mouth of the tube to catch small particles carried along by the gas. Weigh the test-tube and contents together with the drying tube and the tube connecting them, as accurately as possible and record below. Arrange the apparatus as shown in Figure 58. Use water at room temperature.

Heat the test-tube, gently at first, beginning with the portion of the mixture nearest the stopper, and continuing until the bottle is nearly full of oxygen. All of the oxygen must be collected, and it is well to regulate the heat so that the bubbles never come too fast to be counted. When the bottle is nearly full of oxygen, stop heating the test-tube and remove the delivery tube from the water as soon as the last bubble has risen into the bottle. Disconnect the delivery tube from the drying tube so that the water in it may not be drawn back into the drying tube.

Lower the bottle of gas in the water pan until the levels inside and out are the same, inclining the bottle if necessary, close the mouth of the bottle tightly with a glass plate, and set it upright on the table. Find the volume of the oxygen by measuring the quantity of water needed to fill the bottle level full — use a large graduate.

Weigh the test-tube and contents together with the drying tube and the tube connecting them on the same balance and with the same set of weights as were used before, and compute the weight of oxygen given off. Note the temperature of the gas (same as that of the water), and the barometer reading (apply the temperature correction, p. 233). As the gas is moist, look up the vapor pressure for the observed temperature (p. 233) and compute the pressure of the oxygen. Calculate the volume the dry oxygen would occupy at 0° C. and 760 mm. pressure. As the weight of this oxygen is known, the weight of 22.4 L. may be calculated.

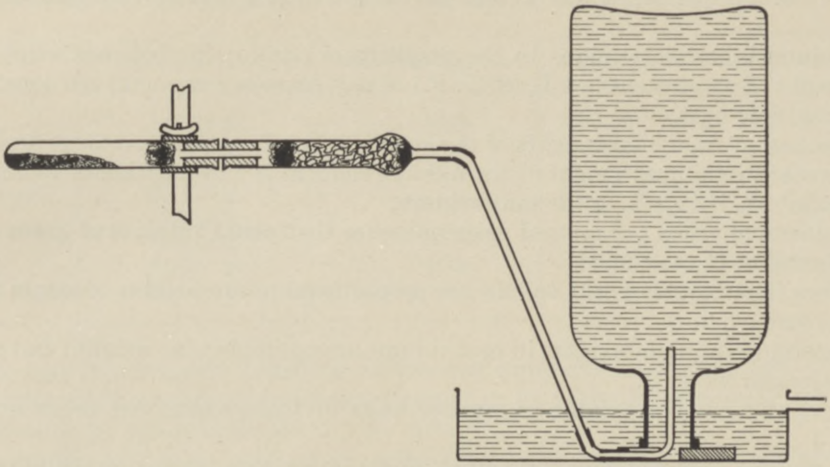


FIG. 58. Collecting oxygen from a definite weight of potassium chlorate.

RECORD OF RESULTS

Weight of test-tube, contents, and drying tube before heating........ g.

Weight of test-tube, contents, and drying tube after heating.......... g.

Weight of oxygen.. g.

Volume of oxygen.. cc.

Temperature... ° C.

Barometric pressure (corrected for temperature).................. mm.

Tension of aqueous vapor.................................. mm.

Pressure of the oxygen.................................... mm.

Volume of oxygen reduced to S. T. P............................ cc.

Weight of 22.4 L. of oxygen at S. T. P........................... g.

Error (accepted value is 32.00 g.)............................ g.

% error (error × 100 ÷ 32.00)................................ %

ATOMIC WEIGHTS

In the preceding discussion on the Theory of Atoms it was pointed out that the elements must consist of very small particles called atoms. We must further assume that the atoms of the same element have the same weight, else the ratio in which two elements combine would not be a constant one. Atoms of different elements must have different weights, or there would not be such a variety of combining ratios as exist among the various elements. Chemical action must consist of combination or separation of the atoms, or a rearranging of them into new groups or molecules. The preceding assumption means that the atom is the unit of matter in chemical action, that there is never less than one atom of each element, or a whole number of atoms, if more than one, entering into a chemical action. The evidence for this last assumption is the fact that the molecular weight of a compound never contains less than the atomic weight of each element present, and that when more than one atomic weight of an element is present in the molecular weight, it is an exact whole multiple of the atomic weight.

The action of magnesium on hydrochloric acid must therefore consist of **one or more** atoms of magnesium replacing **one or more** atoms of hydrogen. The most careful determinations of the equivalent weights of magnesium give 12.16 g. as the weight needed to replace 1.008 g. of hydrogen. **If** one atom of magnesium replaces one atom of hydrogen, then these numbers must be the relative weights of the atoms.

But it is possible that the atoms of magnesium do not replace those of hydrogen one for one, but in some other ratio. If one magnesium atom can replace two (or three, or four) atoms of hydrogen, then only half (or one-third, or one-fourth) of the atomic weight of magnesium must be equivalent to the whole atomic weight of hydrogen (1.008). Hence the atomic weight of magnesium must be numerically two (or three, or four) times its equivalent weight, 12.16 g. Again, it may be that it takes two (or three, or four) atoms of magnesium to replace one atom of hydrogen, and in that case, by a similar argument, the atomic weight of magnesium will be but half (or one-third, or one-fourth) of its equivalent weight.

This would indeed be a very perplexing state of affairs, full of distress and trouble for the chemist, were it not for the fact that ways have been found to determine what the relationship between the atomic and equivalent weights of an element really is. Dulong and Petit discovered that the products of the atomic weights of the solid elements by their specific heats[1] were in almost all cases numbers lying between 6 and 7, the average being 6.4. Hence dividing 6.4 by the specific heat of an element in the solid state gives an approximate value of its atomic weight.

Element	Atomic Weight	Specific Heat	Product	Element	Atomic Weight	Specific Heat	Product
Lithium	7	0.94	6.6	Iron	56	0.112	6.3
Sodium	23	0.29	6.7	Zinc	65.4	0.093	6.1
Silicon	28	0.16	4.5	Silver	107.9	0.0570	6.15
Phosphorus (white)	31	0.19	5.9	Tin	118.5	0.0562	6.66
Calcium	40	0.170	6.8	Gold	197	0.032	6.3

When this product, 6.4, is divided by the specific heat of magnesium, 0.245, a number will be obtained, 26.12, which is **approximately** the atomic weight of magnesium. Since this number does not approximate the equivalent weight, 12.16, the atomic weight will be a multiple of it. The nearest multiple is 24.32. Therefore, this is the accepted atomic weight of magnesium.

This method has been found of great value in the case of those elements whose compounds are not volatile, so that it is impossible to determine their atomic weights according to the method given in the discussion on the Theory of Atoms.

---

[1] A continuous uniform source of heat takes nine times as long to warm a definite weight of water one degree centigrade as it does to warm the same weight of iron one degree. Hence the iron requires only one-ninth as much heat as the water for the same rise of temperature. This ratio, one-ninth, usually expressed decimally as 0.112, is called the specific heat of iron.

ATOMIC WEIGHT OF COPPER

EXPERIMENT

Weigh a clean, dry, hard-glass test-tube carefully, and record in the table below. Fold a piece of paper 20 × 2 cm. lengthwise and put in the V-shaped paper about as much cupric oxid as can be put on a ten-cent piece. Holding the tube horizontally, insert the paper and by tipping up the tube drop **all** the oxid at the bottom of the tube. Remove the paper and weigh again (?). Fasten the tube in a horizontal position by a clamp near the mouth. Insert a 25-cm. piece of glass tubing in a rubber tube attached to the gas jet and fasten it in such a position that the end will be 2–3 cm. from the bottom of the tube (see Fig. 59).

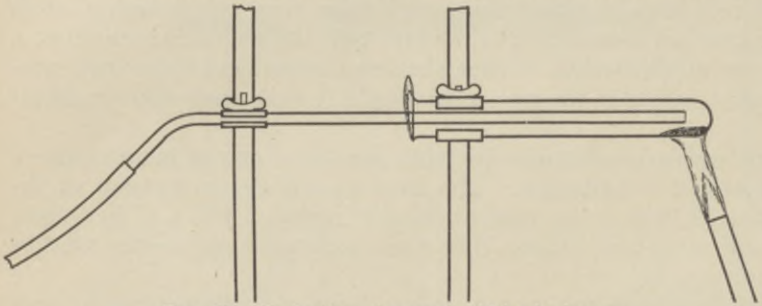


Fig. 59. Reduction of copper oxid by hydrogen.

Hydrogen from a generator may be used in place of the illuminating gas, but be sure that all of the air has been swept out of the generator before applying the light. Illuminating gas, especially water gas, contains from 30 to 40 per cent of hydrogen by volume, and in this experiment, for the sake of simplicity, we shall regard the chemical action as being due to this component.

Light the gas as it issues from the test-tube and turn the gas down until the flame at the mouth of the test-tube is 1 cm. high. Heat the oxid slowly to red heat. Continue for 5 minutes, withdraw the tubing, turn the test-tube in the clamp half around (use a cloth), and tap the tube to cause the oxid to fall down to the other side. Reinsert the tube and continue heating. When the oxid is entirely reduced to red copper, cool the tube and weigh (?). The loss of weight is the weight of the oxygen removed from the oxid.

Calculate the equivalent weight of copper — that weight which combines with one equivalent of oxygen, and is thus equivalent to one gram of hydrogen.

$$\frac{\text{weight of oxygen}}{\text{weight of copper}} = \frac{8 \text{ (equivalent weight of oxygen)}}{x \text{ (equivalent weight of copper)}}$$

Then determine the atomic weight of copper by means of the specific heat given below, applying Dulong and Petit's Law.

RECORD OF RESULTS

Weight of the test-tube.................................... g.
Weight of the test-tube and copper oxid.................... g.
Weight of the test-tube and copper......................... g.
Weight of the oxygen....................................... g.
Weight of the copper....................................... g.
Equivalent weight of the copper............................ g.
Specific heat of copper....................................0.0963
Constant product (6.4) divided by specific heat ....
Factor by which the equivalent weight must be multiplied to get the atomic weight ....
Atomic weight of copper....................................
Accepted value of atomic weight of copper.....................63.57
Error..
Per cent of error (error × 100 ÷ accepted value).................

ACTIVITY SERIES

In the oxygen experiment, the elements used showed considerable variation in the ease and vigor with which they united with the oxygen. The elements which unite most actively with oxygen also enter readily into other chemical actions, so a list of the elements arranged in the order of their activity towards oxygen would be a great help in studying the properties of the elements. As the speed of a chemical action varies greatly with the temperature at which it occurs, all the actions to be compared should take place at the same temperature. Since some elements unite with oxygen at comparatively low temperatures, while others must be heated to rather high temperatures to start the action, it is quite difficult to place the elements in their proper order. As the metals displace hydrogen from acids at ordinary temperatures, this action may be used as a convenient means of arranging the metals in an activity series according to their action on the same acid. The acids may also be placed in a list according to their activity by observing their action on the same metal.

The value of such activity lists lies in the fact that many other facts can be grouped and arranged around the list in such a way as to make them easily understood and remembered. The least active metals form very few compounds and are found in Nature in the uncombined or free state. The metals in the middle of the list form a larger number of compounds which are also more difficult to decompose, and they are generally found in Nature in compounds from which they can be separated without much difficulty. The most active metals form a still greater variety of compounds, which are more stable than the corresponding compounds of the metals below them in the series. They never occur free in Nature, and are separated from their ores only with considerable expenditure of energy. The least active metals, occurring free, were known to the ancients, while the combined metals were not known until means of separating them from their ores were discovered. The effect of moisture and oxygen upon the metals, commonly known as rusting, varies according to their activities, the least active being entirely unaffected. Any metal may be used as the positive plate of a Voltaic cell with carbon as the negative, but a cell made with an active metal has a higher voltage than one made with an inactive metal.

In making an activity list it is very important that all the conditions be kept constant throughout the experiment, or false conclusions may be drawn from the results. As an action with a solid can take place only at its surface, the various metals whose activities are to be compared should be in such form as to present equal surfaces to the acid. Iron in the form of filings displaces hydrogen from acids much more actively than when used in larger pieces. The temperatures should also be the same in all trials, for it is easy to make a metal which is really less active displace hydrogen at a high temperature with greater vigor than that shown by a more active metal at a lower temperature. The concentration of the acid used should be the same with all the metals tried, for the amount of water present has a great influence upon the speed with which a metal can displace hydrogen from the solution. This control of the conditions surrounding a phenomenon is the main characteristic of a scientific experiment, and the more exactly the conditions are controlled, the more valuable the experiment, and the greater the reliance to be placed upon the results.

References: Smith, p. 42; M. & L., pp. 352–354; B. & W., pp. 327–331; Newell, p. 316; H. & S., p. 162.

EXPERIMENT

Purpose:..

Apparatus: Sodium capsule or spoon, 5 glass dishes, 500-cc. flask, iron pipe fitted with stoppers and tubes as shown in Figure 61.

Materials: Iron wire or nails, mossy zinc, mossy tin, copper turnings, aluminum wire, magnesium ribbon, potassium, sodium, calcium.

I. **Activity Series of the Acids.** Measure 15 cc. of pure conc. hydrochloric acid into a beaker, add an equal volume of water, and stir thoroughly. Into separate test-tubes pour 5 cc. portions of each of the following reagents: dilute hydrochloric acid (as prepared above), dilute sulfuric acid, dilute acetic acid, and water. Add a few pieces of mossy zinc to each tube and note the vigor of the actions (?). Test any gases given off with a lighted splinter (?). If there is no bubbling, warm gently and test again. Remember that the gas produced during the heating may be only steam or acid vapors, and not hydrogen; bubbling which continues after the flame has been removed indicates the formation of hydrogen. If the gas comes too slowly to test with the flame, hold an empty test-tube mouth downward over the other for a minute or two and then test the upper tube for hydrogen (?).

Observations: ....

....

....

Questions: 1. What is the order of the activities of the acids used (regarding water as an acid)? ....

2. Complete the following equations excepting in the case where no chemical action occurs:

Zinc + hydrogen chlorid ⟶ ______________ + ______________.
Zinc + hydrogen sulfate ⟶ ______________ + ______________.
Zinc + hydrogen acetate ⟶ ______________ + ______________.
Zinc + hydrogen hydroxid ⟶ ______________ + ______________.

3. If any of the above actions is too slow to be observed under the conditions prevailing, in what two ways might the action be hastened? ....

....

II. **Activity Series of the Metals, Reacting with Acids.** Place in separate test-tubes a few small pieces of each of the following: iron, zinc, tin, copper, aluminum, and magnesium. Pour 5 cc. of the dilute hydrochloric acid prepared above into each tube and **quickly** apply a lighted splinter to the mouth of each tube, first to those showing the greatest activity. Warm and proceed as in Part I in any case where action is slow. Write the names of the metals in the first column of the table below in the order of their activity, most active first, and then fill out the rest of the table.

METALS	ACTION OF COLD ACID?	AFTER HEATING?	RESULT OF FLAME TEST?	WHAT IS THE GAS?

4. Where do the bubbles appear, on the glass or the metal? Why?

....

5. What is the source of the hydrogen, the acid or metal? How do you know?

....

6. Complete the following equations, **except for the case where no action occurred:**

Iron + hydrogen chlorid → ________________ + ________________.
Zinc + hydrogen chlorid → ________________ + ________________.
Tin + hydrogen chlorid → ________________ + ________________.
Copper + hydrogen chlorid → ________________ + ________________.
Aluminum + hydrogen chlorid → ________________ + ________________.
Magnesium + hydrogen chlorid → ________________ + ________________.

III. **Activity Series of Metals, Reacting with Water.** Instructor. (*a*) Put small pieces of potassium, iron, magnesium, calcium, and sodium in water in separate vessels (glass crystallizing dishes) and record the results (?).

....

....

....

(*b*) Fill a test-tube with water and invert it in a trough of water. Put a small piece of sodium in a sodium capsule or spoon and hold the spoon below the test-tube (?). Use a second piece of sodium — dry the spoon — if needed to fill the tube. Remove the tube and apply a lighted splinter to the mouth. What is the gas?

....

....

(*c*) Collect and test the gas formed by the interaction of calcium and water (?).

....

(*d*) Pour about 150 cc. of water into a 500-cc. Erlenmeyer flask. Place it on a gauze-covered ring and heat to boiling. When steam is issuing freely from the flask, adjust the size of the flame so that the water boils quietly. By means of the tongs, lower an ignited strip of magnesium ribbon into the steam in the flask (Fig. 60) (?). Interpose an opaque screen so as to just hide the entire flask from the class and repeat the experiment, noting the action at the mouth of the flask (?). Explain.

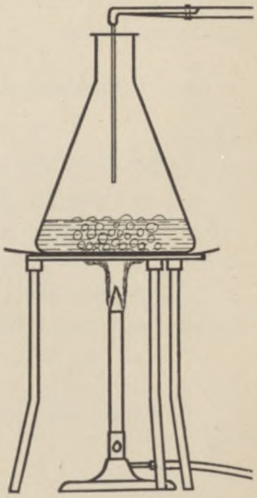


FIG. 60. Interaction of steam and magnesium.

....

....

(*e*) Fill the middle portion (about 20 cm.) of a $\frac{3}{4}$ in. iron pipe 75 cm. long with small brads and then fasten the tube and arrange the rest of the apparatus as shown in Figure 61. To avoid charring the corks, keep the ends of the tube cool by means of wet cloths or strips of filter paper. Boil the water in the flask gently and heat the middle of the pipe as hot as possible. Collect a tubeful of the gas by water displacement and test it with a lighted splinter (?).

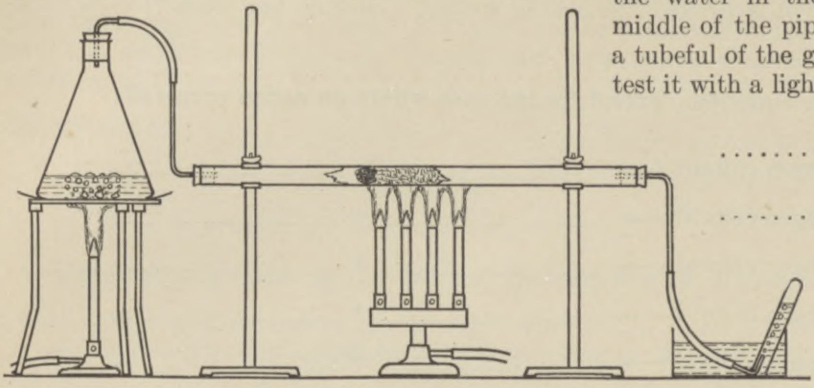


FIG. 61. Interaction of steam and iron.

....

....

....

....

....

7. What is the order of the activity of the metals used in Part III, putting the most active first?

....

8. What is the order of the activity of all the metals used in the entire experiment?

....

9. Would you expect to find potassium, sodium, calcium, and magnesium in a free state in Nature? Why?

....

10. The hydroxids of all but the most active metals are decomposed at high temperatures, leaving the oxid of the metal. Therefore actions which give the hydroxids of the metals at low temperatures, give the oxid at high temperatures. There are examples of that here. Complete the equations below:

Potassium + hydrogen hydroxid ⟶ ________________ + ________________.

Iron + hydrogen hydroxid ⟶ ________________ + ________________.

Magnesium + hydrogen hydroxid ⟶ ________________ + ________________.

Calcium + hydrogen hydroxid ⟶ ________________ + ________________.

Sodium + hydrogen hydroxid ⟶ ________________ + ________________.

11. Aluminum is third in the order of abundance of elements occurring in the earth's crust, but was not discovered until 1827. Why not?

....

12. Why was copper discovered so early?

....

13. Where would gold and silver stand in your activity list? Remember the effect of air upon them.

....

ANHYDRIDS

Of the eighty-odd elements known at the present time, only forty are common enough to require attention in an elementary course, and their names and symbols will be found in heavy type in the list on page 236. The symbols consist of one or two letters taken from the names (sometimes from the Latin or German names) of the elements, and each symbol represents one atom of its element. Just as the atoms of the different elements are Nature's alphabet by means of which every substance is spelled out, so the symbols are put together to spell out formulas representing the molecules of the compounds.

The names of the elements and their derivations are very interesting, as they are in most cases significant of some property of the element or of some fact in its history.

The majority of the elements possess similar physical properties, such as a peculiar metallic luster, good conductivity of heat and electricity, malleability, and ductility. They are called metals, and all others are called non-metals. The metals discovered in modern times have names ending in " ium " to signify metallic properties. Exceptions to this are helium, selenium, and tellurium, non-metals. All those ending in " on " are names of non-metals, except iron.

The distinction between the metals and non-metals is not based so much upon their physical properties, however, as upon their chemical relations — through their oxids — to the bases and acids. Bases and acids are among the most important of all classes of substances. Many acids occur in Nature, entering into important chemical changes, and there are very few lines of industry entirely independent of acid substances. Bases are also quite common and familiar; examples being slaked lime, used in mortar and plaster; caustic soda and caustic potash, used in soap-making; ammonia water, used in cleansing.

The word " anhydrid " literally means " without water," but in chemistry its use is restricted to such substances as may be obtained from acids or bases by the removal of water, or that will unite with water to form acids or bases.

Study Topics: *Derivation of names of common elements; origin of symbols of antimony, copper, gold, iron, lead, mercury, silver, sodium, potassium, tin, and tungsten.*

References: B. & W., pp. 210–212; M. & L., pp. 141–143; Newell, p. 41; Brownlee, p. 207; Smith, p. 161; McP. & H., p. 134; H. & S., p. 211.

EXPERIMENT

Purpose:...

Materials: Asbestos paper, sulfur, red phosphorus, magnesium ribbon, mossy zinc, litmus paper.

I. **Tests for Acids and Bases.** (*a*) Pour a few drops of a dilute solution of an acid into a watch glass and dip short strips of moistened litmus paper of both colors into it. Record any change of color — not of shade — in the table below. Rinse the watch glass, throw the used litmus paper into the waste jars (not in the sink), and repeat the test with two other dilute acids.

(*b*) Repeat the above operations with dilute solutions of the following bases: sodium hydroxid, ammonium hydroxid, calcium hydroxid, and record the results below.

Acids	Effect on Litmus		Bases	Effect on Litmus	
	Red	Blue		Red	Blue

Questions: 1. What is the test for acids?

2. What is the test for bases?

3. Why should both colors of litmus paper be used in making tests?

....

II. **Non-metals and Acid Anhydrids.** (*a*) Sulfur. Line the combustion spoon with asbestos paper and place a small piece of sulfur in it. Ignite the sulfur and lower the spoon into a jar of air, covering the jar as far as possible with a glass plate. When the action ceases, remove the spoon, pour in about 5 cc. of water, and shake vigorously. Test the liquid with both colors of litmus paper and record results in the table below.

(*b*) Phosphorus. Reline the spoon and repeat the above operations, using a very small amount of red phosphorus (size of a grain of rice; remember to burn off all the phosphorus).

Questions: 4. What kind of substance (not state of matter) is formed by the combustion of a non-metallic element?

....

5. What kind of substance is formed by the action of water upon this product of combustion?

....

6. Define a non-metal according to the facts just learned.

....

III. **Metals and Basic Anhydrids.** (*a*) Magnesium. Burn a 4-cm. piece of magnesium ribbon — hold the ribbon with the tongs — and catch the product of combustion in the evaporating dish. Add 5 cc. of water and small pieces of both colors of litmus paper, and stir occasionally until one piece changes color. Record the result below.

(*b*) Zinc. Heat about 1 cc. of mossy zinc in an old spoon, stirring the molten zinc with an iron wire until all the zinc has been consumed. Do not get your face directly over the spoon. When the product of the action has cooled, add to it 5 cc. of water and pieces of litmus paper, stir well, and wait until one piece of litmus paper changes color. Record results below.

Element	Name the Product of Combustion	Kind of Product Formed by Adding Water?
Sulfur....................		
Phosphorus...............		
Magnesium...............		
Zinc......................		

Questions: 7. What kind of substance (not state of matter) is formed by the combustion of the metals?

....

8. What kind of substance is formed when the product of combustion of a metal combines with water?

....

9. Define a metal according to the facts just learned.

....

NAMES OF ACIDS AND BASES

Exercise

Radicals

Acids and bases readily react and produce a third class of substances called **salts**. The members of these three classes are so numerous and useful, and have certain general properties so different from those of all other substances, that they are the chief consideration in the study of elementary chemistry. In most of the chemical actions in which they take part, especially in solution, these substances act as if they were composed of two or more parts, which are called **radicals**. Many facts, the most important of which are presented in a later discussion (see p. 75), indicate that these radicals are more or less completely separated from each other when the substance is dissolved, and lead a fairly independent existence. For this reason the **separated radicals** of a dissolved acid, base, or salt are called **ions**, or wanderers, and the substances from which they are formed are called **ionogens**, or ion producers. There are **simple radicals** consisting of one atom, and **compound radicals** consisting of two or more like or different atoms. The same terms are applied to the separated radicals in solution, the ions.

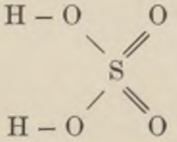
Most chemical actions between acids, bases, and salts consist in a simple exchange of the radicals of which they are composed, without any change in the radicals themselves. As a result each radical is found in quite a large number of substances. As many of the properties of these substances depend upon the radicals they contain, all having the same radical are found to have certain common properties. An excellent example of this is found in the acids, which always contain the hydrogen radical. The action on metals, on litmus, and the sour taste are all due to the hydrogen radical in ionic form.

The bases and salts are named by putting together the names of the radicals they contain: *e.g.* salt contains the sodium (Na) and chlorid (Cl) radicals, and is called sodium chlorid; caustic potash contains the potassium (K) and hydroxyl radicals (OH), and is called potassium hydroxid. Acids may also be named in this same manner, as hydrogen nitrate, HNO_3 (nitric acid), composed of the hydrogen (H) and nitrate (NO_3) radicals; but they are generally named according to the system explained below.

Acids

I. **Ternary Acids**. These acids are composed of three different elements, as the name indicates. Besides the hydrogen radical, which is found in all acids and is responsible for the acid properties, all ternary acids have compound radicals which are composed of some non-metallic element and varying amounts of oxygen. They are sometimes called oxy-acids for this reason. In all of these acids the oxygen atoms are attached directly to the third element, which is the distinctive one of each acid and furnishes the name. The hydrogen radicals are in all cases attached to some of the oxygen atoms. Hydrogen atoms directly united to the distinctive element are not replaceable by metals, hence they are not hydrogen radicals.

Sulfuric acid



(*a*) **Ternary Acids Ending in "ic."** As a rule the suffix "ic" is substituted for the final syllable of the distinctive element of an acid. The first two of the following list add the suffix to the whole name. A vowel immediately preceding the suffix is usually omitted.

Using regular paper with an appropriate heading write these formulas neatly in a column and then write their names after them: H_2SO_4, H_2CO_3, HNO_3, H_3AsO_4, H_3BO_3, $HClO_3$, $HBrO_3$, H_3PO_4, HIO_3, H_4SiO_4, H_2CrO_4. Study this list until you are thoroughly familiar with it.

Monobasic Acids. Four of the above acids have only **one** hydrogen radical in the molecule, and are therefore called **monobasic**. Write the formulas and names of these in a separate list.

Dibasic Acids. Three of the above acids are dibasic. Define the term and write the formulas and names in a separate list.

Tribasic Acids. Three of the above acids are tribasic. Define the term and write the formulas and names in a separate list.

Tetrabasic Acids. The remaining acid is tetrabasic. State why and write its formula and name.

(*b*) **Ternary Acids with Prefix " per."** The prefix " per," meaning " more," is used with the name having the suffix " ic " for acids having a larger proportion of oxygen (usually one more atom) than the acid whose name ends in " ic." Write the formulas and names of such acids of Cl and I.

(*c*) **Ternary Acids with the Ending " ous."** The suffix " ous " is used instead of " ic " for acids having a smaller proportion of oxygen (usually one atom less) than the acid ending in " ic." Write the formulas and names of such acids of S, N, As, Cl, P.

(*d*) **Ternary Acids with the Prefix " hypo."** The prefix " hypo," meaning " less," is used with the name having the suffix " ous " for acids having a smaller proportion of oxygen (usually one atom less) than is found in the acid with the ending " ous." Write the formulas and names of such acids of N, Cl, Br, and P. Double the subscripts of the first.

II. **Binary Acids.** The binary acids are composed of two elements; they do not contain oxygen. Their names have the prefix " hydro " and the suffix " ic." Copy the formulas and write the names of the following binary acids: H_2F_2, HCl, HBr, HI, and H_2S. Note that two of these acids are dibasic.

BASES

All bases have oxygen and hydrogen directly combined (OH). This (OH) group is called the " hydroxyl " radical, and in ionic form is responsible for all the common properties of bases. The other radicals of bases are usually simple and are composed of metallic elements, whose symbols are written first in the formulas of bases. The commonest exception to this is ammonium hydroxid, which contains the ammonium radical (NH_4). The bases are named by adding the word " hydroxid" to the name of the characteristic metal in the base. A hydroxid, therefore, usually means a base.

Water is a compound of the hydrogen and hydroxyl radicals (HOH). Having the hydrogen radical (H) it might be an acid, and having the hydroxyl radical (OH) it might be a base; but the properties of these radicals are quite opposite, so their combination makes a substance that has the properties of neither. Bases may be regarded as derivatives of water whose hydrogen radical has been replaced by a metal. It may be recalled that bases were actually made after this manner in the experiment on Activities.

Monacid Bases. Monacid bases have one hydroxyl radical in the molecule. Write the formulas and names of the monacid bases formed by the following metals: Li, Na, K, Rb, Cs, Tl.

Diacid Bases. Define the term " diacid base." Write the formulas and names of the diacid bases formed by the following metals: Mg, Ca, Sr, Ba, Zn, Cd, Cu.

Triacid Bases. Define the term " triacid base." Write the formulas and names of the triacid bases formed by the following metals: Al, Fe, Cr, Co.

Tetracid Bases. Define the term " tetracid base." Write the formulas and names of the tetracid bases of Sn, Ti, and Zr.

SOLUBILITY

It is a common fact of experience that, when certain solids like salt and sugar are mixed with water, they disappear from view. It is possible to pulverize a solid so finely that it is difficult for one to see a separate particle; but it is impossible, even with the aid of a microscope, to see a separate dissolved particle. The division of a substance in solution, then, must be very great. There is evidence to show that not only are the molecules separated and wandering about in solution, but in a water solution of acids, bases, and salts some of the molecules (not all) are broken into parts which have properties different from the molecules themselves. These parts are called "ions." They seem to be much more active chemically than the molecules, and the number of them present determine how rapidly a chemical change takes place in a solution. A very large number of the chemical changes taking place in Nature and the arts and industries are brought about more readily between the substances in solution, because the finer state of division makes a greater surface at which chemical action may take place, and some of the molecules are changed into the more active ions. If no chemical change (except the change of molecules into ions) occurs when a substance is dissolved in water, the water may be expelled by heating and evaporating and the original solid recovered.

Some substances are said to be insoluble in water. This is not true in the absolute sense, for everything is soluble to some small degree. But where the solubility is so small as to be a negligible factor, that substance is considered insoluble. The real meaning of the ordinary term " solubility " is **degree of solubility**, which is usually defined as " the number of grams that can be dissolved in 100 cc. of water at a given temperature." (See table of solubilities, p. 241.) For convenience the dissolved substance is called the **solute**, and the substance in which the solution takes place the **solvent**.

Solubility is dependent on several factors which should be known if many of the common operations of daily life are to be performed in an efficient manner. For example, the stains on clothing may be made more permanent by the application of a substance that is not a solvent for that stain. Too little attention is given this subject in laundries where all kinds of stains are treated in the same general way. To accomplish the cleaning as quickly as possible powerful chemical agents are used which attack and rot the fibers of the cloth and cause it to wear out much sooner than it would if proper methods of laundering were used. A knowledge of solubilities is indispensable in some occupations. Mixtures are separated and substances are purified mainly by taking advantage of the different solubilities of different substances under conditions that are varied to suit the occasion. The food that is necessary to build up plant and animal structures must be brought into solution, or a condition similar to it, before the plant or animal can assimilate it or actually use it.

Study Topics: *The Effect of Pressure on Solubility. The Similarity of a Substance in the Gaseous State and in Solution. The Effect of a Solute on the Freezing and Boiling Points of the Solvent. The Heat of the Formation of a Solution. Solubility, and the Assimilation of Foods. How to Remove Stains.*

References: Smith, pp. 84–92; H. & S., pp. 67–73; M. & L., pp. 63, 64; McP. & H., pp. 96–100; B. & W., pp. 92, 93; Newell, pp. 42–44; Brownlee, pp. 52–54.

EXPERIMENT

Purpose: Write purpose on separate sheet with heading.

Materials: Copper sulfate (crude), potassium permanganate, rosin, alcohol (grain and wood), carbon tetrachlorid, gasoline, benzene, carbonated water.

I. **Solids.** (*a*) Circulation of Solvent. Select two 5-mm. crystals of copper sulfate. Drop one into a test-tube nearly full of water. Suspend the other by a looped strip of paper just below the surface of water in another nearly filled tube (Fig. 62). Stand the tubes in the rack, noting the time by the clock, and examine the tubes occasionally while continuing work so as to compare the rates of solution (?). In order to see the circulation effect more plainly suspend a small crystal of potassium permanganate as above (?).

FIG. 62. Solid suspended in water.

Questions: 1. Why does the solution move downward in the solvent?

2. How does this movement hasten the formation of a solution?

3. How can this same effect be brought about in other ways?

(*b*) Extent of Surface. Since a solid can dissolve only at the surface, any change of the extent of surface affects the rate of solution.

Select two 1-cm. crystals of copper sulfate. Put one in a 15-cm. tube, pulverize the other crystal and put the powder in another tube. Add 20 cc. of water to each tube and shake both gently until the powdered copper sulfate has dissolved. Estimate the time necessary to dissolve the other crystal, and compute the saving of time by pulverizing a substance to be dissolved (?).

4. Explain how pulverizing a solid increases the surface.

5. State two ways of hastening solution.

(*c*) Effect of Different Solvents. Put 1 cc. of powdered rosin in each of two dry tubes. Pour about 5 cc. of wood alcohol into one tube and the same volume of water into the other. Shake each tube and compare results (?). Pour the alcoholic solution into a beaker nearly full of water (?).

6. State the solubility of rosin in water and alcohol.

7. Is the rosin in the beaker in solution? Look for separate particles.

8. Why was the rosin " precipitated from solution " by pouring the solution into water?

9. Is rosin lighter or heavier than water? Drop a small lump into water to find out.

10. Why does the rosin remain " in suspension "?

II. **Liquids.** (*a*) Put exactly 10 cc. of water in a test-tube. Measure out exactly 5 cc. of alcohol and add it to the water in several portions, shaking after adding each portion (?). Measure final volume of mixture, then pour it into a bottle labeled " dilute alcohol " to save for future use.

State final volume of mixture and explain why it is not 15 cc. (?).

(*b*) To about 10 cc. of water add 1 cc., no more, of carbon tetrachlorid and shake (?). Describe result and state solubility of carbon tetrachlorid in water (?).

(*c*) Repeat procedure for Part (*b*), using gasoline or benzol (?).

III. **Gases.** (*a*) Effect of Pressure. Carbonated water is prepared by forcing carbon dioxid into water under pressure. Uncork a bottle of carbonated water, and pour out about 20 cc. into the smallest beaker; or press the lever of a " seltzer " bottle, allowing this quantity to flow out into the beaker (?). Bubbles of gas coming from a liquid not at the boiling point is called **effervescence.**

11. How is the solubility of a gas affected by the release of pressure? State the evidence on which answer is based (?).

(*b*) Effect of Temperature. When the water has ceased to effervesce, heat it slowly over a wire gauze until effervescence begins again. Test the temperature of water with the finger to learn if the water is hot enough to be boiling (?).

Heat about 100 cc. of cold water from the tap in a beaker over a wire gauze. Do not boil. Look for small bubbles on the inside of beaker (?).

12. How does a rise of temperature affect the solubility of a gas?

13. Explain the formation of bubbles of air from heated tap water.

14. Distinguish between boiling and effervescence.

HYDRATES

The crystal forms of some substances, like washing soda (sodium carbonate) and many others, if exposed to the open air crumble to a powdery mass in a few days. The crystal form disappears. By careful observation of this action it has been found that, with the loss of crystal form, water vapor is given off. This shows that water is essential for the crystal forms of **some** substances. The water in such crystals is able to maintain a pressure (aqueous tension) that is greater than the pressure of the water vapor in the atmosphere. Consequently, the water flows out of the crystal as vapor into the atmosphere. This occurrence is called **efflorescence**.

Since crystals are formed from solutions, the water was built into the crystal at the time of formation, and since the loss of water causes such an alteration of the characteristic properties, the change must be a chemical one. When the water is present, the substance is distinctly different from the substance without it, so it is called a **hydrate**. However, the water is not so firmly combined as the remaining constituents, for moderate heat (often under 100° C.) decomposes most hydrates by the separation of water. For this reason the water is usually expressed separately in the formula of the substance, as $Na_2SO_4 \cdot 10\ H_2O$. One can always tell by the formula in reference books whether a substance has a hydrated form or not.

Some hydrates, having a low aqueous tension in the crystal form, give off the water vapor only when heated. When all removed, the substance is said to be dehydrated or in anhydrous form. If such substances are exposed to the air, they automatically regain water from the air, in some cases in sufficient quantity to make a solution of the substance. Such action is called **deliquescence**. This property makes such substances useful as drying agents. Calcium chlorid was so used to dry hydrogen gas before it was burned.

Some crystals, like those of sodium chlorid and lead nitrate, contain water, but it is not chemically combined. When heated, the crystals burst with considerable noise. This indicates that the water was probably inclosed within a crystal mass as it was formed. On heating, the pressure of the steam bursts the crystal. The peculiar crackling noise is called **decrepitation**.

The aqueous tension in hydrated cobalt chlorid is about equal to the average water vapor pressure of the atmosphere. This substance loses or takes up water, depending on whether the air is dry or moist. Cloth or paper dipped in a solution of it and dried will be pink in moist air and blue in dry air. This effect is used to indicate the relative humidity of the air and hence the probability of rain. If the solution of cobalt chlorid is used as an ink, the writing is invisible (almost) until it has been heated and dried. It is called "sympathetic ink."

Alum crystals contain much water. For this reason they melt at rather low temperature, 90° C. If alum crystals are dissolved in a few cubic centimeters of water above the temperature 90° C., an astonishing fact is found. There seems to be no limit to the solubility. The explanation is that above 90° C., the crystals become a liquid which may be mixed with water in any proportions. Sodium sulfate, $Na_2SO_4 \cdot 10\ H_2O$, breaks up into the anhydrous form at a little above 30° C. Hence, its solubility curve at this temperature shows an abrupt change of direction as the anhydrous compound is formed.

Some common hydrates are:

$BaCl_2 \cdot 2\ H_2O$
$CaCl_2 \cdot 6\ H_2O$
$CaSO_4 \cdot 2\ H_2O$
$MgSO_4 \cdot 7\ H_2O$
$CoCl_2 \cdot 6\ H_2O$
$FeSO_4 \cdot 7\ H_2O$
$Na_2B_4O_7 \cdot 10\ H_2O$
$Na_2CO_3 \cdot 10\ H_2O$
$Na_2SO_4 \cdot 10\ H_2O$
$Na_2S_2O_3 \cdot 5\ H_2O$
$ZnSO_4 \cdot 7\ H_2O$
$K_2Al_2(SO_4)_4 \cdot 24\ H_2O$

Study Topics: *The Six Forms of Crystals. Methods of Drying Gases. The Desiccator and Its Use.*
References: M. & L., pp. 68, 69; Smith, pp. 54, 55; H. & S., pp. 73–77; Newell, p. 47; Brownlee, pp. 57–59; B. & W., pp. 185, 194; McP. & H., p. 271.

EXPERIMENT

Purpose: Write the purpose on a separate sheet with heading.

Materials: Crystals of sodium sulfate, ferrous sulfate, sodium and potassium hydroxids, copper sulfate, cobalt chlorid, alum, potassium sulfate, sodium chlorid.

I. **Efflorescence**. Select a clear crystal each of sodium sulfate and ferrous sulfate. Place the crystals on a labeled paper and leave them exposed to the open air until the end of the laboratory period, or the next period if no change can be noticed. After a decided change occurs, examine, describe change of appearance, and explain it (?).

II. **Deliquescence**. Obtain a 1-cm. piece each of sodium and potassium hydroxids. Do not handle pieces with fingers. Put each on a separate watch glass until the end of the period. Examine, describe, and explain the change of appearance (?).

III. **Hydrates as Chemical Compounds**. (*a*) In a clean, dry 15-cm. test-tube place a 5-mm. crystal of cobalt chlorid. Hold the test-tube nearly horizontal with the mouth slightly lower than the closed end, and heat the crystal slowly (Fig. 5 *a*) in a small flame. Note if there is a liquid condensing in the cool part of the tube (?). Describe change in appearance of the substance (?). Lay the tube aside to cool. When cool, put a little water on the cobalt chlorid in the tube and note the effect on the color (?). Save the solution for Part IV.

1. If water restores the original color of the cobalt chlorid, what must have been the liquid given off when heated?

The remarkable change in color of cobalt chlorid when the hydrate is dehydrated or the operation reversed makes it suitable as a test for water. Filter paper is dipped in a solution of it and then dried to the blue color. This color changes to pink in water or water vapor.

(*b*) Heat a large crystal of sodium chlorid (?). Test any liquid appearing for water, using cobalt chlorid paper for test (?). Pulverize some other crystals of sodium chlorid very finely and heat the powder in a tube (?).

2. Why is "decrepitation" greatly reduced or entirely lacking when a powdered crystal is heated?

3. What is the proof of mechanically inclosed water?

(*c*) Obtain a 1-cm. crystal of alum. Taste it (?). Heat it in an evaporating dish over a wire gauze until its appearance changes decidedly. Cool and taste it again (?).

4. Why is the characteristic taste lacking?

(*d*) Heat a crystal each of copper sulfate, potassium sulfate, sodium borate, and potassium dichromate in separate tubes and record results in tabular form (?).

5. Which two of the four substances in Part (*d*) are hydrates and which two are not? State evidence as reasons for answer.

6. How could copper sulfate be used to show the presence of water?

7. Is the water in hydrates chemically combined? Give the evidence.

8. Are all hydrates efflorescent? Give the conditions.

IV. **Optional Experiments**. (*a*) "Sympathetic Ink." Using a **clean** pen write with the solution of cobalt chlorid, saved from III (*a*), as with ink. While slightly moist, the writing will be almost invisible. Dry the ink over the low flame of a burner (?).

(*b*) Per Cent of Water in Hydrates. Weigh a crucible or evaporating dish carefully to one centigram. Put in it approximately 10 grams of freshly pulverized copper sulfate, barium chlorid, or gypsum, and weigh again carefully to the centigram. Heat the dish slowly at first and increase the rate of heating gradually. After 20 minutes cool the dish and weigh (?). Heat the dish again for 10 minutes and weigh again (?). When the final weight shows no further loss of weight from the preceding one, the substance has been **heated to constant weight**. Compute the percentage loss of weight of the substance (?). This will be the per cent of water in the hydrated substance.

TEMPERATURE EFFECT ON SOLUBILITY

The most common and important condition affecting solubility is the temperature. A solution that contains all the dissolved substance that it can hold at any given temperature is called saturated. If such a solution is heated, it becomes unsaturated. When an unsaturated solution at high temperature is cooled sufficiently, it becomes saturated, and with further cooling the solutions of certain substances present the peculiar condition of actually holding more of the substance in solution than it will take up at that temperature. Such a solution is called supersaturated.

In a previous experiment acids were found to act more rapidly when heated. The common strong acids, sulfuric, hydrochloric, and nitric, are found to be quite neutral substances at 105° C. below zero (–105° C.). A rise of temperature increases acid properties, while a fall of temperature decreases them. Geikie says, " Water at 1000° to 2000° C. will act as a powerful acid," bringing glass and iron into solution by means of chemical action. Water is in about the same condition at ordinary temperatures as acids are at very low temperatures. All that is required to bring out the acid properties is a rise of temperature.

TEMPERATURE EFFECT ON THE SOLUBILITY OF SALTS

Substance	Grams of Salts Dissolved in 100 cc. of Water at					
	0°	20°	40°	60°	80°	100°C.
Alum	4	15	31	67	135	358
Borax	3	8	18	40	76	201
Sodium chlorid	35.7	36	36.6	37.3	38.4	39.8
Sodium nitrate	73	87	102	128	148	180
Sodium sulfate	5	19	48	45	44	42.5
Potassium nitrate	13	31	64	111	172	247
Potassium chlorid	28	35	40	46	51	57
Potassium iodid	136	150	165	179	194	208
Potassium chlorate	3.5	7.5	14.5	25.5	40	55.5

Exercise

On a sheet of metrically ruled paper mark the fourth and fourteenth vertical centimeter lines heavy with ink. Beginning at the lower left-hand corner of the space within these lines as zero mark off units of temperature in the lower margin, ten degrees for every centimeter. The first inked line represents zero temperature and the other one 100° C.

Mark the gram units vertically to the left of the first inked line, ten grams for every centimeter. Plot the values given in above table by placing dots to correspond to the given temperatures and grams. As the solubility of each substance is plotted, draw a distinctive smooth line through all the points found for that substance. Label each line (curve) in the right-hand margin. Answer the following questions in writing on the reverse side of the graph sheet and present to the instructor as a written report.

Questions: 1. Give the points of advantage of the graphs over a table of figures for illustrating the solubility of the substances.

2. What substance is most soluble at ordinary temperature (about 20°)?

3. What substance is least soluble at ordinary temperature?

4. What substance has its solubility increased the most by a rise in temperature?

5. What substance has its solubility affected the least by a rise in temperature?

6. Describe the alum curve between temperatures 90–100° C. What degree of solubility is shown? Explain this effect by the hydrated form of alum. Do likewise for sodium sulfate.

7. State the solubility of each substance at 50° C., as shown by the curves.

References: M. & L., pp. 61–63; Smith, pp. 86–90; Newell, pp. 44, 45; H. & S., pp. 68–70; B. & W., p. 92; McP. & H., pp. 96–98; Brownlee, p. 53.

TEMPERATURE EFFECT ON SOLUBILITY

EXPERIMENT

Purpose: Write the purpose on separate sheet with heading.

Apparatus: Water bath.

Materials: Calcium sulfate, calcium carbonate (powdered marble), crude copper sulfate, sodium sulfate, sodium thiosulfate.

I. **Precision Test for Solubility**. (Two pupils.) Let one pupil shake 1 cc. of powdered calcium sulfate with 10 cc. of distilled water in a test-tube for 2 or 3 minutes while another performs a similar operation with powdered calcium carbonate. Filter the contents of tube into another clean tube (Figs. 12, 13). Pour the water repeatedly through the same filter until the filtrate is clear.

1. Does any of either substance appear to have dissolved in the water?

Evaporate 5 drops of each filtrate on separate watch glasses on the water bath (Fig. 6). The quantity of residue on the watch glasses indicates the exact amount of each substance dissolved. Pupils should compare results and record both (?).

2. Why is the disappearance test for solubility useless in this case?
3. State the degree of solubility of each substance.
4. Which substance is the more soluble?
5. State the method of finding that a substance is slightly soluble.

II. **Temperature Effect**. (Two pupils.) (*a*) Solubility in Cold Water. To exactly 10 cc. of water in a test-tube add well-pulverized crude copper sulfate in 2-gram portions, shaking the tube after each added portion until it dissolves. When the limit of solubility has been reached, record the grams of substance dissolved as the solubility in cold water. Use a tabular form (?). Compute the weight that would dissolve in 100 cc. of water (?). The other pupil performs the same operations with sodium sulfate, taking 5-gram portions instead of 2. Each pupil records the work of the other.

(*b*) Solubility in Hot Water. Heat the solutions prepared above in boiling water in the water bath. Continue to add portions as above until a limit is reached at this temperature (?). Compute the solubility for 100 cc. of water as before (?). Save the solutions.

6. State the effect of a rise of temperature on solubility.

III. **Crystallization**. (Two pupils.) Cool rapidly each of the solutions prepared in II (*b*), by holding the tubes in running water from the tap (?). Note the size of the particles (?). Return the tubes to the water bath until the crystals are again dissolved. Put the tubes away in locker to cool slowly until the next laboratory period (?). Leave space for record, and continue this experiment. Examine the crystals that formed slowly and quietly, comparing the size with those formed quickly (?).

7. Why did the crystals appear as the solution was cooled?
8. What conditions favor the formation of large, well-formed crystals?

IV. **Supersaturation**. (Instructor or pupils.) Fill a 15-cm. tube with pure sodium thiosulfate crystals. Warm the tube with a small flame until the crystals are dissolved (melted). Pour gently on the liquid about 1 cc. of distilled water, then cool the tube in cold tap water. The lighter water stays on top of the solution and protects it from the dust of the air. Now drop a tiny crystal of the thiosulfate into the tube. Watch closely (?). Note the change of temperature (?). When the thiosulfate has crystallized pour off the water, put a cork in the tube, and turn it in to the instructor for use in the next class.

9. What is the source of the heat noticed as the thiosulfate crystallized? Explain with the help of the molecular theory.
10. How could the condition of a solution, saturated, unsaturated, or supersaturated, be determined?

VALENCE AND CHEMICAL EQUIVALENTS

When metallic sodium reacts with hydrogen chlorid, measurement shows that one atom (one atomic weight) of sodium replaces one atom (one atomic weight) of hydrogen.

(1) $$2\,Na + 2\,HCl \longrightarrow 2\,NaCl + H_2.$$

When zinc reacts with hydrogen chlorid, it is noticed that **one** atom of zinc replaces **two** atoms of hydrogen.

(2) $$Zn + 2\,HCl \longrightarrow ZnCl_2 + H_2.$$

One atom of aluminum replaces **three** atoms of hydrogen.

(3) $$2\,Al + 6\,HCl \longrightarrow 2\,AlCl_3 + 3\,H_2.$$

There are other elements which are able to replace four, five, six, seven, or eight atoms of hydrogen.

From these instances it may be seen that atoms of different elements do not always combine with or replace one another, atom for atom. We do not know why atoms of different elements behave differently in this manner. These are the facts as we find them.

Sodium and other elements whose atoms are chemically equivalent to one atom of hydrogen are said to be **univalent**, or to have a **valence** of one. They will replace each other or combine with each other, atom for atom. Thus sodium, hydrogen, chlorin, iodin, and potassium are all univalent and form the following compounds: $NaCl$, KCl, NaI, KI, HCl, HI.

Zinc and those elements whose atoms can replace or combine with two atoms of hydrogen have a valence of two, or are said to be **bivalent**. They will replace or combine with each other, atom for atom; as, ZnO, ZnS, CaO, CaS, MgO, MgS. One bivalent atom combines with or replaces two univalent atoms ($1 \times 2 = 2 \times 1$): $ZnCl_2$, ZnI_2, $CaCl_2$, MgI_2, H_2O, H_2S, Na_2O, Na_2S, K_2O, K_2S.

Aluminum and other elements whose atoms can replace or combine with three atoms of hydrogen are said to be **trivalent**, or to have a valence of three. One trivalent atom will combine with or replace one trivalent atom ($3 = 3$); two trivalent atoms will combine with or replace three bivalent atoms ($2 \times 3 = 3 \times 2$); one trivalent atom will combine with or replace three univalent atoms ($3 = 3 \times 1$). Examples: AlN, BN (aluminum, boron, and nitrogen are trivalent); Al_2O_3, B_2O_3, Al_2S_3, $AlCl_3$, NH_3.

Similarly, an element may be quadrivalent, quinquivalent, etc., if one atom of the element will combine with or replace four or five atoms of hydrogen, respectively. In compounds, as shown above, the total units of valence shown by one element always equal the units of valence of the other. The hydrogen atom never combines with or replaces more than one atom of another element. For this and other reasons it has become a standard atom for comparison in order to express the valence of other atoms. The hydrogen atom, then, has unit valence (one). The valence of an element is the **number** of hydrogen atoms that its atom combines with or replaces.

Chemical Equivalents. In the experiment "Chemical Equivalent of Magnesium," the term "chemical equivalent" was defined as being that weight of the element which would replace or combine with one gram of hydrogen; *i.e.* was chemically equivalent to one gram of hydrogen.

As one atom of an univalent element will combine with or replace one atom of hydrogen, the atomic weight of such an element (in grams) must be equivalent to one gram of hydrogen (atomic weight of hydrogen = 1.008). In other words, the atomic weights of univalent elements are the same as their equivalent weights.

If the valence of an element is two, one of its atoms will replace or combine with two atoms of hydrogen, or its atomic weight in grams is equivalent to two grams of hydrogen, or half its atomic weight is equivalent to one gram of hydrogen. The equivalent weight of a bivalent element is half its atomic weight. By similar reasoning, the equivalent weight of a trivalent element is one-third its atomic weight, while that of a quadrivalent element is one-fourth, etc. The student should work out one of these cases for himself.

From the above statements the following equations may be written:

(1) Equivalent weight = atomic weight ÷ valence.
(2) Valence = atomic weight ÷ equivalent weight.
(3) Atomic weight = valence × equivalent weight.

The atomic weight of an element has been previously defined as the smallest weight of it found in the molecular weight of any of its compounds. As the methods of finding molecular weights are not very exact, the atomic weights derived from them are also inexact. The equivalent weights can be determined with great accuracy, so the approximate atomic weights can be corrected by means of equation (3).

As was pointed out in the experiment on the "Chemical Equivalent of Magnesium," the equivalents are very useful in chemical calculations, for one equivalent of one element will always combine with one equivalent of another element. The term "equivalent weight" is also extended to compounds, indicating that weight of the compound which will enter into chemical action with one equivalent weight of an element. Thus the equivalent weight of sulfuric acid (hydrogen sulfate) is 49 grams, as that weight will interact with 12.16 grams of magnesium (the equivalent of magnesium), giving 1 gram of hydrogen (its equivalent) and 60.16 grams of magnesium sulfate (its equivalent).

Since the metallic and hydrogen radicals (in ionic form) of acids, bases, and salts move toward the negative electrode in electrolysis, they are positively electrified (like electrical charges repel and unlike attract each other), and they are regarded as having positive valence. The non-metallic radicals, which move toward the positive electrode, are negatively electrified and are regarded as having negative valence. If the chemical forces binding the atoms in a molecule are electrical in Nature, as is now generally believed, the sum of the positive electrical charges must be exactly equal to the sum of the negative charges, for the compounds are electrically neutral. Therefore the positive valences must equal the negative valences, making their algebraic sum equal to zero.

Exercises on Valence

I. **Valence Rules.** 1. Hydrogen always has a valence of positive one (+ 1).
2. Oxygen always has a valence of negative two (−2).
3. The metallic radicals are always positive in valence.
4. The non-metallic radicals are always negative.
5. The total of the positive valences equals the total of the negative valences.

II. **Valences in Binary Compounds.** If the formula of a binary compound is known, together with the valence of one of the elements, the valence of the other may be easily computed. Rule 5 may be restated thus: the number of atoms of the positive element times its valence equals the number of atoms of the negative element times its valence. As all the terms of this equality are known except the valence desired, the equation is easily solved. Example: find the valence of iron in ferric chlorid, $FeCl_3$. The valence of chlorin is −1, as it combines with one hydrogen atom to form hydrogen chlorid, HCl. Putting x for the valence of iron, the following equation may be written: $1 \cdot x = 3 \cdot 1$. Therefore $x = 3$.

III. **Valences of Compound Radicals.** The formula of sodium nitrate is $NaNO_3$. As the valence of sodium is + 1, that of the nitrate radical must be −1. The formula of disodium phosphate is Na_2HPO_4. As the two sodium radicals and the hydrogen radical have a total positive valence of 3, that of the phosphate radical must be −3. In this compound it is sometimes convenient to regard the HPO_4 group as being the negative radical (secondary acid phosphate radical), and as it is combined with two sodium radicals, its valence must be −2.

IV. **Making Formulas When the Valences of the Radicals Are Known.** In writing the formula of a compound composed of radicals, the formula of the negative radical is written after that of the positive radical, and such numbers are used as subscripts as will make the total of the positive valences equal to the total negative valences. This may be most easily done by using the valence of the negative radical as the subscript of the positive radical, and the valence of the positive radical as the valence of the negative. This formula is usually simplified if it is possible to do so, by using as subscripts smaller numbers in the same ratio. Example: stannic (tin) sulphid contains the radicals Sn^{++++} and S^{--}, with valences indicated by the plus and minus signs. The formula of the compound is Sn_2S_4, or simplifying, SnS_2.

V. **Valences in Binary Compounds**. After each formula below write the symbols of the elements, omitting the subscripts, and indicate the valence of each by means of plus and minus signs; *e.g.* $FeCl_3 = Fe^{+++}\ Cl^{-}$.

HCl
KCl
$CuCl_2$
CuF_2
Na_2O
NaI
CuS

FeI_3
Fe_2O_3
$PbCl_4$
$NaBr$
$LiBr$
Li_2S
MgS

Al_2O_3
$AlCl_3$
Al_2S_3
SnS_2
Sb_2S_5
$SbCl_5$
As_2O_5

VI. **Valences of Compound Radicals**. After each formula below write the formula of the compound radical inclosed by the parenthesis and indicate the valence in the usual manner. The valences of the simple radicals found in these compounds are as follows: Ag^{+}, Hg^{+}, Rb^{+}, Zn^{++}, Cd^{++}, Sr^{++}, Mn^{++}, Cr^{+++}, Bi^{+++}, Cl^{-}.

$Ag_2(SO_4)$
$Sr(CrO_4)$
$Cd(HSO_4)_2$
$Mn(C_2O_4)$
$Cr(C_3H_2O_2)_2$
$Sr(CO_3)$
$Sr(HCO_3)_2$

$Zn(OH)_2$
$Cr(PO_4)$
$Hg_2(SO_3)$
$Rb(HSO_3)$
$Cd(HPO_4)$
$Sr(H_2PO_4)_2$
$Rb_3(AlO_3)$

$Bi(NO_3)_3$
$Rb_3(AsO_4)$
$Rb_2(HAsO_4)$
$Rb(H_2ASO_4)$
$Cd(ClO_3)_2$
$Rb_2(Cr_2O_7)$
$(NH_4)Cl$

VII. **Correction of Formulas**. The following formulas are incorrect. Rewrite each one in the space to the right, using the valences given or determined in the preceding sections.

$Cd(PO_4)$
$Rb_2(NO_3)_3$
$Zn(ClO_3)_4$
$Ag(CrO_4)_2$

$Sr(OH)_3$
$Ag(SO_3)$
$Mn_3(HSO_4)$
$Rb(C_2O_4)$

$Cr(SO_4)_2$
$(NH_4)_2(CO_3)_3$
$Sr_2(HPO_4)_3$
$Bi_3(HSO_4)$

VIII. **Writing Formulas**. The radicals across the top of the table on the other side of this sheet are positive. Indicate their valences by plus signs as usual. The radicals in the first column are negative. Indicate their valences by minus signs. In the squares write the formulas of the compounds formed by the radicals at the head of the column and at the beginning of the row. Crossed squares indicate that compounds of those radicals are not known. Write small.

	K	Zn	NH_4	Fe^{+++}	Pb	Al	Ag	Au^{+++}
Cl								
CO_3				X X X		X X X		
HCO_3				X X X		X X X	X X X	X X X
SO_4								X X X
HSO_4								X X X
SO_3								
HSO_3								X X X
S								
NO_3								X X X
ClO_3								X X X
PO_4								X X X
HPO_4								X X X
H_2PO_4								X X X
AsO_4								X X X
$HAsO_4$								X X X
H_2AsO_4								X X X
OH								

PURIFICATION OF WATER

Water is one of the absolutely essential substances, not only for the life of both animals and plants, but also for all the arts and industries of men. Natural waters are never absolutely pure. Much of it has been polluted by the carelessness of an ever increasing population, so that it is unsafe as a drinking water. Industrial needs have become so exacting that various treatments must be applied to water before it is suitable for use. For these reasons the processes of purifying water have become of great importance.

For drinking and cooking purposes water must be free from all germs that may cause disease. It has been found easier to keep water pure than to purify it after it has been polluted. Hence the modern emphasis upon sanitation and sanitary living. But until the population as a whole becomes better educated on the subject of sanitation, or there are more stringent laws enacted regarding sanitation, much water must be purified before it is safe for drinking. The large cities, which need water in large quantities, usually obtain it from some near-by river or lake. It is passed through extensive filter beds, which remove suspended matter and many of the germs. Chemicals, like chlorin or chlorid of lime, are sometimes used to kill the germs, or the individual user boils the water to make it sterile. The ultra violet light rays have been found to be effective in sterilizing water. The rain water comes in contact with limestone under ground, dissolves it, and becomes " hard." When used for washing purposes, this hardness wastes much soap, and causes much injury to the steam boilers in industrial plants.

In purifying water the kind of impurities present must first be known. The different kinds of bacteria can be determined only after a very careful analysis. This is not attempted in most cases because of the length of time needed to complete the work satisfactorily. Chemical tests are made instead for chlorids, ammonia, and organic matter which, if present, indicate that the water is polluted with sewage. Such a condition is considered dangerous because of the possibility of the presence of disease-producing bacteria in the sewage. Sulfates and calcium compounds, derived from limestone, make the water hard. If it exceeds a certain degree, the water is chemically treated to " soften " it before use in steam boilers. The tests for chlorids, sulfates, and ammonia are made by mixing chemical reagents with the water so that insoluble compounds will be formed from these substances. The insoluble compounds become visible as solids having characteristic properties, and they are called **precipitates**. The quantity of the precipitate formed may be taken as an index of the quantity of the impurity present.

The directions for making these tests must be carefully followed, or erroneous conclusions will be formed. For example, silver chlorid gives a precipitate with a number of substances besides chlorids. But the chlorids are the only class of these which form a precipitate that is soluble in ammonia and insoluble in nitric acid. Hence, the complete procedure of test should be followed to obtain correct results.

Study Topics: Sadtler, *Chemistry of Familiar Things:* "Water," pp. 82–97.
Baskerville, *Municipal Chemistry:* " Drinking Water and Disease," pp. 20–31; " Purification of Water," pp. 71–89.
Hazen, *Clean Water and How to Get It.*

References: McP. & H., pp. 41–45, 327; B. & W., pp. 95–97; M. & L., pp. 56–60; Smith, pp. 52–53, 235; H. & S., pp. 56–58, 432; Newell, pp. 36–38, 333; Brownlee, p. 51.

EXPERIMENT

Purpose: State it on separate sheet used for the record.

Apparatus: Large bottle or jar, distilling apparatus as in Figure 63 or Figure 64.

Materials: Impure water containing calcium sulfate and sodium chlorid, each 2 grams per liter; 1 cc. of ammonium hydroxid, and sugar 5 grams per liter. Solutions of ammonium oxalate, barium chlorid, silver nitrate, potassium permanganate, and Nessler's reagent. Pumice stone heated red hot and dropped quickly into water.

I. **Sedimentation**. Fill a large, **clean** bottle with water from the tap. Examine it carefully for turbidity, comparing it with distilled water (?). Do not mistake air bubbles for turbidity. If the turbidity is not readily noticeable, stir a little powdered clay into the water and allow the bottle to stand quietly until the end of the laboratory period. Describe what took place (?). Apparently clear water that is allowed to stand quietly for a few days may frequently show a fine sediment at the bottom.

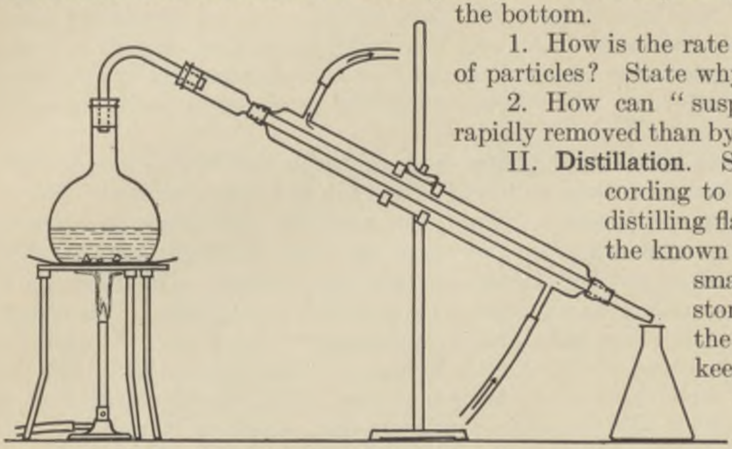


Fig. 63. Distillation with Liebig condenser.

1. How is the rate of sedimentation affected by the size of particles? State why.

2. How can "suspended matter" in water be more rapidly removed than by sedimentation? Define the process.

II. **Distillation**. Set up the distilling apparatus according to one of the given figures. Fill the distilling flask one-half full of water containing the known impurities. Put in the flask 3 or 4 small pieces of the prepared pumice stone to obtain steady boiling. Heat the water to boiling over a gauze, and keep it boiling steadily at a moderate rate. Collect the distilled water (distillate) in **clean** 15-cm. tubes, marking the tubes in the order that they are filled. While the distilling is in progress make the following tests with the impure water.

III. **Tests for Impurities**. (*a*) Calcium (lime). To 10 cc. of the impure water in a tube add about 1 cc. of acetic acid and several cubic centimeters of ammonium oxalate solution, a few drops at a time. An insoluble substance (calcium oxalate) is formed, which is easily seen. Test its solubility in hydrochloric acid by adding about 5 cc. of the acid. The complete test is summarized by stating the **properties** of the precipitate, its **color** and **general appearance**, and its **solubility** in acetic and hydrochloric acids.

(*b*) Sulfates. To another 10-cc. portion of the impure water add 1 cc. of hydrochloric acid and heat the mixture to boiling. Then add several cubic centimeters of barium chlorid solution, a few drops at a time. An insoluble substance (barium sulfate) is formed.

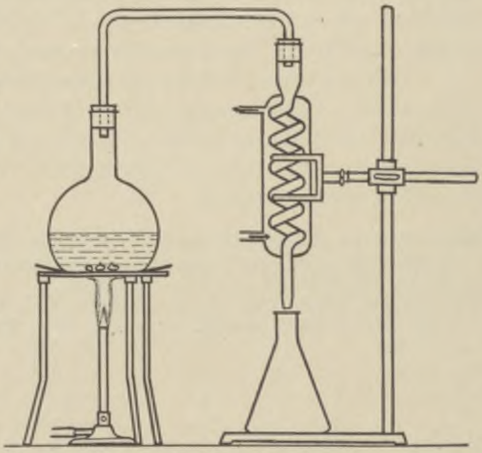


Fig. 64. Distillation with coil condenser.

(*c*) Chlorids. To another 10-cc. portion of the water add 1 cc. of nitric acid and about 1 cc. of silver nitrate solution, a few drops at a time. In soluble silver chlorid is formed. Let it settle for a minute and pour off as much of the liquid as possible, retaining the precipitate. Test the solubility of the precipitate in ammonium hydroxid by adding about 5 cc.

(*d*) Ammonia. To another similar portion of the impure water add 1 or 2 drops of Nessler's reagent solution. A precipitate or coloration is formed, depending on the quantity of ammonia present.

(*e*) Organic Matter. Make about 50 cc. of the water decidedly pink, but not purple, by adding potassium permanganate solution to it in a beaker and stirring. Put half of it in each of two tubes. To one add 2 cc. of concentrated sulfuric acid, and heat contents of both tubes to boiling. Organic impurities act as reducing agents, combining with the oxygen in the permanganate in the presence of sulfuric acid. As the permanganate is decomposed the color fades.

IV. **Impurities in Distilled Water**. Make the above tests with 5-cc. portions of the first tube of the distillate collected. Test also the succeeding tubes for all impurities found in the preceding one. Make a record of results in the provided form.

EXPERIMENT

Purpose:..

....

I. **Sedimentation.**

Questions: 1. ....

2. ....

III. **Tests for Impurities.** Summarize tests as directed for calcium.

(*a*) Calcium ....

....

(*b*) Sulfate ....

....

(*c*) Chlorids ....

....

(*d*) Ammonia ....

(*e*) Organic Matter ....

IV. **Distilled Water.**

Calcium:
Acetic acid and
Ammonium oxalate....................HCl.....................Conclusion..........

Sulfates:
Hydrochloric acid and
Barium chlorid ....Conclusion..........

Chlorids:
Nitric acid and
Silver nitrate..........................NH_4OH..................Conclusion..........

Ammonia:
Nessler's reagent..Conclusion..........

Organic Matter:
Sulfuric acid and
Potassium permanganate...............................Conclusion..........

3. What impurity seems to distill over with the water? Why? (See boiling point of this impurity.)

....

4. How may this impurity be separated from the distilled water?

....

5. Why is distillation the best method of purifying water?

....

6. Define the process of distillation.

....

....

Analysis of Tap Water

Calcium:
Acetic acid and
Ammonium oxalate......................HCl......................Conclusion..........

Sulfates:
Hydrochloric acid and
Barium chlorid..Conclusion..........

Chlorids:
Nitric acid and
Silver nitrate............................NH_4OH...................Conclusion..........

Ammonia:
Nessler's reagent..Conclusion..........

Organic Matter:
Sulfuric acid and
Potassium permanganate....................................Conclusion..........

In further analyses use a form similar to that for tap water.

HYDROGEN CHLORID

This compound was known to the early chemists under a variety of names. Basil Valentine in 1644 described its preparation from sodium chlorid and called it "spirit of salt." Priestley called it "marine acid air" because of its preparation from sea water. An impure solution of the gas in water is known as "muriatic" acid. The term is derived from the Latin word, "muria," meaning "brine."

In the first part of the nineteenth century, when sodium carbonate began to be made by the Leblanc process in England, this compound was given off in large quantities as a chimney gas. It killed vegetation and rendered the air unfit to breathe, so laws were enacted forbidding the soda makers to allow more than 0.2 grain of this gas per cubic foot to be discharged into the atmosphere. The gas was then absorbed in water. Uses were soon found for the acid, and now it seems a necessity of industrial life.

This instance affords a remarkable example of how it is possible to convert a substance causing a nuisance into a necessity. It is possible that all matter is useful to man if it is in the proper form and place. It is the business of the chemist to put matter into this condition. The profits in soda making by the Leblanc process declined very seriously many years ago, but while they declined, the by-product, hydrogen chlorid, became valuable by its increased use. The process is still used, but mainly to manufacture this former worse than worthless by-product. The soda is now the by-product. Many other similar instances could be given of the utilization of industrial by-products. The solving of such problems proves the worth of a chemist.

By the Leblanc process hydrogen chlorid is prepared by the chemical action of sulfuric acid on the cheapest obtainable chlorid, sodium chlorid. In the study of hydrogen it was stated that water was the chief hydrogen compound in Nature. The important class of compounds, the acids, all contain hydrogen as a common element. Of these the ternary acids may be said to derive their hydrogen from water by the combination of it with the acid anhydrids.

$$\underset{\textbf{Sulfur trioxid}}{SO_3} + H_2O \longrightarrow \underset{\textbf{Sulfuric acid}}{H_2SO_4}.$$

Sulfur is very abundant, and since sulfuric acid is easily and cheaply made in this way, the other acids are generally made from it rather than by the anhydrid method. From this it may be said that other acids derive their hydrogen radical from sulfuric acid and indirectly from water. This hydrogen unites at red heat with the chlorin of sodium chlorid, and being gaseous at this temperature, the hydrogen chlorid is carried away through pipes to be absorbed in water, leaving the sodium to combine with the remaining radical of the sulfuric acid.

$$\text{Hydrogen sulfate} + \text{sodium chlorid} \longrightarrow \text{hydrogen chlorid} + \text{sodium sulfate.}$$

It may be noticed that the parts of the reacting substances have been interchanged. This kind of a chemical change is known as **double replacement**. When the mixture of sodium chlorid and sulfuric acid is not heated beyond a moderate temperature, the acid reacts with only half as much of the salt as at the higher temperature. Only one of the two hydrogen radicals of the sulfuric acid takes part in the double replacement action, and the residue left is sodium hydrogen sulfate, $NaHSO_4$.

Since the boiling point of sulfuric acid is 330° C. while that of hydrogen chlorid solution is only a little over 100° C., the latter may be separated from the former by heating it above its boiling point, but not enough to bring the temperature of the mixture near 330° C.

Hydrogen chlorid can be prepared synthetically, but with no source of free hydrogen and chlorin in nature, this method has been used only in the student laboratory until recently. Since electrochemical processes have come into general use, these two gases are now on the market as by-products from the preparation of sodium and other metals. The two gases are very active and combine with but little urging.

Study Topics: *Effect of Light on a Mixture of Hydrogen and Chlorin. A General Method of Preparing Acids. The Chlorin Industry.* Thorp's *Industrial Chemistry*, pp. 98–121.

Bird, *Modern Science Reader*: "Making Money Out of Waste," pp. 11–16.

References: M. & L., pp. 86–92; Smith, 94–100; Brownlee, pp. 80–87; McP. & H., 175–180; B. & W., pp. 129–139; H. & S., pp. 108–114; Newell, pp. 77–81.

HYDROGEN CHLORID

EXPERIMENT

Purpose: State it on the record sheet.

Materials: Sodium chlorid (commercial), magnesium ribbon, methyl orange solution, litmus paper, pine splinter, litmus solution.

I. **Preparation**. Pour a 15-cm. test-tube, full of concentrated sulfuric acid, into half as much water in a beaker. Pour slowly and stir the mixture. **Caution!** Do not pour the water into the acid. Note the evolution of heat caused by the acid forming hydrates with the water. Let the mixture cool while setting up apparatus. Weigh out about 25 grams of sodium chlorid and put it into the flask. Insert a two-hole stopper, carrying thistle and delivery tubes, into the mouth of the flask; and secure it by a clamp over a wire gauze at sufficient height to permit placing a burner underneath for heating (Fig. 65). The thistle tube **must** extend almost to the bottom of the flask so that the end will be under the liquid poured in, or the generated gas will escape through it. Put the delivery tube through the center of a piece of cardboard and have its lower end reach the bottom of a clean, dry receiving jar. Pour the diluted cool acid into the flask through the thistle tube and heat flask **slowly**. Hold some moist blue litmus paper at the mouth of the jar. When the paper turns red, the jar is filled with the gas and should be replaced promptly by another. Collect three or four jars of the gas and cover them with glass plates. After filling the last jar, place the end of the delivery tube just below the surface of about 30 cc. of water in a beaker or jar. Continue heating, and look toward the light at the water at end of delivery tube. Make diagram of apparatus, describe action in generator, and write the correct equation.

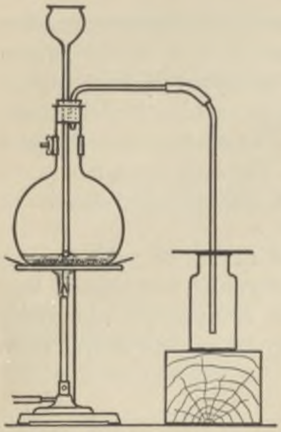


Fig. 65. Preparation of hydrogen chlorid.

II. **Properties**: (*a*) Note the color of gas, if any (?).

(*b*) Smell the gas cautiously if you have not already done so (?).

(*c*) Invert a jar of gas in a pan of water, remove plate, and note what happens (?). Explain. Describe what was seen at end of delivery tube and explain (?).

(*d*) Compare the density of the gas and air, remembering the manner of displacement in collecting (?).

(*e*) After 10 minutes passing the gas into water, stop heating, unstopper the flask, and blow the breath across its mouth (?). Blow the breath similarly across the mouth of an unstoppered bottle of concentrated hydrochloric acid (?). The action is similar to the formation of a fog.

(*f*) Plunge a blazing splinter into a jar of the gas (?).

(*g*) Put a few drops of concentrated ammonium hydroxid in an empty jar, cover, and shake it. Invert it over a jar of the collected gas, having the mouths of the jars together. Remove both cover plates (?). Ammonium chlorid, NH_4Cl, is formed. Write the equation (?).

(*h*) Pour a little of the solution of the gas on 1 cc. of litmus solution in a test-tube, also on a few drops of methyl orange solution in a separate tube. Give the color changes (?).

(*i*) Pour some of the solution of the gas on a 3-cm. piece of magnesium ribbon in a 10-cm. tube (?). Test the gas set free with a flame (?). Write the equation.

Questions: 1. Does any of the sulfuric acid distill over with the hydrogen chlorid? Make a decision on this point, then test the solution of the gas for a sulfate to verify your answer (?).

2. Why was not the gas collected by displacement of water?

3. The effect of the gas and the solution of the gas on the indicators, litmus and methyl orange, shows that it belongs to what class of substances?

4. Summarize the five physical properties of hydrogen chlorid as shown above.

5. Summarize the five chemical properties of hydrogen chlorid and its solution.

CHLORIN

The element chlorin does not occur in Nature in a free state. It can be prepared from its compounds with sodium, potassium, and magnesium, which are quite common, but they are so stable that considerable energy must be applied to them to separate the chlorin. This is most readily done by means of electricity, and practically all of the chlorin now on the market is obtained as a by-product in the manufacture of magnesium, sodium hydroxid, and potassium hydroxid by the electrolysis of the chlorids named above. The chlorin appears at the positive electrode, is purified, dried, liquefied, and marketed in steel cylinders. The electrolysis of the chlorids, however, is impracticable for the preparation of small quantities used, in the laboratory.

Only the chlorids of gold and platinum furnish chlorin by heating, and these compounds are too costly for laboratory use. The cheaper chlorids, being more stable, require more energy to separate the chlorin from them. If hydrogen chlorid is used, the necessary energy can be obtained from unstable compounds rich in oxygen. They are called oxidizing agents. The oxygen with the aid of its accompanying energy separates the hydrogen from the chlorin to form water and free chlorin.

When metallic elements are present in the oxidizing agent, they combine with chlorin, according to their valences, to form chlorids of the metals. So the free chlorin obtained by this method is that which remains after satisfying these metals.

The writing of the equations for these oxidations is simply an application of the foregoing statements. The action of sodium chromate on hydrogen chlorid may be considered as an example, producing chlorin, water, sodium chlorid, and chromium chlorid. Skeleton equation:

$$Na_2CrO_4 + HCl \longrightarrow Cl_2 + H_2O + NaCl + CrCl_3.$$

As there are four oxygen atoms on the left, there must be four molecules of water on the right to use all of them. Four molecules of water contain eight hydrogen atoms, which can be obtained only from eight molecules of hydrogen chlorid. Two sodium atoms and one chromium atom on the left will make two molecules of sodium chlorid and one of chromium chlorid, respectively. All together there are eight chlorin atoms in the hydrogen chlorid taken, of which five are combined with the sodium and chromium, leaving three chlorin atoms free, or $\frac{3}{2}$ Cl_2 molecules.

$$Na_2CrO_4 + 8\ HCl \longrightarrow \tfrac{3}{2}\ Cl_2 + 4\ H_2O + 2\ NaCl + CrCl_3.$$

The fractional coefficient may be eliminated by doubling all the coefficients, which gives for the balanced equation:

$$2\ Na_2CrO_4 + 16\ HCl \longrightarrow 3\ Cl_2 + 8\ H_2O + 4\ NaCl + 2\ CrCl_3.$$

In writing the formulas of the chlorids produced in this experiment, the following valences should be used: potassium (K), sodium (Na), and hydrogen (H) are univalent; lead (Pb), magnesium (Mg), copper (Cu), and manganese (Mn) are bivalent; antimony (Sb), chromium (Cr), iron (Fe), and phosphorus (P) are trivalent.

Study Topics: Thorpe, *Essays in Historical Chemistry:* "Carl Wilheim Scheele," pp. 60–78. Smith, *General Inorganic Chemistry:* "Bleaching," pp. 269–270.

References: Smith, pp. 106–114; M. & L., pp. 83–86; McP. & H., pp. 169–174; H. & S., pp. 98–106; Brownlee, pp. 71–77; Newell, pp. 73–77; B. & W., pp. 164–174.

EXPERIMENT

Purpose: State it on the sheet used for the record of the results.

Apparatus: Dropping funnel, 10 gas jars and glass covers, hydrogen generator.

Materials: Potassium permanganate, antimony powder, copper foil, iron picture wire, sodium, white phosphorus, candle, litmus paper, figured calico, potassium chlorate, potassium dichromate, lead dioxid, lead tetroxid, magnesium oxid, manganese dioxid.

I. **Preparation and Properties of Chlorin.** (Instructor.) (*a*) Fill 10 medium-sized dry gas jars with chlorin gas, either from a cylinder of liquid chlorin, or by slowly adding diluted hydrochloric acid (1 water : 3 acid) from a dropping funnel upon 30 grams of potassium permanganate crystals in a 250-cc. flask (Fig. 66). Fill the jars by downward displacement of air, holding a white sheet of paper behind the jars to see when they are full. Remove the wooden block while changing jars. Cover the jars with glass plates and set them mouth up on the table. Sodium hydroxid solution in beaker absorbs waste chlorin. Describe the action in the generator and make a diagram of the apparatus.

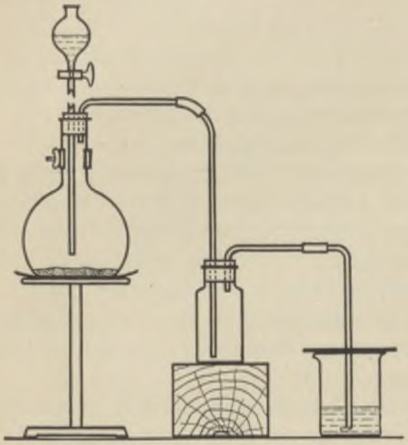


Fig. 66. Chlorin prepared by oxidation of hydrochloric acid.

(*b*) Slip aside one of the glass plates and fill the jar one-third full of cold water, cover it tightly with the palm of the hand, and shake vigorously (?). Lower the jar into a pan of water and remove the hand (?). Explain.

(*c*) Place a very thin, 1-cm. square slice of clean sodium in a jar of the gas and cover at once (?). After a half hour cautiously taste the **white** product. Name the product and write the equation.

(*d*) Sprinkle some antimony powder into a jar of the gas (?). Name the product and write the equation.

(*e*) Heat a loose wad of copper foil or of iron picture wire and immediately plunge it into a jar of the gas (?). Name the product and write the equation.

(*f*) Introduce a small piece of white phosphorus into the gas by means of a combustion spoon (?). Name the product and write the equation.

(*g*) Introduce a burning jet of hydrogen (test the hydrogen for purity before igniting it) into a jar of chlorin (?). Blow gently across the mouth of the jar (?). Name the product and write the equation for the burning.

(*h*) Lower a lighted candle into a jar of the gas by means of a wire handle (?). Blow gently across the mouth of the jar (?). Identify two products of the action.

(*i*) Put dry litmus paper, figured calico, and newspaper with writing ink marks on it in a jar of the gas. Now put some moistened pieces of the same articles in another jar of the gas. In a few minutes compare the contents of the two jars (?).

Questions: 1. What are the four physical properties of chlorin?

2. Summarize the action of chlorin with metals, non-metals, hydrogen, and hydrogen compounds.

II. **Study of the Oxidation of Hydrogen Chlorid.** (Hood.) To 2 cc. of concentrated hydrochloric acid in a 20-cm. test-tube add a pinch of powdered potassium chlorate (?). Identify the gas by its color and effect on moist blue litmus paper. Empty the test-tube into the water in the sink as soon as possible after observation. Repeat the experiment, using potassium dichromate, magnesium oxid, lead dioxid, lead tetroxid, manganese dioxid, and potassium permanganate, one at a time, instead of the potassium chlorate. **Note**: To relieve the irritation by chlorin, sprinkle alcohol or ammonia water on a handkerchief and inhale the fumes. Tabulate the results under the headings: oxidizing agent, color of gas, action on moist litmus paper, name of gas. Warm the tube slightly if the action is slow. Remember that heating may drive off steam and hydrogen chlorid.

Questions: 3. What element is present in each of the substances added to the hydrochloric acid? See their formulas below.

4. Do all compounds of this element liberate chlorin from hydrochloric acid? Why?

5. Write the equations for the actions according to the instructions in the introduction. The formulas are $KClO_3$, $K_2Cr_2O_7$, MgO, PbO_2, Pb_3O_4, MnO_2, $KMnO_4$.

NEUTRALIZATION

Metal + oxygen ⟶ basic oxid (anhydrid) + water ⟶ base (hydroxid).

Non-metal + oxygen ⟶ acidic oxid (anhydrid) + water ⟶ acid (hydroxid).

Base + acid ⟶ water + salt.

The above summary recalls how bases and acids, the two most fundamental classes of compounds, have been made. It is remarkable that their properties should be so different, since both have hydrogen and oxygen (except the binary acids) as constituent elements; but in bases the oxygen and hydrogen make up the negative radical $(OH)^-$, while acids have the hydrogen as the positive radical $(H)^+$ and the oxygen is a part of the negative radical.

Potassium hydroxid $K^+(OH)^-$. Hypochlorous acid $H^+(OCl)^-$.

From these considerations it may be said that the hydroxid of a metal is a base, and that the hydroxid of a non-metal is an acid.

The physical properties of acids and bases are not distinctive enough for use as tests. Their chemical behavior toward certain organic substances, called indicators, affords a simple and sure means of knowing their presence. They show a decided color change when mixed with an acid or a base. Litmus is a well-known example. Methyl orange, cochineal, and phenolphthalein are in common use also.

A remarkable transformation takes place when the proper proportions of an acid and a base are mixed. There is little external evidence of a chemical change except the rise in temperature of the mixture, but an indicator present shows that there is neither an acid nor base now left in the mixture. Their properties have disappeared, so the acid and base have ceased to exist. For example, sodium hydroxid and hydrochloric acid, both violent poisons, when mixed form two necessary articles of food, water and sodium chlorid. Such are the wonders of chemical change.

The reaction is called neutralization because of the neutral properties of the products as compared with those of the acid and base. When the solvent water is driven off by heating, a solid is left which is found to be composed of the metal from the base and a non-metallic part from the acid. In some cases the same salt without water can be formed by action between the anhydrids, but since water increases the activity, the reaction is usually brought about between the hydroxids. With binary acids, which have no anhydrids, action may take place between the elements themselves to form a salt; as,

$$2\,Na + Cl_2 \longrightarrow 2\,NaCl.$$

The heat of the neutralization has been measured. The unit quantity of heat is the **calorie**. It is the amount of heat necessary to raise the temperature of one gram of water one degree centigrade.

$$KOH + HNO_3 \longrightarrow KNO_3 + H_2O + 13{,}700 \text{ calories.}$$

With other active acids and bases in dilute solution the **same amount of heat is liberated in the formation of a gram-molecular weight of water and of the salt**, whatever active acid or base is used. Since the water is the only common compound and the same weight is formed in each case, this uniform quantity of heat evolved must be the **heat of the formation of water** from the hydrogen of the acid and the hydroxyl radical of the base. From this it appears that water and not the salt is the chief product in the neutralization of an acid and base.

Study Topics: *The Liming of Soils. Antidotes for Poisons.*
References: Smith, pp. 129–136; M. & L., pp. 88, 174, 265; Brownlee, pp. 139, 153–155; B. & W., pp. 207–216, 306; H. & S., pp. 142–159; McP. & H., pp. 109, 110; Newell, pp. 142–144.

,A BASE AND NEUTRALIZATION

EXPERIMENT

Purpose: Write the purpose on a separate sheet with heading.
Material: Distilled water, sodium, litmus paper, solutions of methyl orange and phenolphthalein.

I. **Preparation of a Base.** Fill a clean evaporating dish half full of distilled water, and drop into it a freshly cut 5-mm. piece of metallic sodium. Protect the face from flying sodium by holding a glass plate near the dish (Fig. 67), and observe the action through it (?). When the action has ceased, dip into the liquid in the dish strips of both colors of litmus paper (?).

FIG. 67. Sodium reacting with water.

Questions: 1. What kind of substance is present in the liquid in the dish? Give a reason for the answer.

2. Write the equation of the chemical action.

3. Write the equation for the preparation of this base by the anhydrid method and compare the two actions.

II. **Properties of a Base.** (*a*) Try the feeling of the solution of a base by dipping the thumb and forefinger in the liquid and rubbing together (?). The feeling is due to soap formed by the action of a strong base upon the oil of the skin.

(*b*) Touch the forefinger to the tongue and note the taste of the base (?). Eject the tasted material into the sink. Remember to taste no substance except when directed to do so.

(*c*) Test the effect of the base on other indicators, methyl orange and phenolphthalein, by adding a drop or two of each to a little of the solution of the base in separate 10-cm. test-tubes. Record the color of the indicator that is due to the action of the base (?).

III. **Neutralization of a Base.** To the remainder of the solution in the dish add two drops of phenolphthalein solution and then dilute hydrochloric acid slowly with stirring until the red color of the indicator disappears. This change of color marks the " end point " of the chemical action at which the acid and base have destroyed each other. Note if the dish and contents feel warmer just after the action (?). Evaporate the solution to dryness and taste the residue when it is cold (?). Write the equation, showing as one of the products of the action the 13,700 calories of heat given off. By similar equations show the neutralization of potassium hydroxid, KOH, with nitric acid, HNO_3, and ammonium hydroxid, NH_4OH, with acetic acid, HCO_2CH_3 (?).

Questions: 4. State two methods of making a base.
5. Summarize the properties of a base as shown above.
6. What is the residue from the evaporated solution?
7. By the formation of what compound may the **same amounts of heat liberated** be explained?
8. Define neutralization.

Optional: If time permits, neutralize 20 cc. of potassium hydroxid solution with dilute nitric acid, using two drops of methyl orange solution as an indicator. Evaporate the water and examine the residue (?).

Neutralize 20 cc. of ammonium hydroxid solution with acetic acid or dilute sulfuric acid, using two drops of cochineal or litmus solution as an indicator. Evaporate solution to dryness and examine the residue (?).

THE NAMING OF SALTS

Exercise

The names of salts are compound words. The first word is the name of the metallic radical and the second is derived from the characteristic radical of the acid from which the salt is made.

I. **Binary Acids and Salts.** **Hydro**chlor**ic** acid forms chlor**id** salts.
Hydrosulfur**ic** acid forms sulf**id** salts.

Binary salts always have the ending "id." In salts the metal may be considered as having replaced the hydrogen of the acid. The number of hydrogen atoms that any metal can replace is called its valence. The valences of the metals in the first, second, and third columns below are one, two, and three, respectively. Write the correct formulas and then the names of the following salts, arranging the work neatly in two columns.

The hydrochloric acid (HCl) salts of Au, Pb, and Al.
The hydrosulfuric acid (H_2S) salts of Ag, Co, and Cr.
The hydriodic acid (HI) salts of Na, Ca, and Bi.
The hydrobromic acid (HBr) salts of Li, Ra, and Fe.

II. **Ternary Acids and Salts.** **Hypo**chlor**ous** acid forms **hypo**chlor**ite** salts.
Chlor**ous** acid forms chlor**ite** salts.
Chlor**ic** acid forms chlor**ate** salts.
Perchlor**ic** acid forms **per**chlor**ate** salts.

The prefixes " hypo " and " per " are retained in the names of the salts. Acids ending in " ous " form salts ending in " ite," and ternary acids ending in " ic " form salts ending in " ate." Following this system, write the correct formulas first in a column, and then the names of the following salts:

The nitric acid (HNO_3) salts of K, Ba, and Bi.
The carbonic acid (H_2CO_3) salts of Li, Pb, and Fe.
The phosphoric acid (H_3PO_4) salts of Ag, Zn, and Al.
The silicic acid (H_4SiO_4) salts of Na, Mg, and Al.
The sulfurous acid (H_2SO_3) salts of Na, Cd, and Cr.
The hypochlorous acid ($HClO$) salts of K, Zn, and Bi.
The sulfuric acid (H_2SO_4) salts of Ag, Cd, and Sb.

III. **Acid and Basic Salts.** Since a monobasic acid has only one replaceable hydrogen atom, only one kind of salt can be made; for if a metal replaces any hydrogen in a molecule of the acid, there can be none left. With a dibasic acid one hydrogen atom in the molecule may be replaced and the other left, $NaHC_2O_4$, sodium bioxalate. This forms a salt that is somewhat intermediate in its properties between the acid and the normal salt in which all replaceable hydrogen has been replaced.

With monacid and diacid bases the same conditions are true with respect to the hydroxyl radical (OH), except that the product has properties intermediate between those of a base and a salt, so they are called basic salts. An example is basic lead nitrate, $Pb(OH)NO_3$. The relations of normal, acid, and basic salts may be conveniently expressed graphically as in Fig. 68. The normal salts are produced when equivalent quantities of acid and base are used. One equivalent of a base with two equivalents of a dibasic acid makes an acid salt. One equivalent of an acid with two equivalents of a diacid base makes a basic salt.

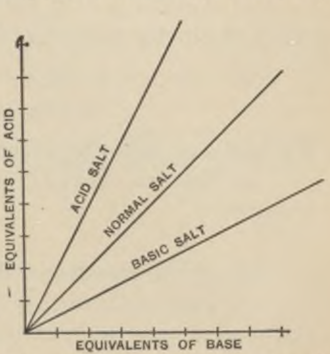


FIG. 68. Graphic relationship of acids, bases, and salts.

Write the formulas and names of the acid sodium salts of the dibasic acids in I and II above; the basic nitrates of Pb, Zn, Ca, and Ba; the basic carbonates of Bi, Al, and Pb.

WAYS OF MAKING SALTS

The total number of salts known is very large, and many of them find more or less important applications in the industries, the arts, and in medicine. Some are relatively unimportant by-products of the manufacture of other substances, while others are the main products of great factories. There are a number of different chemical actions by which salts are produced, the most important being given below.

I. **Binary Salts by Direct Combination**. Many of the binary salts can be made by the direct combination of their elements.

$$2\,Na + Cl_2 \longrightarrow 2\,NaCl.$$

In some cases higher temperatures than the ordinary must be used to induce the action. Write the equations for the preparation of the following salts: ZnS, $CuCl_2$, KI, $MgBr_2$, $SbCl_3$, FeS, $ZnCl_2$.

II. **Replacement**. Metals above hydrogen in the activity list will replace hydrogen from acids. Example:

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2\uparrow.$$

Write the equations for $FeSO_4$, $ZnCl_2$, $Mg(NO_3)_2$, $SnCl_2$, $AlCl_3$.

III. **Double Replacement**. This is by far the most commonly used reaction in the making of salts. It must be borne in mind that double replacement actions are reversible actions and do not run to completion unless one of the products of the action is removed as a gas or precipitate, or for some reason does not give back the radicals from which it was formed and which are necessary for the reverse action. For a satisfactory method of preparing a salt, one must bring together under the proper conditions two substances which (1) contain the radicals of the salt desired, (2) will enter into a complete reaction, (3) and will give products which can be readily separated from each other. The third requirement is easily satisfied if one of the products is a gas or readily vaporizable substance, or an unstable compound giving a gas on decomposing, or if one product is insoluble under the conditions prevailing when it is formed. The table of solubilities on page 241 will help in choosing substances which will satisfy all the requirements.

(*a*) Acid on a base, giving a salt and water. The water can be evaporated, leaving the salt. The action is complete because the water does not give back the hydrogen and hydroxyl radicals it is formed from, and which are necessary for the reformation of the acid and base. Examples:

$$2\,NaOH + H_2SO_4 \rightleftarrows Na_2SO_4 + 2\,H_2O.$$

$$NaOH + H_2SO_4 \rightleftarrows NaHSO_4 + H_2O. \quad \text{(Excess of acid.)}$$

Write the equations for the preparations of $ZnSO_4$, $Ca(HSO_3)_2$, $MgCl_2$.

(*b*) Acid on a salt, giving a new salt and a new acid. A precipitate or gas must be produced to make the method applicable. Examples:

$$HCl + AgNO_3 \rightleftarrows HNO_3 + AgCl\downarrow.$$

$$ZnCO_3 + H_2SO_4 \rightleftarrows ZnSO_4 + H_2CO_3 \rightleftarrows H_2O + CO_2\uparrow.$$

Write the equations for preparing $Ca(NO_3)_2$, AgI, $PbCl_2$.

(*c*) Salt on another salt, giving two new salts. One of the new salts must be insoluble. Example:

$$NaCl + AgNO_3 \rightleftarrows NaNO_3 + AgCl\downarrow.$$

Write the equations for preparing $ZnSO_4$, $CaCO_3$, $AgBr$, $BaSO_4$.

(*d*) Base on a salt, giving a new base and a new salt. One of the products must be insoluble, or the new base must be unstable and give a gas on decomposing. Examples:

$$Ca(OH)_2 + Na_2CO_3 \rightleftarrows CaCO_3\downarrow + 2\,NaOH.$$

$$3\,NaOH + FeCl_3 \rightleftarrows Fe(OH)_3\downarrow + 3\,NaCl.$$

$$2\,NH_4NO_3 + Ca(OH)_2 \rightleftarrows Ca(NO_3)_2 + 2\,NH_4OH \rightleftarrows 2\,H_2O + 2\,NH_3\uparrow.$$

Write the equations for preparing $BaSO_4$, $NaNO_3$, NH_4Cl, Na_2SO_4.

AIR AND NITROGEN

The early philosophers had very vague ideas about the nature of the atmosphere. The Greeks attributed its movements to the work of the god, Æolus. In Shakespeare's time it was thought that witches had control of the winds. These superstitious notions were the conceptions of persons ignorant of many of the facts and most of the laws of Nature.

The scientific study of the atmosphere presented more than ordinary difficulties. Being in gaseous form, it could not be felt except when in motion. It has no color, no odor, and four-fifths of it is chemically inert; so it was a trying subject for mental grasp. Although the atmosphere is a mixture of gases in no fixed proportions, yet samples of it from different localities showed so little variation that it was long thought to be a simple elemental substance. Now it is known that molecular motion and the winds keep the gases of the atmosphere so diffused that their proportions are remarkably uniform. Our present knowledge of it represents the accumulation of facts sifted out from numerous fictions through the centuries in many a test and trial by experiment.

Nitrogen is an essential constituent of all living matter, which as it lives decomposes. The energy stored in the nitrogen compound helps to furnish some of the energy which enables an animal to move. In particular, the nerves are composed of nitrogenous substances. Their sensitivity and liability to change give the nerves the ability to receive stimuli and transmit them to the brain. It seems a reasonable belief that each impulse which flashes along a nerve represents an almost explosively rapid decomposition of a train of nitrogenous molecules.[1] The broken-down molecules make up the body wastes which are excreted and eliminated. This waste material is utilized as a food by plants, and it is built again, with energy from the sun, into nitrogenous compounds which are food for animals. From these facts it may be concluded that molecules which contain the nitrogen atom are storehouses of energy. The **inactivity** of nitrogen, or its reluctance to combine with other elements, gives **instability** to the molecule into which it may be forced by the application of much energy.

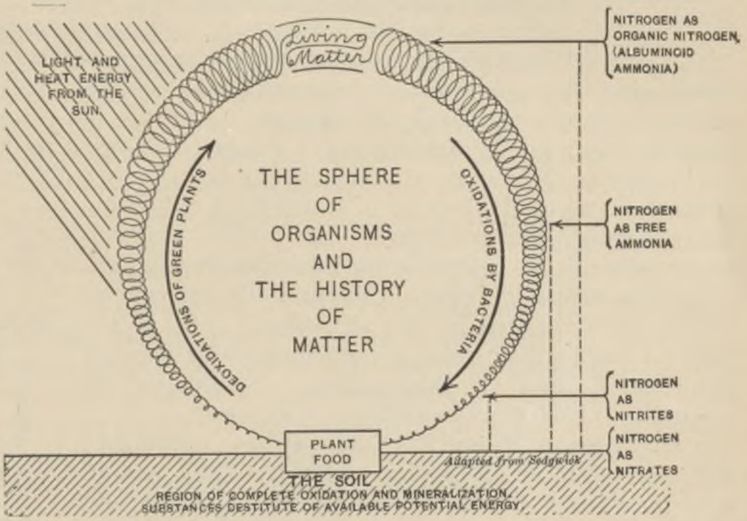


FIG. 69. The complete cycle of organic nitrogen.

Study Topics: Martin, *Triumphs and Wonders of Modern Chemistry:* "The Air," pp. 126–188; "The Element Nitrogen," pp. 189–205. *Modern Chemistry and its Wonders:* "The Romance of Explosives," pp. 51–88.
Duncan, *The Chemistry of Commerce:* "The Fixation of Nitrogen," pp. 42–66.
Bird, *Modern Science Reader:* "Modern Explosives," pp. 17–27.
Ramsay, *The Gases of the Atmosphere:* "Phlogisticated Air," pp. 121–146.
Baskerville, *Municipal Chemistry:* "Ventilation," pp. 337–355.

References: Smith, pp. 189–199; M. & L., pp. 128–135; H. & S., pp. 169–181; McP. & H., pp. 86–94; Brownlee, pp. 221–228; Newell, pp. 86–88, 104–111; B. & W., pp. 67–77, 373–374.

EXPERIMENT

Purpose: Write it on a separate sheet.

Apparatus: Rubber stopper #5, piece of glass tubing, glass plug, 16 cm. of 6 mm. diameter rubber tubing, pinch clamp.

Materials: Solution of pyrogallic acid (20 grams to 100 cc. of water), solution of potassium hydroxid (250 grams in 1 liter of water), gummed labels.

---

[1] Mathews, *Physiological Chemistry.*

I. **Nitrogen from Air.** Assemble the apparatus as in Figure 70. The rubber tubing connects the stem of the funnel with a short piece of glass tubing inserted in one of the holes of the stopper. The funnel should be held at such a height that the test-tube just touches the base of the stand. Remove the stopper and fill the 20-cm. test-tube with water and empty it. The tube is now filled with a fair sample of air of the room. Stand the tube upside down to drain. See that the rubber tube is closed by the pinch clamp, then pour into the funnel a mixture of 5 cc. of the pyrogallic acid solution and about 25 cc. of the potassium hydroxid solution. Place a beaker under the apparatus and cautiously open the pinch clamp to allow the liquid to fill the entire length of the tube. Insert the stopper securely in the tube and close the other hole by the glass plug. Make sure that all connections are air-tight. Move the pinch clamp up on the funnel stem. Some of the liquid will flow at once into the test-tube below. More will follow as the acid absorbs the oxygen from the air and turns black. If all the liquid flows into the tube, the connections are not air-tight. Close the rubber tube with the clamp, return the liquid to the funnel, and repeat the experiment more carefully from the beginning. Fasten a test-tube holder to the tube and raise it to an inverted position so that the air may be brought into better contact with the liquid. After inverting the tube about 20 times during a period of 5 minutes, hold it so that the levels of the liquid in the tube and funnel are at the same height. Close the rubber tube with pinch clamp while levels are at same height. Paste a strip of a label on the outside of the tube to mark the position of the bottom of the stopper. The remaining gas in the tube is practically all nitrogen. Remove the stopper and insert a blazing splinter (?). Pour the black liquid carefully into a graduate, allowing the tube to drain for a minute. Record the volume as the volume of oxygen in the air. Rinse out tube and graduate, then find the volume of the tube up to the bottom of stopper as marked, by filling it with water and using the graduate. This volume represents the volume of air taken for the analysis. Record in a tabular form. Compute the per cent of oxygen and nitrogen in the air. About 1 % of the residual gas is argon, which is very similar to nitrogen in properties. Show all computations. Summarize the properties, physical and chemical, of nitrogen.

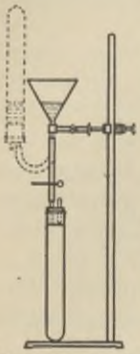


Fig. 70. Analysis of air.

II. **Nitrogen from Its Compounds.** (Optional.) Put about 7 g. of ammonium chlorid, NH_4Cl, and 10 g. of pulverized sodium nitrite, $NaNO_2$, in a flask and equip it as in Figure 71. Pour into the flask about 50 cc. of water. Heat the flask **slowly** and collect three jars of the gas over water, throwing away the contents of the first, since it is impure. Have a vessel ready (large beaker or water bath) half filled with cold water. If the contents of the flask foams much and appears to be about to overflow, swing the supporting ring with the wire gauze to one side and bring the vessel of cold water up under the flask until the lower part is in the water. Regulate the flow of gas in this manner. Ammonium nitrite is formed by the action. It decomposes into nitrogen and one other product. Write two equations for this action. Observe the physical properties of the gas and test its behavior toward combustion (?).

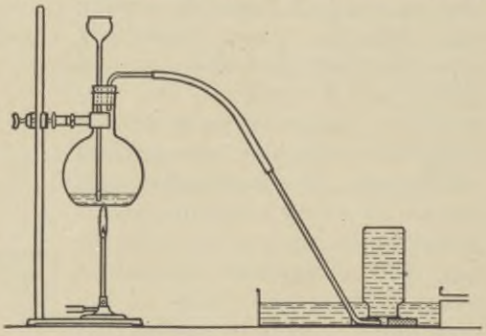


Fig. 71. Nitrogen from the decomposition of ammonium nitrate.

Questions: 1. From the degree of heat used in the experiment, judge the stability of ammonium nitrite.

2. What constitutional evidence is there that this compound might be expected to be more unstable than most nitrogen compounds?

AMMONIA AND AMMONIUM HYDROXID

Although the sharp odor of ammonia is distinctly noticeable around stables and other places where plant and animal wastes are decaying, it was not isolated and described until 1774, when Priestley discovered it and called it "alkaline air." Later it received the name **ammonia** from sal ammoniac, a compound from which it is quite easily prepared. In ancient times considerable quantities of sal ammoniac were imported into Europe from Egypt, especially from the neighborhood of the temple of Jupiter Ammon which stood in an oasis in the northern part of the desert of Sahara. Caravans crossing the desert rested here. As wood was scarce, the refuse from the stables where the camels were kept was used as fuel. From the soot a white crystalline substance was obtained, which was called sal ammoniacum, salt of Ammon.

Formerly when deer were much more numerous than now, the antlers which they shed every year could be picked up in great quantities. These horns of the hart, a general term for deer, were utilized by being dry distilled to make animal charcoal. A by-product of this distillation was ammonia, hence it came to be called **spirits of hartshorn**. Most of the ammonia used at present is obtained by a similar distillation of coal. Formerly most of the ammonia was combined with acids and changed into ammonium salts, especially the chlorid and sulfate, but now much of it is used in refrigeration and in making the ammonia water used as a cleanser.

It will be recalled that when sulfuric acid was heated with sodium chlorid, an exchange of radicals took place, giving hydrogen chlorid and sodium hydrogen sulfate. An acid and a salt gave a new acid and a new salt. In an exactly similar manner, a base and a salt will enter into a double replacement, giving a new base and a new salt. When a mixture of an ammonium salt and a soluble, non-volatile base is heated, the following action takes place:

$$\text{Ammonium salt} + \text{metal hydroxid} \rightleftarrows \text{metal salt} + \text{ammonium hydroxid.}$$

$$\text{Ammonium hydroxid} \rightleftarrows \text{ammonia} + \text{water.}$$

The escape of the ammonia gas allows the second action to run to completion, and as this removes the ammonium hydroxid produced in the first action, this can also run to completion.

It will be noted that the metal and the ammonium radical (NH_4) have exchanged places. By this action the ammonium radical is seen to be in the same class as the hydrogen and metal radicals, in that they unite with the non-metallic radicals, so is written first in the formulas of ammonium compounds. An especial effort must be made not to confuse the substance ammonia, NH_3, and the radical ammonium, NH_4, which does not occur free but only in combination with non-metallic radicals.

References: H. & S., pp. 183–194; Smith, pp. 200–206; M. & L., pp. 150–155 and 335–348; McP. & H., pp. 121–127; Brownlee, pp. 233–239; Newell, pp. 89–94; B. & W., pp. 358–362.

EXPERIMENT

Purpose: State it on separate sheet used for record of results.

Materials: ammonium chlorid, ammonium nitrate, ammonium sulfate, slaked lime (calcium hydroxid), magnesium oxid, gelatin, soda-lime (sodium hydroxid and calcium oxid), "Ammo."

I. **Preparation**. Fill a 20-cm. test-tube one-third full of a mixture of equal volumes of ammonium chlorid and hydrated lime (slaked lime or calcium hydroxid). Insert a one-hole stopper carrying an L-tube. Clamp the test-tube as shown in Fig. 72 so that the L-tube extends upward nearly to the bottom of a dry jar which is standing mouth down on a cardboard on the ring of the stand. The test-tube should incline slightly downward to the mouth so that condensed moisture may not run back and cause the tube to crack. Leave a space above the mixture to the end of tube. Provide a small test-tube with a few drops of concentrated hydrochloric acid, and fill the water pan. Warm the mixture gently with a small flame at first and then with a larger flame, which should be kept moving continually. Dip the end of the glass rod in the hydrochloric acid, and hold it close to the hole in the card to find out when the jar is full. Dense smoke will show this. Explain this test (recall the hydrochloric acid experiment). When the jar is full, remove and set it mouth down on a glass plate on the table. Fill a dry test-tube with the gas in the same way, cover it with the thumb, and place it as directed in II (*a*). Fill a second dry jar with the gas. Make a diagram of apparatus, and write the correct equation.

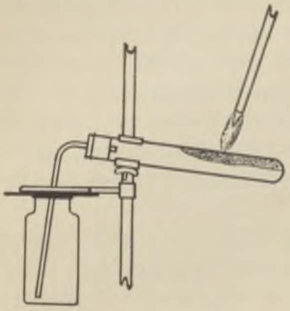


FIG. 72. Ammonia from an ammonium salt.

II. **Properties.** (*a*) Lower the test-tube, mouth down, into the water and remove the thumb (?). Is the gas soluble in water? Compare its solubility with that of chlorin and hydrogen chlorid.

(*b*) Remove the second jar as soon as it is full, turn the L-tube down, connect the calcium chlorid tube to it by a short piece of rubber tubing, and let the gas bubble through 15 cc. of water in a beaker for 10 minutes. Does the gas dissolve readily? Look through the beaker horizontally, observe the movement of the solution, and state whether the solution is lighter or heavier than water. Pour the solution into a bottle, cork it tightly and save it for (*d*). Meanwhile proceed with (*c*).

(*c*) Use one jar of the gas to determine the color, taste, and odor of the gas, if you have not already noticed them. Raise the other jar from the table, mouth down, and bring up to it, and then into it, a burning splinter (?). State whether the gas burns or supports combustion.

(*d*) Record the color, odor, taste, and feeling (rub a little between the fingers) of the solution made in (*b*). Test it with litmus papers (?). What is it? Write the equation for its formation, its formula being NH_4OH.

(*e*) Pour half of the solution into an evaporating dish and boil off about half the liquid. Smell the residue (?), and test it with litmus papers (?). Explain the result by contrasting it with that in (*b*). If the equation of the preceding paragraph is not in accord with your statements, correct the equation. This compound is sometimes called the "volatile alkali." Explain.

(*f*) Put a bit of litmus paper in the rest of the solution, and slowly add dilute hydrochloric acid until neutral. Stir constantly. Evaporate the resulting solution to dryness in an evaporating dish. What is the residue? Make the proper tests for the two radicals (?). Write the equation for the neutralization and for the tests. Heat the residue in the dish to a red heat (?).

III. **Sources of Ammonia.** (*a*) Apply gentle heat to a mixture of very small equal amounts of gelatin and soda-lime (sodium calcium hydroxid) in a test-tube, noting the odor of the gases given off and also their effect on moistened litmus paper. Almost all organic compounds containing nitrogen give ammonia when heated with soda-lime, so this is a test for such compounds.

(*b*) Note the odors of ammonium chlorid and slaked lime (?). Rub the two together in the palms of the hands, and note the odor (?) and the effect on moistened litmus papers held over the mixture. Write the equation.

Repeat, using ammonium sulfate and magnesium hydroxid (moistened magnesium oxid). Repeat with ammonium nitrate and sodium hydroxid solution, using a test-tube instead of the palms.

(*c*) Note the color, odor, etc., of dry "Ammo" powder. Moisten a little of it and again note the odor (?). State the probable composition of "Ammo" and explain its action. Try to verify your statements by whatever tests may suggest themselves.

Questions: 1. Give four physical properties of ammonia, and four of ammonium hydroxid.
2. Give four chemical properties of ammonia, and three of ammonium hydroxid.
3. By what test may ammonium salts be detected?

ANALYSIS OF HOUSEHOLD AMMONIA

Acids are important ingredients of certain substances, such as vinegar and baking powder. Bases are present in washing compounds and household ammonia. The value of such substances depends largely on the amount of the acid or base that they contain. In almost every locality it is possible to find brands of household ammonia on the market having only one-tenth the ammonia content that others have, yet they are retailed at the same price. It is not possible for the average person to judge the quality of the ammonia by appearances, so the consumer depends on the grocer's honesty and fair dealing. The grocer himself frequently does not know the quality of the goods that he sells. Hence, the market-place phrase of the ancients, " Caveat emptor " (Let the buyer beware), is still applicable at times.

It is possible for one with a little knowledge of chemistry to make an accurate determination of the quality and value of ammonia by the method known as **titration**. By this method a small accurately measured volume of the ammonia is neutralized by a solution of an acid whose concentration is known. Then from the volume of acid used, the content of the ammonia is calculated on the basis of chemical equivalence. The solutions of known concentration used are prepared having **one gram equivalent** of hydrogen per liter and are called **normal** solutions. A normal solution of a monobasic acid then will contain the molecular weight in grams of the acid in one liter of water, and the normal solutions of dibasic and tribasic acids will contain one-half and one-third of the molecular weight in grams of the acids respectively in a liter of water. All normal solutions of acids, bases, and salts are similarly prepared on the basis of chemical equivalence. Consequently, all normal solutions are chemically equal, volume for volume. For example, 1 cc. of a normal solution of any acid will exactly neutralize 1 cc. of a normal solution of any base. The ease and rapidity of analysis by this method gives it so much favor with chemists that the method is employed wherever the conditions are such as to permit its use.

In order to know when the exact quantity of the normal solution needed to react with the unknown substance has been added, an indicator is used. These substances are very sensitive and show a remarkable change of color with less than a drop of the solution used. When the unknown substance has been completely used up by the chemical action of the known solution, any slight excess used will be shown by the color change of the indicator. The careful operator can determine this " end point " of the chemical action to a fraction of a drop. To make sure of an accurate analysis several trials are made, or more, until a close agreement of the results of analysis shows that it has been conducted properly.

After the analyses of the class are finished, the results may be compared. The poorest brands of household ammonia are noted in order to avoid them, and the best noted so that the best brands may be purchased. The value of this information will be appreciated if the student considers what the dealer's attitude toward him will be when he has a better knowledge of the quality of the goods than the dealer himself.

Many other substances may be analyzed like ammonia. For finding the per cent of acetic acid in vinegar a normal solution of some base would be used. If the unknown substance is but slightly acidic or basic, a tenth normal or hundredth normal solution would be used. A great many substances used in the home, like soaps, washing powders, tooth powders, vinegar, fruit juices, baking powders, toilet preparations, medicines, inks, shoe blacking, etc., can be mixed with water and the liquid tested with litmus paper to determine whether a soluble acid or base is present. Its quantity may be determined by titration with a normal solution.

EXPERIMENT

Purpose: State it on a separate sheet.

Apparatus: 50-cc. burette, 10-cc. pipette.

Materials: 100 cc. of ammonia solution from the home or grocer's, normal solution of sulfuric or hydrochloric acid, solution of methyl orange, indicator.

Each student should bring from home or obtain at the grocer's about 100 cc. of household ammonia. Put about 20 cc. of distilled water in a clean beaker. Get a 10-cc. pipette and rinse it with the ammonia solution, throwing away the rinsings. Measure out exactly 10 cc. of the ammonia with the pipette (Introduction, p. 5) and add it to the water in the beaker. Put 2 drops of methyl orange solution in the mixture.

Fasten the burette (Fig. 17), after thorough rinsing with the solution to be used in it, upright by a clamp with its lower end just high enough to allow a beaker to be placed beneath it. Fill the burette with the normal acid solution. Open wide the tube for an instant at the cock to fill the tip with the solution. Adjust the surface of the solution to the zero mark of the burette, then place the beaker containing the ammonia sample on a piece of white paper under the burette and let the acid run slowly, with stirring, into the beaker until the indicator shows by a color change that enough acid has been added.

In this preliminary trial too much acid will be used. In the next trial the acid is run rapidly into the ammonia sample while stirring, until a volume of acid 2 cc. less than that used in the preliminary trial has been used. Then the acid should be added drop by drop, so that it may be seen when a single drop changes the color of the indicator. This is the end point of the chemical action. Proceed with other trials until the results show agreement to within 0.5 cc. Make at least three trials after the preliminary one. Record each reading at once in the given form. Omit preliminary trial from record and average.

If more than 10 cc. of the normal acid is required to neutralize the 10 cc. of the ammonia sample, then the ammonia solution is more concentrated than a **normal** ammonia solution, and the volume of normal acid required to neutralize 1 cc. of the ammonia expresses the multiple of normality. If less acid than ammonia is required for neutralization, the number of cubic centimeters of the acid required for 1 cc. of ammonia expresses the fractional normality of the ammonia solution. In either case the product of the normality of the ammonia and 1.7 % (the per cent of ammonia by weight in a normal solution, 17.0 g. per L.) gives the per cent of ammonia in the household ammonia.

Burette reading at the end of second titration	cc.
Burette reading at beginning of second titration	cc
Volume of normal acid used to neutralize 10 cc. ammonia	cc
Burette reading at the end of third titration	cc.
Burette reading at beginning of third titration	cc.
Volume of normal acid used to neutralize 10 cc. ammonia	cc.
Burette reading at the end of fourth titration	cc.
Burette reading at beginning of fourth titration	cc.
Volume of normal acid used to neutralize 10 cc. ammonia	cc.
Average volume of normal acid to neutralize 10 cc. ammonia	cc.
Volume of normal acid to neutralize 1 cc. of ammonia	cc.
Multiple or fractional normality of ammonia	
Normality of ammonia × 1.7 % equals the per cent of ammonia	%
Brand of ammonia .... Maker	

THE THEORY OF IONIZATION

Most of the acids, bases, and salts when in water solution have properties which are different not only from the properties of all other substances, but also from their own properties in the dry state. The great increase of chemical activity manifested by acids, bases, and salts in solution was shown in an early experiment by the sodium bicarbonate and cream of tartar mixture when it was put in water. Mention has already been made of the fact that these substances are composed of radicals, and that many of the chemical actions in which they engage consist in a mere exchange of these radicals. These facts, with others discussed in the following pages, have given rise to much speculation among scientists. Various hypotheses as to the reasons for the peculiar properties of these substances have been offered, but the only one which has survived the test of trial and experiment is that known as the **Ionization Theory**.

I. **Boiling Points of Solutions**. The boiling point of a pure liquid is a very definite temperature and may be used in establishing the identity of an unknown substance, or in testing the purity of a known one. The presence of even relatively small amounts of other substances produces a decided change in the temperature at which a liquid boils. Solutes which are not themselves volatile always raise the boiling points of their solvents. The following table gives some of the increases as shown by careful experiment.

SOLUTE	WEIGHT OF SOLUTE	FRACTION OF MOLECULAR WEIGHT	VOL. OF WATER	RISE OF BOIL. PT.
Sucrose ($C_{12}H_{22}O_{11}$)	34.2 g.	$\frac{1}{10}$	100 cc.	0.52° C.
Sucrose ($C_{12}H_{22}O_{11}$)	68.4 g.	$\frac{1}{5}$	100 cc.	1.04° C.
Glucose ($C_6H_{12}O_6$)	18.0 g.	$\frac{1}{10}$	100 cc.	0.52° C.
Glucose ($C_6H_{12}O_6$)	36.0 g.	$\frac{1}{5}$	100 cc.	1.04° C.
Glycerin ($C_3H_5(OH)_3$)	9.2 g.	$\frac{1}{10}$	100 cc.	0.52° C.
Glycerin ($C_3H_5(OH)_3$)	18.4 g.	$\frac{1}{5}$	100 cc.	1.04° C.

These figures show that the rise of the boiling point produced by any given substance is directly proportional to the weight of solute present in a given volume of the solvent. But with different solutes it is apparent that quite different weights are needed to produce the same elevation of the boiling point of the same amount of water. However, these weights are all the same fraction of the molecular weight of the solutes. As the molecular weight of a substance taken in grams represents 6.3×10^{23} molecules, the fraction used must have represented the same number of molecules in each case, and the second weight of each solute contained twice as many molecules. Raoult, in 1886, asserted that the **elevation of the boiling point is directly proportional to the number of dissolved particles present**. If $\frac{1}{10}$ the molecular weight of a solute will elevate the boiling point of 100 cc. of water 0.52° C., it will require one molecular weight of solute in 1000 cc. of water to produce the same rise of the boiling point, and this is often given as a definition of the molecular weight of a substance.

Acids, bases, and salts, however, are abnormal in their effect on the boiling point of their solvent; for one molecular weight of sodium chlorid in 1000 cc. of water raises the boiling point 0.94° C. This is nearly twice (181 %) the normal increase. If it is true that the elevation of the boiling point is due to the dissolved particles, only one conclusion can be drawn from this fact, namely, that there are nearly twice (181 %) as many dissolved particles as there are in a solution of sucrose, for example, containing one molecular weight per 1000 cc. and raising the boiling point but 0.52° C. The most reasonable way to account for the extra 81 % of particles is to assume that 81 % of the molecules of the sodium chlorid have separated into two smaller particles. These sub-molecules are called **ions**.

Those acids, bases, and salts which are composed of two radicals give nearly twice the normal rise of the boiling point. Those having three radicals give nearly three times the normal rise, and those having four radicals give nearly four times the normal rise. The molecules give as many ions as there are radicals, and so it is reasonable to think that the ions are the separated radicals.

Solutes lower the freezing point of the solvent, and all except acids, bases, and salts produce the same depression, namely, to −1.89° C. when one molecular weight of the solute is dissolved in 1000 cc. of the solvent. This is called the normal effect. Acids, bases, and salts produce abnormal depressions approaching two, three, and four times the normal effect, according as they contain two, three, or four radicals. The solution of sodium chlorid mentioned above freezes at −3.42° C., which is again 181% of the normal lowering. This leads to the same conclusion as that drawn previously, that 81% of the molecules separate into their radicals or ions.

Some acids, bases, and salts give solutions which raise the boiling point of the solvent or lower its freezing point only slightly more than the normal effect. Following out the reasoning already employed, one is forced to the conclusion that these substances have only a small per cent of their molecules separated into ions.

If successive equal volumes of water are added to solutions of salt and sugar made by dissolving one molecular weight of each in one liter of water, it is found that the elevation of the boiling point of the salt solution approaches more and more nearly twice the elevation found in the sugar solution at the same dilution. If, on the other hand, successive equal volumes of water are removed from the two solutions by evaporation, the elevation of the boiling point drops nearer and nearer to that of the sugar solution. This is readily explained by the ionization theory; for the ions will meet less frequently as they are given more room to move about in by the addition of more water, hence an increasingly larger per cent of the molecules must be separated into ions. As the amount of solvent is diminished, however, the ions are crowded together, the collisions become more frequent and result in the formation of a larger per cent of molecules, until the substance is entirely in the molecular form, when all the solvent is gone.

II. **Conductivity of Solutions.** (*a*) Fill the tumblers of the apparatus shown in Fig. 73 with the following solutions: hydrochloric acid (1 acid : 4 water), sodium chlorid (10%), sodium hydroxid, alcohol (10%), sugar solution (10%), and distilled water. Insert the electrodes in each cell, see that all connections are made so that the electricity will have a chance to pass through each cell, and then close the circuit. The brightening of a lamp means that the electricity is passing through that cell. Disconnect those cells through which the current is passing to see whether this will have any effect on the others (?). Then add a few drops of hydrochloric acid to any cell through which the current has not passed. Record.

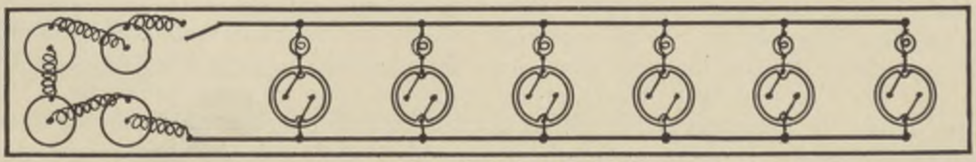


Fig. 73. Conductivity of solutions.

....

....

Numerous experiments like the above with various solutions prove that only solutions containing ions (those which give abnormal boiling and freezing points) conduct the electric current. The current then must be carried through such solutions by the ions.

(*b*) Empty and rinse the tumblers of the apparatus used above, and fill four of them to the same depth with the following solutions: hydrochloric acid (1 acid : 3 water), sulfuric acid (1 acid : 11 water), orthophosphoric acid (1 of 85% acid : 8 water), and acetic acid : (10 acid : 58 water). Insert the electrodes and close the circuit. Note and record any differences in the brightness of the lamps, as the brightness depends upon the conductivity of the solutions. These solutions are equally concentrated (about 3 N) and any differences in conductivity must be due to differences in the degree to which the acids are ionized.

....

....

When the conductivities of the solutions are compared with the elevations of the boiling point of the solvent, it is found that good conductors have large abnormal effects, while the poor conductors have nearly normal effects. There is an exact parallel between the conductivity of the solutions and the degree of departure from the normal elevation of the boiling point. This means that **the conductivity is proportional to the number of ions present**.

(*c*) Rinse a tumbler and its carbon rods with distilled water, replace the rods, adjusting them so that they reach the bottom of the tumbler and are as far apart as possible. Connect it in series with a lamp and a source of electricity. Then pour in just enough concentrated hydrochloric acid to make the lamp glow a dull red. Pour in distilled water and stir with a glass rod, noting any change in the brightness of the lamp (?).

....

....

Any change in the brightness denotes a change in the conductivity, which can be due only to a change in the number of ions. When the ions are crowded together, they are meeting frequently, and many such encounters result in the formation of molecules. In dilute solutions, such encounters must occur less frequently, hence the ionization must be more nearly complete.

(*d*) Fill a U-tube nearly full of a 15% solution of sodium sulfate colored with neutral litmus solution. Support by a clamp in an upright position and insert a platinum electrode in each arm, immersing it but slightly. The electrodes may be disks of platinum foil or flat spirals of platinum wire. Close the circuit and let the current run till there is a decided change in the color of the litmus solution at the two electrodes.

....

The gas at the negative electrode (cathode) is hydrogen, that at the positive electrode (anode) is oxygen. Which of the ions of the sodium sulfate, Na_2SO_4, is shown by the change in the color of the litmus to have moved to the cathode?

....

Which of the ions of the sodium sulfate has moved to the anode?

....

The fact that the current sorts out the ions can be explained only by the assumption that the ions are themselves charged with electricity of opposite kind to that of the electrode to which they are attracted. The hydrogen ions of acids and the metallic ions of salts and bases are always attracted to the cathode, and must therefore be positively charged, while the hydroxyl ion of bases and the non-metallic ions of acids and salts are always attracted to the anode and must therefore be negatively charged. As the solution shows no other signs of being charged with electricity, the total amount of positive charges must exactly equal the total amount of negative charges.

In electrolysis the ions move to the oppositely charged electrode, give up the electrical charges, and then become atoms or groups of atoms. What happens thenceforth depends upon the chemical properties of the atoms or groups of atoms produced. At the cathode, one of three actions may occur: (1) hydrogen (from acids) escapes as a gas; (2) inactive metals like copper, silver, etc., are deposited on the cathode as a film; (3) active metals like sodium, calcium, etc., act on water, giving hydrogen and the hydroxid of the metal.

$$2\ Na + 2\ H_2O \longrightarrow 2\ NaOH + H_2.$$

Three actions are possible at the anode: (1) If the anode consists of a metal capable of forming a compound with the radical constituting the negative ion, it gradually disappears, forming a salt which immediately ionizes and so continues the action. (2) If the anode cannot be attacked, some of the negative ions, after losing their charges, escape as gases or decompose into gases and other substances. (3) Others act upon water to form the acid of which they constitute the negative radical, and liberate oxygen.

$$2\ SO_4 + 2\ H_2O \longrightarrow 2\ H_2SO_4 + O_2.$$

III. **Importance of the Theory to Chemistry**. This theory offers a satisfactory explanation of a number of characteristics of acids, bases, and salts which would otherwise be difficult to understand.

(*a*) Acids owe their common properties to their hydrogen ions, and their individual properties either to their negative ions (anions) or to their molecules. Strong acids are those which are highly ionized.

(*b*) Bases owe their common properties to their hydroxyl ions, and their individual properties either to the positive ion (cation) or to their molecules. The strong bases are those which are highly ionized.

(*c*) Salts owe their properties to their ions, mainly, though some also to their molecules. Although formed by the action of an acid on a base, they are neither acid nor base because they have neither hydrogen nor hydroxyl ions.

(*d*) Double replacement occurs so readily between acids, bases, and salts in solution because the molecules are already more or less completely separated into the ions which can reunite — positive with negative — to form new combinations. In general the products tend to separate again into ions, so that a reverse action is at once set up which soon comes to a state of equilibrium with the forward action. Hence such an action will not run to completion unless the reverse action is in some way prevented. This occurs if one of the products fails for some reason to give the ions from which it was formed and which are necessary for the reverse action. Three cases are known: (1) Some substances are not ionized at all, or else to such a slight degree as to be negligible, like water, very weak acids and bases, and some salts. (2) Other substances are insoluble and form precipitates, so that the actions run to completion. (3) In the last case gases are formed at either the ordinary or slightly raised temperatures.

IV. **Résumé of the Ionization Theory**. When acids, bases, and salts are dissolved in water or certain other solvents, (1) some of their molecules are dissociated into smaller particles called ions, which lead an independent existence in the solution. (2) Some of the ions are positively charged with electricity, the others negatively, the total positive electricity equaling the total negative electricity. (3) The degree of ionization depends upon the solvent, the solute, and upon the dilution. (4) The abnormal boiling and freezing points are due to the presence of ions, and the degree of abnormality proportional to their number. (5) The conductivity of solutions is due to the ions and proportional to their number. (6) Ionization is a reversible action, increasing with the dilution. (7) The chemical properties of acids, bases, and salts are mainly due to the ions they dissociate into, and depend upon the degree of this dissociation.

V. **Comparison of Theory and Facts**. Take a sheet of paper, divide it into two columns, and head one **Theory** and the other **Facts**. In the Theory column write the seven assumptions of the theory, and in the Facts column write the facts upon which the assumptions are based.

FIG. 74. Section of the coal measures in Ohio. The black bands represent coal beds. (U. S. Geol. Survey.)

At the time of the middle ages of the earth's history only small portions of the present continents were visible above the ocean waters. Their surfaces were bleak, rocky exposures without verdure or life. Bird, beast, and insect were still creatures of the future. There was, as yet, no food for them. The atmosphere, heavy with the immense quantities of carbon dioxid, was unfit to be breathed.

The most important epoch in the physical history of our planet was about to begin. Chemical action began to operate miraculously to transform a verdureless and voiceless world into a clothed and animated one. The interior heat of the earth gave the surface a tropical temperature. The rapidly forming soil, the atmosphere reeking with moisture and filled to plenitude with carbon dioxid furnished a most favorable condition for the growth of a luxuriant vegetation. The remains of this vegetation, as found in the coal beds, taxes our imagination in the effort to conceive of its luxuriance. Fern fronds were six to eight feet long, rushes grew thirty feet tall instead of a few feet, and club-mosses attained the heights of our palm trees.

At times parts of the surface of a continent with this vegetation sank below the ocean level. Floods of water overwhelmed it and covered its prostrate remains with pebbles, sand, and mud to great depths. The great pressure of the overlying débris with the consequent heat distilled off the volatile parts of the plant stems and trunks and formed the carbon residue into coal beds. When another uplift of the earth's crust caused this region to emerge from the ocean depths, a new generation of plants began the work of the fixation of the carbon upon a new surface many feet above where the previous one was buried. Thus a complete oscillation of the earth's crust makes a seam of coal among its other rocky layers.

The rate of fixation of carbon from carbon dioxid into coal gives one a conception of the time involved in these earth-making processes. It has been calculated that ten centuries are required for a luxuriant vegetation to make a seam of coal one foot thick. In the coal measures of Nova Scotia one seam is thirty-seven feet thick. Besides this there are seventy-five other seams of varying thickness. This would indicate that over a million years were required to make these beds of coal, leaving out of account the time taken to form the much thicker layers of rock between the seams of coal.

The same coal-making process is going on in the swamps and peat bogs of the earth to-day, but at a very slow rate.

Study Topics: Martin, *Triumphs and Wonders of Modern Chemistry:* "The Element Carbon," pp. 206–233.
Bird, *Modern Science Reader:* "Coal, Its Composition and Combustion," pp. 115–124; *The Romance of the Diamond,* pp. 1–11; *Plant Growth and Decay,* pp. 103–114.
Martin, *The Story of a Piece of Coal.*
References: Consult the index of any available text of chemistry.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Apparatus: Clay crucible, hard-glass test-tube, fitted with stopper carrying L-tube (see Fig. 75), kerosene lamp.

Materials: Animal and vegetable matter, candle, kerosene, brown sugar or solution of cochineal or litmus, animal charcoal, wood charcoal, solution of hydrogen sulfid, copper oxid, tin oxid, lead oxid, lime-water.

I. **Occurrence of Carbon.** On a thin layer of sand in a clay crucible or small iron pan place small pieces of various animal or vegetable material that may be available. Cover them with sand to keep out the air and heat them strongly until all smoking ceases. When the crucible is cool, carefully scrape away the upper layer of sand and examine the residues, recording their color, weight, volume, porosity, combustibility, and name of residue.

Hold a test-tube filled with cold water in the luminous gas flame, a candle flame, and a kerosene flame. Describe the deposit, name it, and tell where it comes from.

II. **Properties of Carbon.** (*a*) Shake 15 cc. of water with enough brown sugar, cochineal, or litmus solution to give the water a decided color. Add 5 cc. of animal charcoal to the solution, shake thoroughly, and boil for three minutes. Filter the solution, and if the filtrate is not clear and transparent, repeat the process, using fresh animal charcoal. If sugar or molasses was used, determine by the taste whether or not the sugar has been removed (?). State the reason for filtering solutions of brown sugar through animal charcoal in the refining of sugar.

(*b*) Fill a test-tube one-fourth full of powdered charcoal (wood) and add 5 cc. of hydrogen sulfid solution. Cork the tube and shake it vigorously from time to time for fifteen minutes. Compare the odor with that of the original solution, and if there is no perceptible difference, add more charcoal and continue shaking until there is a decided difference between the odors of the contents of the tube and of the original solution (?).

(*c*)[1] Mix intimately in a mortar 5 g. of powdered copper oxid, CuO, and 0.5 g. of powdered wood charcoal. Slide the mixture from a creased paper into a hard-glass test-tube, insert a one-hole stopper carrying an L-tube, and provide a test-tube one-fourth full of lime-water. Heat the mixture in the tube gently at first to avoid cracking the tube, and then as strongly as possible, allowing the gas given off to bubble up through the lime-water. Describe the change taking place in the lime-water (milky color indicates carbon dioxid) and draw a conclusion as to the chemical action taking place between the carbon and copper oxid (?). Describe any change in the materials which supports your conclusion (?). After heating about 15 minutes, remove the L-tube from the lime-water and stop heating. When the tube is cold, grind the contents in the mortar with water, letting the stream of water wash away the lighter particles. Describe and identify the residue, and write the equation for the action by which it was produced (?). If any large particles are produced, lay them on a hard surface and strike them with the pestle or a hammer (?).

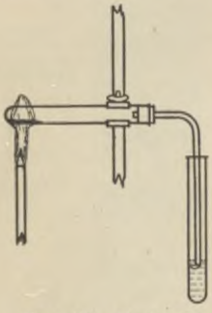


FIG. 75. Reduction of metallic oxids by carbon.

(*d*) Mix 5 g. of tin oxid, SnO_2, and 1 g. of powdered charcoal. Heat this mixture in a hard-glass test-tube as described in the preceding paragraph. Make similar observations and explanations.

(*e*) Perform the same experiment, using an intimate mixture of 20 g. of litharge (lead monoxid, PbO) and 1 g. of powdered wood charcoal. Make the same observations and explanations as in (*c*).

Questions: 1. How may the presence of carbon in organic matter be shown?

2. Why is impure water often filtered through thick layers (1½ to 2 ft. thick) of fresh charcoal?

3. What was oxidized and what reduced in (*c*), (*d*), and (*e*)?

4. What important industrial process depends upon the action in (*c*), (*d*), and (*e*)?

---

[1] Parts (*c*), (*d*), and (*e*) may be worked by different pupils and the results compared.

CARBON DIOXID

"There is, perhaps, no gas more interesting, by reason of its history, its peculiar properties, its effects on animal and vegetable life, and its great abundance and wide diffusion, than carbon dioxid. It is the first gas that was ever carefully studied."

It is being constantly produced by the burning of coal and carbon compounds. A ton of average coal with the necessary oxygen will make over two tons of the gas. It is estimated that 513,523,477 tons of coal were burned in 1914. From these relations one can more readily comprehend the enormous quantity of this gas formed each year. However great this may seem, the amount of this gas formed by the decomposition of carbonate rock (limestone) by heat in volcanic regions is much greater. The quantity of it formed is further increased to the extent of several million tons daily by that which is exhaled by all forms of animal life.

Geologists tell us that in the remote ages of the earth's history the atmosphere contained a very much larger quantity of this gas than now. Its abundance encouraged the growth of vegetation to a wonderful degree. Aided only by the energy of the sun the green chlorophyl in the leaves of plants extracts the carbon from the carbon dioxid and builds it into the leaves and stems of the plant. It is quite certain that the beds of coal in the earth are the remains of this ancient luxuriant vegetation. The heat obtained from the burning of the coal is the heat of the sun that was stored up in the growing plants in that distant age, hundreds of thousands or perhaps a million years ago. This long have these carbon atoms lain dormant in the earth. Being formed into carbon dioxid by burning, these carbon atoms may again be built into the structures of plants and animals, and thrill in living tissues as they have before in many a strange form of flower, bird, or beast.

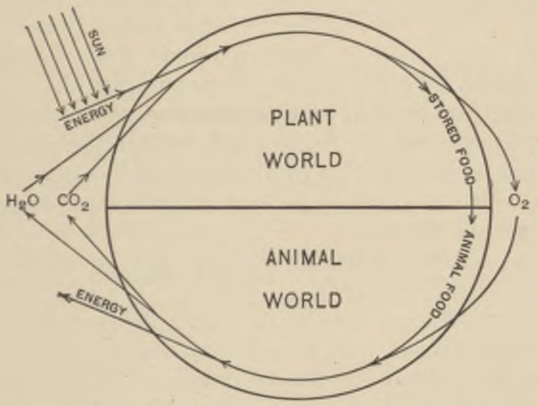


FIG. 76. Cycle of carbon in Nature.

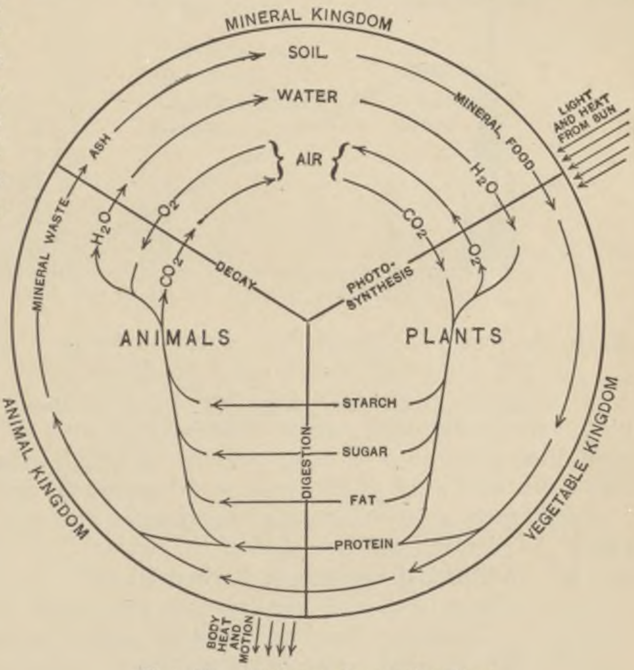


FIG. 77. Metamorphoses of matter.

A common method of preparing any certain acid is to cause some acid to react with a salt containing the radical of the acid desired. If an acid is mixed with a carbonate, then carbonic acid might be expected to be formed. Instead of this product the experiment furnishes water and carbon dioxid. From these facts it is known that the carbonic acid is too unstable to exist under the conditions of the experiment, so it decomposes into water and the anhydrid. There are evidences to show, however, that the acid does exist in solution form at ordinary temperatures. From this relation carbon dioxid has been called "carbonic acid gas."

Study Topics: Martin, *Triumphs and Wonders of Modern Chemistry:* "Carbon Dioxid," pp. 234–264.
Ramsay, *Gases of the Atmosphere:* "Fixed and Mephitic Air," pp. 39–68.

CARBON DIOXID

EXPERIMENT

Purpose: State it on a separate sheet used for the record of results.
Materials: Marble chips, lime-water, 10-cm. strip of magnesium ribbon, wood splinters, white lead.

I. **Preparation.** Fill the generating bottle one-fourth full of marble chips, cover them with water, and insert a two-hole stopper carrying a delivery tube and a thistle tube. The end of the thistle tube must be under the water. Add a little commercial hydrochloric acid from time to time so as to maintain a fairly brisk evolution of gas, and collect five jars of the gas by water displacement. Make a diagram of the apparatus and write the equation for the preparation of the gas (marble is calcium carbonate, $CaCO_3$).

II. **Properties.** (*a*) Light a wood splinter and pour the gas from one jar over the flame as you would water (?). State the relation of the gas to combustion. Compare its density with air.

(*b*) Fill a jar of the gas one-third full of water, avoiding unnecessary loss of gas, cover quickly with the palm of the hand, and shake vigorously. State the evidence of solubility of the gas. Test the water in the jar with both colors of litmus paper (?). Explain the result and give the equation.

(*c*) Pour some lime-water (calcium hydroxid, $Ca(OH)_2$) into a jar of the gas and shake (?). The product may be regarded as the result of action between the calcium hydroxid and the compound formed by the action of carbon dioxid and water. Name the product and write the equation for its formation. This is a test for carbon dioxid.

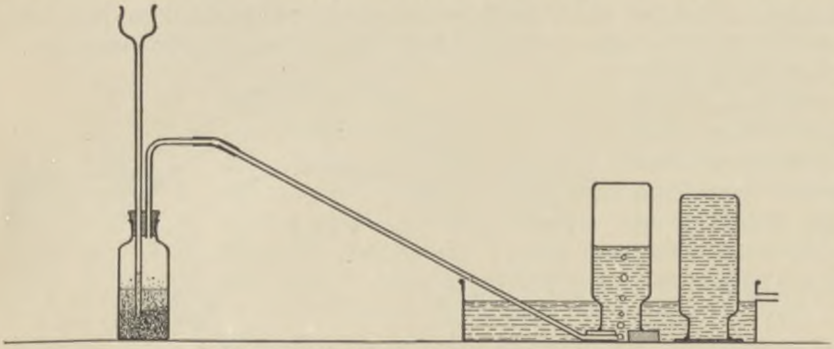


FIG. 78. Carbon dioxid from a carbonate.

(*d*) Set a jar full of air over a jar full of gas, mouth to mouth, and remove the glass plate. After 10 minutes (meanwhile perform (*e*) and clean and put away the apparatus) replace the glass plate between the jars, remove the upper jar with the plate, and set it upright on the table. Pour some lime-water into this jar, cover it and shake vigorously (?). Explain any difference between this result and what you learned in (*a*) about the density of the gas.

(*e*) Ignite a 10-cm. strip of magnesium ribbon and quickly insert it into the last jar of the gas (?). Identify two products of the action (?). Write the equation.

III. **Additional Sources of Carbon Dioxid.** (*a*) Pour 10 cc. of lime-water into a gas jar, burn a wood splinter in the jar (do not drop any ashes), cover the jar with the palm of the hand, and shake (?). Introduce the burning stick again if there is no change in the lime-water. Write the equation for the combustion of the cellulose, $C_6H_{10}O_5$, of the wood.

(*b*) Heat a pinch of white lead (basic lead carbonate) in a 15-cm. test-tube, and test the gas given off by means of a drop of lime-water held in a small loop at the end of a wire. Describe and explain any change in the lime-water. Is a new lead compound formed? If so, try to identify it by comparison with the lead compounds on the side shelf. Write the equation for the action resulting on heating the white lead, regarding it as lead carbonate, $PbCO_3$.

Questions: 1. What five physical properties of the gas are noted?

2. What three chemical properties of the gas are noted?

3. How can you test an unknown substance to determine whether or not it is a carbonate?

4. What must happen to the carbon dioxid before other substances can burn in it? Why is it that most substances cannot burn in this gas?

HOMOLOGOUS SERIES OF HYDROCARBONS

The following named hydrocarbons and many others are known. Their relationship is remarkable in that they differ in properties, not in kind, but in degree only. They are obtained from petroleum, coal, and wood by destructive distillation.

In some of these hydrocarbons, like C_2H_4, the valence (4) of the carbon atom is not fully satisfied by the number of hydrogen atoms. The valences exist, however, as is shown by the fact that active elements like chlorin or bromin are readily taken in combination without displacing hydrogen. Such hydrocarbons are said to be **unsaturated.**

I. PARAFFIN SERIES

Name	Formula	Molecular Weight	Boiling Point[1]	Commercial Name	Use
Gases.					
1. Methane....	CH_4	16	− 164	Natural gas	Illuminant, fuel
2. Ethane.....	C_2H_6	30	− 84		
3. Propane....	C_3H_8	44	− 37	Cymogene	Artificial cold
4. Butane.....	C_4H_{10}	58	+ 1		
Liquids.				Rhigolene	Artificial cold
5. Pentane....	C_5H_{12}	72	+ 37	Petroleum ether	Enriching water gas
6. Hexane	C_6H_{14}	86	69	Gasoline	Motor fuel
7. Heptane....	C_7H_{16}	100	98	Naphtha, ligroin	Cleaning
8. Octane.....	C_8H_{18}	114	125	Benzine	Solvent
9. Nonane	C_9H_{20}	128	150		
10. Decane.....	$C_{10}H_{22}$	142	173		
11. Undecane ..	$C_{11}H_{24}$	158	195		
12. Dodecane ..	$C_{12}H_{26}$	170	214		
13. Tridecane...	$C_{13}H_{28}$	184	234		
14. Tetradecane	$C_{14}H_{30}$	198	252	Kerosene	Illuminating and fuel oil
15. Pentadecane	$C_{15}H_{32}$	212	270		
Solids.					
16. Hexadecane.	$C_{16}H_{34}$	226	287		
* * * *	* * *	* * *	* * *		
			Melting Point	(Vaseline, $C_{19}H_{40}$ — $C_{21}H_{44}$)	
20. Eicosane ... Etc.	$C_{20}H_{42}$	282	37	Paraffin	Candles, sizing

II. OLEFINE SERIES		III. ACETYLENE SERIES		IV. TERPENE SERIES		V. BENZENE SERIES	
Name	Formula	Name	Formula	Name	Formula	Name	Formula
Gases.							
1. Unknown	—	1. Unknown	—	1. Impossible	—	1. Impossible	—
2. Ethylene	C_2H_4	2. Acetylene	C_2H_2	2. Impossible	—	2. Impossible	—
3. Propylene	C_3H_6	3. Propine	C_3H_4	3. Unknown	—	3. Unknown	—
4. Butylene	C_4H_8	4. Butine	C_4H_6	4. Unknown	—	4. Unknown	—
Liquids.							
5. Pentylene	C_5H_{10}	5. Pentine	C_5H_8	5. Valylene	C_5H_6	5. Unknown	—
6. Hexylene	C_6H_{12}	6. Hexine	C_6H_{10}	6. ?	C_6H_8	6. Benzene	C_6H_6
7. Heptylene	C_7H_{14}	7. Heptine	C_7H_{12}	7. ?	C_7H_{10}	7. Toluene	C_7H_8
8. Octylene	C_8H_{16}	8. Octine	C_8H_{14}	8. ?	C_8H_{12}	8. Xylene	C_8H_{10}
9. Nonylene	C_9H_{18}	9. Nonine	C_9H_{16}	9. ?	C_9H_{14}	9. Mesitylene	C_9H_{12}
10. Decylene	$C_{10}H_{20}$	10. Decine	$C_{10}H_{18}$	10. Turpentine	$C_{10}H_{16}$	10. Cymene	$C_{10}H_{14}$

---

[1] Bernthsen.

Members of the olefine series are found naturally in the petroleum at Baku, Russia. They can be prepared from the paraffin series and are obtained from wood and coal by destructive distillation. They are used chiefly in chemical work.

The first known member of the acetylene series is usually prepared from calcium carbid and water. In general they can be made from other series. They are used for illuminating gas.

The terpene series of hydrocarbons are extracted from plants and are the chief constituents of essential oils. They are used for solvents, and in the preparation of oils and perfumes.

The benzene series are derived from coal tar. They are of immense importance in the preparation of dyestuffs, drugs, perfumes, and photographic developers.

Exercise: 1. State the derivation of the words paraffin and olefine.

2. Why is it not easy to distinguish between consecutive members of a series?

3. Compare the number of hydrogen atoms with those of carbon in the first series and explain the meaning of the general formula, C_nH_{2n+2}.

4. Derive a general formula for each of the other series.

5. Explain the basis of the division of hydrocarbons into series.

6. Why are the series called **homologous**?

7. Write the formula of the 25th member of series I, the 20th of series II, the 16th of series III, the 24th of series IV, and the 12th of series V.

Study Topic: Martin, *Modern Chemistry and its Wonders:* "The Romance of Hydrocarbons," pp. 175–224.

EXPERIMENT

Purpose: State it on a separate record sheet.

Apparatus: Dropping funnel.

Materials: Fused dry sodium acetate, soda lime, lime-water, syrupy orthophosphoric acid (sp. gr. 1.75), ethyl alcohol, calcium carbid, turpentine, benzene, bromin water.

I. **Methane**. Mix thoroughly by pulverizing about 10 g. of fused sodium acetate with the same quantity of soda lime. Distribute the mixture along half the length of a 20-cm. tube near the closed end, leaving a space above the mixture to the end of the tube. Fasten the clamp near the mouth to hold the tube in a horizontal position (Fig. 79). Fit the tube with a stopper and delivery tube to collect the gas over water. Heat the tube with a moving flame so as not to melt the glass. Collect two jars of the gas, then quickly replace the delivery tube by an L-tube, having the short arm pointing upward. Continue the heating for a moment, then bring a flame to the end of the outlet tube (?). Hold a large, clean, dry jar inverted over the flame for a minute and note the interior surface of the jar (?). Quickly pour about 20 cc. of lime-water into the jar, cover it, and shake it (?). Use the jar of gas first collected for a study of the physical properties, after constructing a tabulated form for the record similar to that below. Add 10 cc. of bromin water to the second jar, cover and shake the jar (?). This result will determine by the change of color of the bromin water whether the members of this series are saturated or unsaturated hydrocarbons. Let the jar stand until the next day. Make a diagram of apparatus, and write the equation, considering soda lime as sodium hydroxid.

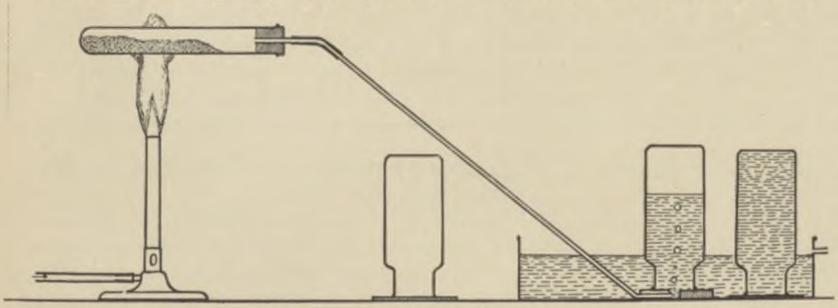


Fig. 79. Preparation of methane.

PROPERTIES

Physical	Chemical
Color..	Combustibility....................................
Odor...	Color of flame..................................
Solubility.....................................	Saturation..
	Equation of combustion.....................

II. **Ethylene.** Pour syrupy orthophosphoric acid into a flask until the liquid is 1 cm. deep. Equip the flask with a dropping funnel (Fig. 80). Put 20 cc. of ethyl alcohol in the dropping funnel and boil the acid until its temperature rises to 225° C. Allow the alcohol to drop slowly into the hot acid, which acts as a dehydrating agent. Keep the flame away from the collecting gas until two jars are collected, then shut off the supply of alcohol. Use the first jar of gas collected for the study of physical properties. Shake the gas with bromin water to determine whether ethylene belongs to a saturated or unsaturated series. Let the jar stand until the next day. Ignite the gas in the other and test the combustion product with lime-water (?). Make diagram, write equation, and tabulate observations as before.

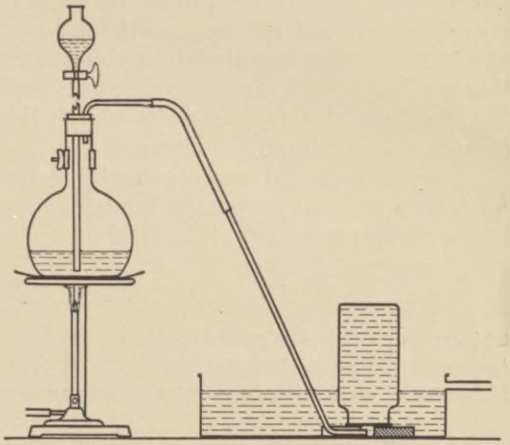


FIG. 80. Ethylene by dehydration of alcohol.

III. **Acetylene.** Invert a jar filled with water in the pan half filled with water. Drop a piece of calcium carbid (CaC_2), size of a pea, into the water and quickly place the mouth of the jar over the carbid (?). Water takes part in the reaction to form acetylene and a hydroxid. Write the equation, tabulate properties, shake with bromin water, and test the combustion product with lime-water.

IV. **Turpentine.** Heat 1 cc. of turpentine, no more, in a test-tube until it boils. Bring the mouth of the tube to a flame (?). Make all the observations and tests except with the lime-water. Tabulate.

V. **Benzene.** Repeat the procedure for turpentine, using benzene instead of it (?). Use no more than 1 cc.

Questions : 1. Why are the combustion products of all hydrocarbons identical?

2. Explain the different luminosity and sootiness of different hydrocarbons.

3. Why were the above five hydrocarbons selected for study?

CARBON MONOXID

There is an acid having one less oxygen atom in its molecule than carbonic acid, and is known as formic acid, $H(CO_2H)$. If this acid is slowly dropped into warm concentrated sulfuric acid, a powerful dehydrating agent, a gas is liberated which is known as carbon monoxid. From this action the gas might be considered as an anhydrid of formic acid. It does not, however, unite with water to make the acid, but it does unite with active bases like potassium hydroxid to make salts.

$$KOH + CO \longrightarrow K(CO_2H) \text{ (potassium formate).}$$

Carbon monoxid can be formed by burning carbon with a limited supply of oxygen. It is half-oxidized carbon. It has been found that this first half of the oxidation of carbon produces 2430 calories of heat per gram, while in the second half, the burning of carbon monoxid to carbon dioxid, 5650 calories are produced. The reason is that much of the heat of burning is used up in changing the solid carbon into the gaseous carbon monoxid. The carbon monoxid, being gaseous at the start, does not absorb any of the heat of combustion but gives off all of it. Consequently this gas has a high thermal value as a fuel, like hydrogen. Since it furnishes about two-thirds of the heat from the combustion of coal, it is important to regulate the burning so that none of this gas escapes up the chimney unburned.

Carbon monoxid can be made by the reduction of carbon dioxid as it passes over red-hot carbon. As anthracite coal burns, the lambent blue flames hovering over the coals are caused by the burning of carbon monoxid.

EXPERIMENT

Object: State it on the record sheet.
Apparatus: Dropping funnel.
Materials: Formic acid, lime-water, litmus, pine splinter.

I. **Preparation.** Pour concentrated sulfuric acid into the 250-cc. flask until it is 1 cm. deep. Fasten it securely by a clamp and place a wire gauze underneath (Fig. 81). Put a 2-hole stopper, carrying the dropping funnel and delivery tube, firmly in the mouth of the flask. Arrange to collect the gas over water. Pour 25 cc. of formic acid into the dropping funnel and adjust the stopcock so that it will drop slowly into the sulfuric acid while it is slowly heated. **Caution! Allow none of the gas to escape in the room and do not smell the gas, for it is poisonous. Shut off the supply of formic acid when the last jar is filled.** Collect three jars of the gas, leaving a little water in the last jar.

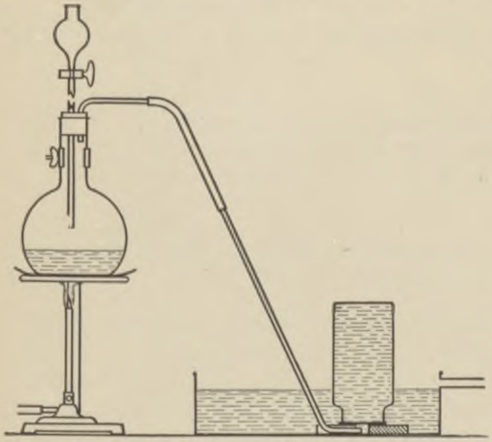


FIG. 81. Carbon monoxid from formic acid.

II. **Properties.** (*a*) The gas has no odor. Record the other physical properties (?).

(*b*) Quickly insert both colors of litmus paper in the jar having the water. Hold the cover on tightly and shake the jar. State any effect on the litmus (?).

(*c*) Pour about 20 cc. of lime-water quickly into one of the remaining jars of the gas, cover tightly, and shake it. State any effect on the lime-water (?).

(*d*) Ignite the gas in these two jars. When the flame goes out, cover and shake again each jar. Explain the effect on the litmus and the lime-water (?).

Remove the stopper from the flask and ignite the gas at its mouth. Burn the remaining jar of gas. Cautiously pour the acid from the flask into water in the sink and let the water run for a time. Return any formic acid left to the stock bottle.

FLAMES

Man is the fire-using animal. Fire has done more than any other one thing for the development of civilization. There are no tribes of savages so low in the scale of development that they do not use fire for warmth and the preparation of food. It seems to have had a strange fascination for primitive man, for in every tribe it was closely connected with his home and religious life.

The nature of fire remained a mystery until after the discovery of oxygen. Before that time it was believed that all materials were composed of varying proportions of the four "elements": earth, air, fire, and water. After the discovery of oxygen, it was learned that oxygen is necessary for fire, and that the burning substance combines with the oxygen, liberating a great amount of heat. In most instances this heat is sufficient to vaporize the substance which is burning, or to decompose it into volatile substances. Flames appear only when a vapor burns. When a combustible substance cannot be vaporized, there is no flame but rather a glow.

Study Topic: Tidy, *The Story of a Tinder Box.*

EXPERIMENT

Purpose: State it on separate sheet used for record of results.
Apparatus: Large lamp chimney fitted with cork and tubes as shown in Figure 82.
Materials: Cardboard, wood splinters, powdered charcoal.

I. **Structure of the Flame**. (*a*) Hold the wire gauze vertically in the middle of the non-luminous burner flame, resting the gauze on the top of the burner. Make a drawing to show the shape of the portion of the gauze which is heated to redness. Now hold the gauze horizontally in the blue flame, first near the top of the flame, above the inside cone, and then through the inside cone, just above the top of the burner. Make drawings of the hot portions. Repeat the operations with the luminous flame, using a heavy cardboard instead of the wire gauze. Remove the cardboard as soon as it becomes slightly charred and make drawings showing the shape of the charred portion of the cardboard.

(*b*) Hold a wood splinter across both flames, first near the top and then through the inside cone near the mouth of the burner. Make diagrams showing how the wood is charred in each position.

(*c*) Hold a piece of fine glass tubing about 15 cm. long at an angle of 45° from the horizontal, so that its lower end is about in the middle of the inner cone. After a short time apply a light to the upper end (?).


FIG. 82. Exploring the inner cone of a burner flame.

Questions: 1. What is the condition of the inner cone of a flame?

2. Of what is the inner cone composed?

II. **Luminosity**. (*a*) State what effect opening and closing the holes in the base of a burner has upon the luminosity of the flame (?). Determine by means of a smoking splinter whether air is entering the holes or gas escaping from them (?).

(*b*) Hold a test-tube full of cold water in the non-luminous flame (?). In the luminous flame (?). Sprinkle some powdered charcoal in the non-luminous flame (?).

Questions: 3. To what is the luminosity of the yellow flame due?

4. Explain the presence of free carbon in the luminous flame.

5. Make and label diagrams showing the various cones of the luminous and non-luminous flames.

III. **Burning and Supporting Combustion**. (Instructor.) Arrange the apparatus as shown in Fig. 83. Fill the chimney with gas and light it at the top. Turn the gas partly off to allow the flame to descend to the air inlet tube at the bottom, then turn the gas on more. Place a wire gauze or perforated asbestos board on the chimney top and have the gas ignited as it issues through it.

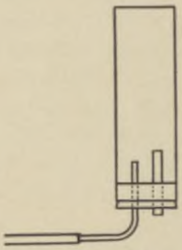


FIG. 83. Conditions of burning reversed.

Questions: 6. What is burning in the gas inside the chimney?

7. Is there any essential difference between the flame at the top of the chimney and that inside?

8. What are the usual meanings of the terms " combustible " and " supporter of combustion "? Can they be reversed?

9. What other combination of gases can be used in such an experiment?

ETHYL ALCOHOL

This alcohol is easily and cheaply made by the fermentation of glucose. The ferment commonly used is a secretion of the yeast cell and is called **zymase**. Starch can be converted into maltose and then into glucose by the action of other ferments. Since starch is cheaper than glucose, it is the general practice to make alcohol from the starch of cereals and vegetables by the successive action of ferments. Many substances of organic nature will not dissolve in water but will in alcohol, so it is used to dissolve drugs and to extract essences and perfumes as well as to preserve them.

EXPERIMENT

Purpose: State it on a separate sheet used for the record of results.
Apparatus: Condenser with connecting tubes, hydrometer, thermometer.
Materials: Corn syrup or molasses, moist compressed yeast, lime-water, pumice stone, iodin solution.

I. **Fermentation and Distillation**. Put about 40 cc. of corn syrup in a 250-cc. flask, add warm water until it is nearly full, and mix by shaking. Make a thin paste of a 1-cm. piece of yeast and mix it well with the syrup solution. Cover the flask with a bit of paper stuck on with a drop of syrup and let the mixture stand quietly for a day or more, then examine the contents of the flask (?). Test the gas in the flask with a lime-water film in a wire loop or pass the gas into lime-water in a test-tube (?). After about 4 days of fermenting, pour off the liquid carefully from the sediment into another vessel and wash out the flask.

Return the liquid to the flask or a larger one, drop in a few pieces of red-hot pumice stone, and insert a two-hole stopper carrying a thermometer and a bent tube for connection with the condenser. Take great care to prevent breakage. Use a condenser as shown in Fig. 84 or Fig. 85. For either of the two forms connect the **lower** jacket tube to some source of cold water and allow the discharge to come from the upper tube. Heat the flask over a wire gauze, and collect the distillate in a bottle or another flask. Note and record the temperature as the first few drops of liquid appear in the condenser and again occasionally as the distillation is continued (?). Distil over about one-half of the liquid, throw away the liquid left in the flask, but save the distillate for further work.

II. **Tests for Alcohol**. (*a*) As soon as some of the distillate is collected pour a few drops into the evaporating dish and try to ignite it (?). If it does not burn, the percentage of alcohol is low. Make the following chemical test carefully.

(*b*) To about 5 cc. of the distillate add 3 or 4 cc. of iodin solution and then sodium hydroxid solution until the brown color of the iodin disappears. Warm the mixture and let it stand for several minutes. The precipitate is iodoform (CHI_3). State the color and odor (?). This precipitate indicates the presence of ethyl alcohol (C_2H_5OH).

(*c*) Hold a clean graduate horizontally and place the hydrometer in it, touching the bottom. Set the graduate upright on the table and pour in the distillate until the hydrometer floats. Read and record the density (specific gravity) indicated on the stem of hydrometer at the surface of the liquid. Find the per cent of alcohol from this density, using the percentage tables, page 246 (?).

Combine two distillates and redistil until a temperature of 98° C. is reached. Repeat the ignition test with the second distillate (?). Determine the per cent of alcohol in this distillate by the hydrometer (?). Save the second distillate in a labeled bottle provided by instructor.

Questions: 1. Explain the use and purpose of the lime-water test, as above.
2. Write the equation of the fermentation of glucose ($C_6H_{12}O_6$).

Study Topics: Bird, *Modern Science Reader:* "The Yeast Cell and Its Lessons," pp. 169–180.
Martin, *Modern Chemistry and Its Wonders:* "Alcohol," pp. 242–261.
Duncan, *The Chemistry of Commerce:* "Industrial Alcohol," pp. 128–157.
Sadtler, *Chemistry of Familiar Things:* "Fermentation," pp. 230–243.

METHYL ALCOHOL

The internal revenue tax, $2.08 per gallon, on ethyl alcohol makes it rather costly. Methyl alcohol, being poisonous, is not used in beverages, so it is not taxed. Although it has a little greater cost of production than ethyl alcohol, it sells for $0.75 per gallon. Because of its cheapness, it is used, wherever possible, instead of ethyl alcohol in the effort to lessen the cost of manufactured articles. Unscrupulous persons yield to the temptation and use it in the preparation of food products and medicinal preparations. Whether it is used externally or internally, blindness, severe illness, or death may result. Its characteristic odor is due mainly to acetone, present in it as an impurity. When this impurity is removed, the alcohol is not readily noticed in lemon extract, liniments, bitters, toilet waters, bay rum, witch hazel, paregoric, and whisky. Pure deodorized methyl alcohol is called Columbian spirits. There is no simple direct test for this alcohol, but there is for formaldehyde, into which it is easily converted.

EXPERIMENT

Object: State it on a separate sheet used for a record.

Apparatus: Coil of copper wire.

Materials: Columbian spirits, 1% solution of resorcin, milk, commercial hydrochloric acid (sp. gr. 1.2), or C. P. acid containing about 5 grams of ferric chlorid per liter, ethyl alcohol.

I. **Conversion of Methyl Alcohol into Formaldehyde.** Mix about 1 cc. of methyl alcohol (Columbian spirits) with 6 cc. of water in a 15-cm. tube that has been placed in a bottle of cold water. Get a coil of copper wire (Fig. 84) and heat it to redness (do not melt it) in the upper oxidizing flame of the burner. Plunge the hot coil quickly into the mixture in the tube (?). Repeat the operation six times. By this operation most of the alcohol will be changed into formaldehyde (HCHO). Compare the odor of the mixture with that of the formaldehyde in the stock bottle (?).

FIG. 84. Coil of copper wire.

II. **Tests for Formaldehyde.** (*a*) Resorcin Test. To 5 cc. of the formaldehyde solution add a drop or two of a 1% solution of resorcin and shake the mixture. Pour this mixture carefully down the inside of a small test-tube having 2 cc. of concentrated sulfuric acid in it. A rose-red layer will form at the junction of the two liquids, indicating the presence of formaldehyde. If the formaldehyde is present in quantity, a turbidity or white precipitate is formed above the colored zone.

(*b*) Casein Test. To 10 cc. of pure milk in an evaporating dish add an equal volume of concentrated hydrochloric acid containing 5 g. of ferric chlorid per liter and a few drops of the formaldehyde solution. Heat the mixture slowly with occasional stirring until nearly boiling. The presence of formaldehyde is indicated by a violet coloration, varying in depth with the amount present. If no formaldehyde is present, the solution slowly turns brown. If there is doubt regarding the color, a blank test should be made without the formaldehyde. A comparison of the two results will clear away the doubt.

III. **The Combustion Products of Alcohols.** Ignite a little methyl alcohol in a clean combustion spoon and lower it into a large clean dry bottle. Keep the mouth of the jar well covered. Note what collects on the inside walls of the jar (?). Remove the spoon and add about 30 cc. of limewater; cover the jar and shake it (?). **Keep the alcohol in bottles away from flames.**

Questions: 1. Copper oxidizes rapidly in the air when heated. Write the equation.

2. The methyl alcohol, CH_3OH, is oxidized to formaldehyde, HCHO, by the copper oxid. Write this equation.

3. Summarize the resorcin test for formaldehyde.

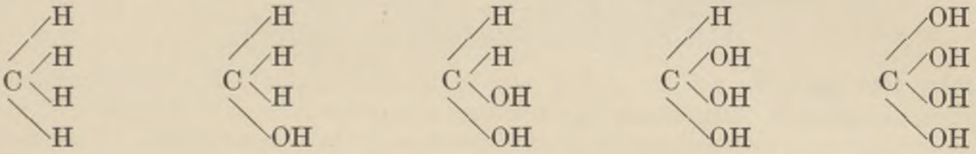
4. Summarize the casein test for formaldehyde.

5. Write the combustion equation for methyl alcohol.

THE RELATION OF SOME CARBON COMPOUNDS

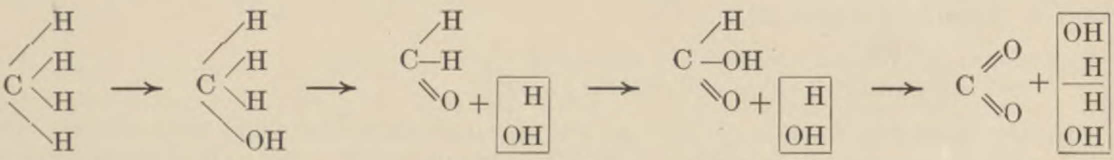
Carbon atoms form the nucleus or backbone of a great many compounds which make up the structures of living organisms (plants and animals). These atoms are able to unite chemically with each other in chains or groups and with other atoms to form compounds of great complexity. If the positions of the atoms in a molecule of one of these compounds are changed, a new compound is the result. The formula $C_{13}H_{28}$ can be expressed in 799 different arrangements of the atoms in the molecule, all of which are possibilities as substances having different properties. Only a few of these are actually known to exist, but the possibility of different arrangements of atoms in the molecule is another reason for the great number of carbon compounds. There are far more compounds of carbon than there are of any other element.

The final product of the complete oxidation of carbon is carbon dioxid. If this compound is completely reduced the final product is the hydrocarbon, methane (CH_4). Between these extremes of the oxidation and reduction action are a number of interesting and important derivatives. The oxidation of methane may be considered to take place by the following stages:



It has been found that the above compounds having two or more hydroxyl groups decompose, yielding water as one of the products.

The oxidation products then are:



Methane — Methyl alcohol — Formic aldehyde — Formic acid — Carbon dioxid

The classes of compounds, alcohols, aldehydes, and acids are found to exist for other series of hydrocarbons. While it is possible to carry out the above oxidation changes of methane, the method is not always used if an easier one is known. The relationship of these compounds as shown above makes it easy to understand and remember these compounds and their properties. The reactions by which methane can be converted into methyl alcohol are given on the next page. Methyl alcohol has been oxidized to formaldehyde in a previous experiment. In making vinegar the ferments found in the atmosphere change the sugars in fruit juices to alcohol, and then the alcohol into acetic acid.

If an alcohol and an organic acid are mixed with concentrated sulfuric acid to extract the elements of water, a compound corresponding to the inorganic salt is formed. It is called an **ester**, or **ethereal** salt.

$$C_2H_5OH + H(CO_2CH_3) \longrightarrow \underset{\text{(ethyl acetate)}}{C_2H_5(CO_2CH_3)} + H_2O.$$

From this action the alcohol may be considered an organic base, since it neutralizes acids to form a salt. Alcohols neither ionize nor change the color of litmus, however.

Other dehydration products that have been prepared are ethylene, by the action of phosphoric acid on ethyl alcohol, and carbon monoxid, by the dehydration of formic acid.

EXPERIMENT

Purpose: State it on a separate sheet for record.
Materials: Ethyl, amyl, and methyl alcohols, butyric, salicylic, and valeric acids, potassium dichromate, ether.

I. **Substitution Products of Hydrocarbons.** Chlorin or bromin gas reacts readily with the gaseous members of the paraffin series at ordinary temperatures to form a mono-substitution product, such as CH_3Cl or C_2H_5Br. Other members react similarly at higher temperatures.

$$CH_4 + Cl_2 \longrightarrow CH_3Cl + HCl.$$
$$C_2H_6 + Br_2 \longrightarrow C_2H_5Br + HBr.$$

When these substitution products are heated with a strong base like potassium hydroxid, the corresponding alcohol and a salt is formed.

$$CH_3Cl + KOH \longrightarrow CH_3OH + KCl.$$
$$C_2H_5Br + KOH \longrightarrow C_2H_5OH + KBr.$$

By these two reactions a hydrocarbon can be transformed into its corresponding alcohol. Ethyl chlorid, C_2H_5Cl, is a very volatile liquid, known under the trade name, Kelene. It is used by surgeons to produce local anesthesia by making a spot insensible to pain by cold. The tri-substitution products of methane, $CHCl_3$ and CHI_3, are the common drugs chloroform and iodoform.

II. **Preparation of Acetaldehyde.** To 6 cc. of water in a test-tube add 2 cc. of concentrated sulfuric acid and cool the solution in cold water. Next, add about 2 cc. of ethyl alcohol and pour the mixture on about 1 cc. of small crystals of potassium dichromate, $K_2Cr_2O_7$, in another tube (?). Hold the tube over the sink. The vapor escaping from the tube contains the acetaldehyde, CH_3CHO. Describe the odor (?).

Questions: 1. What is the oxidizing agent and what is oxidized?

2. What is the reducing agent and what is reduced?

III. **Preparation of Esters.** (*a*) To 1 cc. of acetic acid in a test-tube add twice as much ethyl alcohol and 2 cc. of concentrated sulfuric acid.. Heat the materials slowly to boiling temperature. Note the pleasant fruity odor of the ester, ethyl acetate (?).

$$C_2H_5OH + H(CO_2CH_3) \longrightarrow C_2H_5(CO_2CH_3) + H_2O$$

(*b*) To 1 cc. of acetic acid in a tube add twice as much amyl alcohol, $C_5H_{11}OH$, and 2 cc. of concentrated sulfuric acid. Heat mixture slowly to boiling temperature. Describe the familiar odor, name the ester formed, and write the equation (?). Omit sulfuric acid from the equation.

(*c*) Repeat the above procedure in (*b*) using the same quantities of butyric acid, $H(CO_2C_3H_7)$, ethyl alcohol, and sulfuric acid. Describe the familiar odor, name the ester, and write the equation (?). Note the remarkably different odors of the acid and ester.

(*d*) Repeat the above procedure, using methyl alcohol and salicylic acid, $H(CO_2C_6H_4OH)$ (?). Record as before.

(*e*) Repeat the procedure, using valeric acid, $H(CO_2C_4H_9)$, and amyl alcohol (?).

3. By what type of chemical action were the esters prepared?

4. Why are the esters sometimes called " ethereal salts " ?

IV. **Preparation of a Fatty Acid by Oxidation.** Note the odor (?) and test the liquid with litmus paper (?) of a dilute solution of glucose and yeast that has been fermenting for two weeks or more in an open vessel.

5. What acid was formed in Part IV?

6. What ferments caused this action to take place?

CARBOHYDRATES

This class of compounds is most essential to man as food, fuel, clothing, and shelter. The name signifies the presence of the elements carbon, hydrogen, and oxygen, the latter two in the proportion to form water.

The sugars are very valuable as a food because of their great solubility. Little time is required for their digestion, so they are assimilated at once. Through this property they are able to prevent or relieve fatigue quickly in cases of exhausting physical exercise. The starches make up over half of the food of the average man. They require considerable time for digestion. Cellulose is not digested in the human stomach. Sugars occur in roots, grasses, stems of plants, trunks of trees, leaves, and fruits, usually as sucrose, glucose, and fructose. Other kinds of sugars are known, but they occur in small quantities. See the per cent of sugars in common fruits, p. 247.

CARBOHYDRATES

I. **Sugars.** *A*. Those having **six** atoms of carbon per molecule (monosaccharids).
  Glucose, $C_6H_{12}O_6$ (grape sugar, dextrose).
  Fructose, $C_6H_{12}O_6$ (fruit sugar, levulose).
  Galactose, $C_6H_{12}O_6$
 B. Those having **twelve** atoms of carbon per molecule (disaccharids).
  Sucrose, $C_{12}H_{22}O_{11}$ (cane sugar, beet sugar).
  Maltose, $C_{12}H_{22}O_{11}$ (malt sugar).
  Lactose, $C_{12}H_{22}O_{11}$ (milk sugar).
II. Polysaccharids, those having $n \times$ **six** carbon atoms per molecule.
  Cellulose, $(C_6H_{10}O_5)_n$
  Starch, $(C_6H_{10}O_5)_{n'}$
  Dextrin, $(C_6H_{10}O_5)_{n''}$

Sucrose is the most used as a food. It is the sweetest of all the sugars. It is extracted from sugar cane and beets. Maltose is less important. It is formed from starch, along with dextrin, by the catalytic action of a ferment, diastase, found in malt (sprouted barley). This is a most important chemical change in plant life. When the plant matures, it stores up starch in the seed, particularly in the class of plants known as the cereals. This starch is to be the food supply of the new plant springing from the seed, until it can establish enough root to obtain its food from the soil. But starch is not soluble in cold water, so it cannot be used in this form. Nature has provided the ferment diastase, which under the conditions of warmth and moisture converts the starch into soluble maltose, that can be assimilated by the plant. Maltose is the sugar fermented in making liquors from grains. Lactose is the sugar in milk. It seems to be assimilated with the least difficulty, so it is used in the foods of infants and invalids.

Glucose is commonly prepared in the United States by heating the starch from corn under pressure with very dilute hydrochloric acid as a catalytic agent. Subsequently, the acid is neutralized with sodium carbonate to form sodium chlorid. This salt is usually present in American glucose. Equal quantities of it and fructose are formed from sucrose by the influence of acids, as follows:

$$C_{12}H_{22}O_{11} + H_2O \longrightarrow C_6H_{12}O_6 + C_6H_{12}O_6.$$

This action is called **inversion**. Similarly, glucose and galactose are formed from lactose. By inversion, maltose may be converted entirely into glucose. The sugars formed by inversion are frequently called **invert** sugars. The inversion process is not used commercially to make glucose, since the disaccharids are more expensive than the monosaccharids, while starch is very cheap.

Study Topic: Martin, *Modern Chemistry and its Wonders*: "Sugar," pp. 225–241.

EXPERIMENT

Purpose: State it on the sheet used for the record.
Apparatus: Water bath.
Materials: Solutions of silver nitrate, glucose, sodium carbonate, lactose, and maltose, Fehling's or Haines' solution, sucrose, cobalt chlorid paper, and various food products for testing.

I. **Reduction by Glucose.** Add extremely dilute ammonium hydroxid a drop at a time to 2 cc. of silver nitrate in a 15-cm. tube until the precipitate formed at first is dissolved. To this solution add a little glucose, a few cubic centimeters of potassium or sodium hydroxid solution, and warm the mixture. The glucose reduces the silver nitrate to metallic silver. Describe the evidence of this reducing action (?).

II. **The Test for Glucose.** The reducing power of glucose, as shown above, is used to convert the cupric hydroxid in the reagent, Fehling's or Haines' solution, to cuprous oxid. A remarkable change of color of the solution takes place during the test. Haines' solution is more convenient and does not deteriorate appreciably with age. To 5 cc. of a solution of glucose add an equal quantity of Haines' solution and heat the mixture to boiling. Give the color changes up to the final red-brown.

III. **Inversion of Cane Sugar.** (*a*) Make Haines' reduction test as above, using sucrose (cane sugar) instead of the glucose. Note carefully the difference in appearance of the mixture as compared with the test for glucose (?).

(*b*) To about 10 cc. of a fresh solution of sucrose add 5 drops of hydrochloric acid and heat the mixture in a test-tube in the water bath for about 10 minutes. (While waiting perform Part IV.) Neutralize the acid in the solution by adding sodium carbonate solution slowly until no further effervescence is noticed. Make the test for glucose with the sugar solution. Describe the result (?). Most other acids will bring about the same chemical action as hydrochloric. All monosaccharid sugars have the reducing power.

IV. **Reducing Power of Lactose and Maltose.** Test lactose and maltose, if available, for reducing power without inversion. State colors shown in the tests (?).

Questions: 1. Define the chemical process of reduction as shown in Part I.

2. State the test for glucose.

3. Is sucrose a reducing sugar? Give the proof.

4. Define the inversion of sucrose.

5. Sucrose is sweeter than glucose by the ratio 5/3. When stewed fruit is sweetened, why should the sucrose be added near the end rather than at the beginning of the stewing process?

V. **Commercial Forms of Sucrose.** (Optional.) Heat very slowly about a teaspoonful of moist sugar in the evaporating dish until the sugar melts. While it is yellowish-brown take some of it out on a paper to cool. Taste it (?). It is called "barley sugar" and is used in confectionery. Continue to heat the remainder **slowly** until it becomes dark brown. Do not burn it. Taste it (?). At this stage it is known as "caramel" and is used to color and flavor foods and beverages.

VI. **Constituents of Sugar.** Heat half a teaspoonful of dry sugar in a test-tube held horizontally in the flame. Note the changes in appearance and the odor (?). Test the colorless liquid condensing at the mouth of the tube with cobalt chlorid paper (?). Identify the liquid and the black substance in the tube (?). State three elements that are here shown to be constituents of sugar (?).

VII. **Testing Foods for Glucose.** Test any or all of the following substances in solution for glucose by the reduction test: Honey, maple syrup, molasses, jam, jelly, pie and cake filler, beets, carrots, parsnips, juices of grapes, cherries, apples, and pineapple. Record results of the test with the name or brand of food.

Glucose is a wholesome and cheap food. It is wrongfully used in foods as a substitute for sucrose. When used in this way, it is classed as an adulterant. Be sure to read the labels on all food products before purchasing, in order to know the ingredients.

CELLULOSE

This substance makes up the walls of plant cells. It serves as a framework for the plant in the same way that the bones do for the animal body. It is most abundant in the stems, less abundant in the leaves, and there is little in the fruit of plants. It is built into the stems where strength is needed to keep the plant in a suitable position for growth. As the plant matures, the cellulose changes into ligno-cellulose, which is very tough and durable.

Since cellulose is practically indigestible in the human stomach, it has no food value. Foods containing much cellulose, like asparagus, celery, and roots, are eaten to obtain the juices, which serve to regulate the physiological processes of the body. The cellulose itself acts as a mechanical aid to digestion by giving bulk to the food and making it porous for the entrance of the digestive juices. Its rough fibers stimulate the walls of the intestines, causing peristaltic motion.

The greatest value of cellulose to man is the use of its tough fibers for clothing and shelter. Paper is entirely cellulose, except for the materials used in giving it a smooth finish. This cheap article has brought knowledge on the printed page within the reach of persons of the most limited means. As wrappings for merchandise it serves to protect materials of much greater value than itself. In these and other ways paper has become an indispensable article of the modern world. As important as it is, paper that is used for wrappings would fail in its purpose if not held in place by cellulose in the form of cord and twine.

The cellulose known as cotton is almost ideal for clothing. It grows on the plant in the boll, with the fibers already separated, so it needs only a little mechanical treatment to fit it for spinning and weaving. Hence it is cheap. It is a good conductor of heat, so, as clothing, it allows the body to cool off readily in summer. If it is woven with a " nap " or " fleece," it has air spaces between the fibers, which lessens the conductivity of heat enough to make it a comfortable winter clothing. Since its high capillary capacity makes it very absorbent, it readily takes up the secretions of the skin and promotes body hygiene. Without this property, a towel is of little use. Finally, cotton is easily washed, keeps its shape, and is very durable. The cellulose known as linen has a higher capillary capacity and greater conductivity of heat, but it is more expensive, owing to the greater labor of preparing the fibers for use. The gums and resins which bind the fibers together in the flax stem make much chemical, as well as mechanical, treatment necessary before the fibers can be spun into thread.

Ingenious minds have found ways to alter the properties of cellulose for special uses. Moderately concentrated sulfuric acid will change soft absorbent paper into tough, durable parchment paper. If cellulose is treated with nitric acid, it becomes nitro-cellulose. The lower nitrated product is dissolved in ether and alcohol to make collodion for lacquers and photographic plates; or mixed with camphor it becomes celluloid. The higher nitrated product is the terrible explosive, gun cotton.

John Mercer discovered that cotton cloth became 50 % stronger and shrunk 20 % when treated with a concentrated solution of sodium hydroxid. It then has a greater dyeing capacity. If the shrinking is not permitted, by keeping the cloth stretched on a frame while under treatment, the surface of the fabric takes on a silken sheen like real silk. It is called mercerized cotton.

In spite of the extensive study of cellulose and its adaptation to so many uses by the chemist, he has been unable to learn more about the composition of cellulose than that it is composed of hydrogen, oxygen, and carbon in a definite proportion. The complexity of the molecule makes it still a mystery. The future holds prizes for the careful student who finds new adaptations and uses for cellulose.

Study Topics. Duncan, *The Chemistry of Commerce*, pp. 219–240.
Bird, *Modern Science Reader:* "Artificial Silk," pp. 36–42.
Findlay, *Chemistry in the Service of Man:* "Cellulose and Cellulose Products," pp. 78–87.
Sadtler, *Chemistry of Familiar Things:* "Cellulose," pp. 273–280.

CELLULOSE AND STARCH

EXPERIMENT

Object: State it on a separate sheet used for the record.
Materials: Filter paper, sodium hydroxid, sp. gr. 1.33, iodin solution, starch, sodium carbonate solution, paraffin, Haines' solution, starch from wheat, corn, potato, and arrowroot, white muslin.

Cellulose

I. **Solvents.** Measure out 10 cc. of water into the evaporating dish. Add slowly to the water 15 cc. of concentrated sulfuric acid. **Be cautious**! Pour about 20 cc. of the mixture into a tube and stand the tube with contents in a jar of cold water to cool. Pour the remainder into another tube and put 2 bits of filter paper in the acid. Warm the contents of the tube, while shaking gently. Describe the result (?). Other solvents of cellulose are Schweitzer's reagent and acid zinc chlorid.

II. **Parchment Paper.** Pour the cold solution prepared in (*a*) into the evaporating dish. Immerse one end of a strip of filter paper in the acid for 5 seconds, rinse the paper in water, then let it lie in a solution of ammonium hydroxid for 5 minutes, and finally rinse in water and dry it between blotters. Describe the change in appearance, texture, and toughness (?). Mount uniformly trimmed pieces (3 × 2 cm.) of the treated and untreated paper. Arrange and label neatly.

III. **Mercerized Cotton.** Place a small piece (4 × 6 cm.) of fine, white, unsized muslin in a solution of 30 % sodium hydroxid, specific gravity 1.33. Leave it in the solution for 10 minutes, then wash and dry the cloth. Describe the change in appearance, texture, and toughness (?). If the cloth can be held tightly in a frame to prevent shrinkage while being treated, the change will be more pronounced. When the cloth is dry, mount neatly trimmed pieces of the mercerized and unmercerized cotton.

Starch

IV. **Iodin Test for Starch.** Shake a small pinch of starch with half a tubeful of cold water, and add **one drop** of iodin solution. Color (?). Repeat the test, but first boil the starch and cool the solution. Color (?). Heat the starch solution until the color is gone, then cool it in cold water. Note the return of color. Heat the solution again and boil it gently for a minute. Cool the solution and note that the color does not return as before. Now add a drop of iodin solution (?). The best test for starch is the dark-blue color with iodin. Make a careful statement of the conditions under which this color will appear (?).

V. **Conversion of Starch into Glucose.** Start heating the water bath. Add a large pinch of starch to half a tubeful (15 cm.) of water and heat it to boiling. Add 2 drops of concentrated hydrochloric acid and keep the tube in the gently boiling water of the bath for 10 minutes. Add sodium carbonate solution until no further effervescence shows. Test the solution for starch and glucose (?). Write the equation for the conversion of starch into glucose (?). See formulas and explanation on a preceding page. A chemical action of this nature is called **hydrolysis**, because one of the reagents is water.

VI. **Digestion of Starch with Saliva.** Chew a piece of paraffin to promote the flow of saliva. Collect about 5 cc. of it in a test tube and add **one drop** of iodin solution. Add a **few drops** of a boiled starch solution (?). Let the tube stand for 15 minutes or longer (?). When the color is gone, make a test of the mixture for glucose (?). Explain the gradual fading of the color (?). What procedure is advisable when eating starchy foods?

VII. **Microscopic Examination of Starch.** (Optional.) Examine the starch grains from corn, wheat, potato, and arrowroot. Make diagrams of the relative sizes of grains and their markings.

SOAP

It has been said that the state of civilization of a nation or people can be determined by the quantity of soap used. The length and efficiency of human lives are certainly largely controlled by cleanly and sanitary habits of living. To this end no material contributes more service than soap. It was first used as a cosmetic and medicament, but it soon found its proper place in human economy as a necessary cleansing agent or detergent.

In the year 1000, Marseilles had a flourishing soap industry, but little was known of the chemical nature of soap until the nineteenth century, when a French chemist, Chevreul, determined its composition. This knowledge gave the industry a scientific basis. One of the two chief ingredients of soap is an oil or a fat, which is composed of an organic base, glycerin, $C_3H_5(OH)_3$, and a mixture of various organic acids depending on the source of the fat. The three chief ones found as constituents of fat are palmitic, stearic, and oleic acids. From this it may be seen that a fat is an organic salt or ester. The other ingredient of soap is a strong base called an alkali. Sodium hydroxid is the one commonly used. The action belongs to that class in which a base acts on a salt. Another salt, soap, and a base, glycerin, are the products.

In pioneer days the housewives saved the fats that were not suitable for food until the time of the annual event of soap making. The fats were boiled with lye, extracted by water from wood ashes, or the commercially prepared lye was purchased. The lye from wood ashes is potassium carbonate (potash), which produces potassium hydroxid in a water solution. The soap made from this lye is soft, while that made from soda is hard. The soda is cheaper also, so it is practically the only base now used in soap making. Successful soap making in the home is largely a matter of experience because of the variable proportions of the different fats used. An excess of the alkali was generally used so that the soap would not be greasy. It was effective in removing dirt and at the same time injurious to the clothing and the skin. In factories where the fat stock is carefully tested and is more uniform in composition, scientific methods are used to produce soaps of excellent quality. Cocoanut oil is one of the easiest fats to convert into soap. A temperature of 80° C. is sufficient to complete the reaction, if considerable time is allowed.

Kerosene is used as a spray to kill insects on fruit trees. When diluted with water, the kerosene is not dissolved but separates from the water soon after being mixed with it. If a soap solution is shaken with the oil and water, the oil remains distributed through the mixture in small droplets, forming an **emulsion**. The soap is said to " carry the oil." In a similar way soap forms an emulsion with the fats or oils that cause the dirt to adhere to the fibers of clothing. This action loosens the hold of dirt, thus permitting its easy removal by mechanical washers.

Study Topic: Sadtler, *Chemistry of Familiar Things:* "Soap," pp. 262–266.

SOAP

EXPERIMENT

Object: State it on a separate sheet used for the record.

Materials: Cocoanut oil, saturated solution of sodium hydroxid, calcium chlorid solution, kerosene, waste fat or grease, commercial lye.

I. **Preparation.** Measure out in a graduate 8 cc. of a cold saturated solution of sodium hydroxid. Pour into this solution 10 cc. of melted cocoanut oil. Empty the mixture into an evaporating dish and heat it with a small flame, not touching the dish, until the mixture becomes thick. Stir the soap to prevent spattering. The oil is a mixture of the glyceryl salts, of palmitic, stearic, and oleic acids. The sodium salts of these monobasic acids constitute the soap. The glycerin remains mixed with the soap. Consider the oil as glyceryl palmitate, $C_3H_5(CO_2C_{15}H_{31})_3$, and write the equation.

II. **Properties.** (*a*) Take a little of the soap and wash the hands with it. Note how readily it lathers (?). If it makes the hands rough, too much alkali, sodium hydroxid, was used; if greasy, too much oil. From these observations state the quality of the soap formed.

(*b*) **Formation of Fatty Acid.** Using about half of the prepared soap, make a solution in the flask, using 100 cc. of warm distilled water. Fill a test-tube half full of the solution and add about 5 cc. of dilute hydrochloric acid. Shake the mixture well. The fatty acid is the precipitate formed. Recall a general method of preparing an acid and write the equation. Separate most of the acid and mix it with water in another tube. Add a little sodium hydroxid solution and heat the mixture. Note whether soap forms again or not (?). Write the equation.

(*c*) **Effect of Hard Water.** To another tube half full of soap solution add half as much calcium chlorid solution and shake it (?). An insoluble calcium soap is formed. Write the equation. Calcium compounds (lime) are frequently present in natural waters, making them "hard." The insoluble soap is useless for cleaning and makes sinks and wash basins grimy.

(*d*) **Cleansing with Soap.** Put 1 cc. of kerosene in a test-tube, add 10 cc. of water, and shake. Note the separation of the oil into tiny globules (?). Such a mixture is an **emulsion.** Let it remain quietly for a minute (?). Is the emulsion permanent? Add about 2 cc. of the soap solution to the mixture and shake again. Note permanency of the emulsion (?).

Procure 300 cc. of hot water and heat the soap solution to boiling. Dip a dry test-tube brush into some oil and rub it in some dirt. Move the brush back and forth in half a tubeful of hot water. Does the hot water clean the brush? Use the hot soap solution in a similar manner (?). Pour the used soap solution into a beaker and rinse the brush with hot water, catching rinsings in the beaker. When the contents of the beaker settles, determine the condition of the oil and the dirt. Explain how soap cleanses (?).

III. **Soap Made from Waste Fat.** (Optional.) The fat or grease should be melted and strained through cheese-cloth to remove any solid particles. No matter how rancid or dirty it is, it will serve. Weigh the fat. Then weigh out one-third as much commercial lye and stir it into sufficient water to dissolve it. Heat the fat in an iron kettle and pour the lye solution slowly into it while stirring. Be careful to prevent the hot solution from spattering or boiling over. If the whole mixture becomes thick and pasty, add water to make it thin. After boiling the soap for 30 minutes or 1 hour, add twice as much sodium chlorid as lye and stir it in well. Allow soap to cool. Remove the layer of soap on top and boil it again with considerable water and add salt as before, if a whiter soap is desired. Otherwise pack the hot soap into small pasteboard boxes to form cakes of the desired size. The soap will be as good as the average laundry soap.

SULFUR

Almost all substances can be obtained in the three states of solid, liquid, and gas, each change occurring at a definite temperature. Many substances, however, both elementary and compound, have the property of existing in two or more solid or liquid or gaseous states. Thus ozone is another gaseous form of oxygen; diamond, graphite, and charcoal are three solid forms of carbon; and sulfur exists in two solid, two liquid, and three vapor forms. The two solid forms differ in certain physical properties, and only one of them, rhombic, is stable at ordinary temperatures. At 96° C. it changes into the other solid form, monoclinic, which is the stable form of sulfur between that temperature and its melting point, 119.25° C. Slow and careful heating of either solid form will give one of the liquid varieties, S_λ, but too high a temperature changes this into the other liquid variety, S_μ, which is quite different in many respects from S_λ. The S_μ boils at 445° C., giving a vapor having the formula S_8. As the temperature of the vapor is raised, the molecules of S_8 split up into S_6 and finally S_2. These changes may be represented diagrammatically as follows:

$$\underbrace{S_{\text{rhom.}} \overset{96^\circ}{\rightleftarrows} S_{\text{mono.}}}_{\text{Solids}} \overset{119^\circ}{\rightleftarrows} \underbrace{S_\lambda \overset{160^\circ}{\rightleftarrows} S_\mu}_{\text{Liquids}} \overset{445^\circ}{\rightleftarrows} \underbrace{S_8 \rightleftarrows S_6 \overset{1000^\circ}{\rightleftarrows} S_2}_{\text{Vapors}}.$$

Sulfur belongs to the oxygen group of elements and there are many points of resemblance between its compounds and those of oxygen, selenium, and tellurium. It will be noted that the properties of the four elements exhibit quite regular variations in the same order as the atomic weights.

OXYGEN–SULFUR GROUP

	OXYGEN	SULFUR	SELENIUM	TELLURIUM
Atomic weight	16	32	79	127.5
Valences	− 2	− 2, 4, 6	− 2, 4, 6	− 2, 4, 6
State	gas	solid	solid	solid
Melting point	− 227° C.	119° C.	217° C.	452° C.
Boiling point	− 183° C.	445° C.	680° C.	1390° C.
Hydrogen compound	H_2O	H_2S	H_2Se	H_2Te
Stability	very	less	less	least
Oxids	OO_2 (ozone)	SO_2 SO_3	SeO_2 SeO_3	TeO_2 TeO_3
Acids		H_2SO_3 H_2SO_4	H_2SeO_3 H_2SeO_4	H_2TeO_3 H_2TeO_4
Strength of acids		greatest	less	least
Chlorids	OCl_2	SCl_2	$SeCl_2$	$TeCl_2$
Stability	least	more	more	most
Affinity for metals	most	less	less	least

SULFUR

EXPERIMENT

Purpose: State it on a separate sheet used for the record.
Materials: Sulfur, carbon disulfid, copper foil, iron powder.

I. **Allotropic Forms.** (*a*) Rhombic. Shake a pinch of powdered sulfur with 2 cc. (no more) of carbon disulfid in a clean dry test-tube (?). Pour the liquid into a watch glass and allow it to evaporate (keep it 6 feet from flames; **inflammable**). Examine the crystals, using a lens if they are very small, draw several, and record the color and other properties. Examine them again after 24 hours.

(*b*) Monoclinic. Place a folded filter paper in the funnel, fill a beaker with water, and half fill a clean dry 15-cm. test-tube with powdered sulfur. Melt the sulfur slowly by holding the test-tube well above the flame and turning the tube so that the sulfur will not be overheated in any place. If properly done, the liquid is a pale yellow (S_λ); a dark color is due to overheating, producing S_μ.

Pour the liquid sulfur into the filter paper, and as soon as the crystals extend from the sides to the center, break the crust and pour the sulfur which is still liquid into a beaker of water and immediately open up the paper. Examine the crystals, make a drawing of them, and record their color, transparency, and any change in appearance after 24 hours. Dry the sulfur in the beaker and use it in the following part.

(*c*) Amorphous or Plastic Sulfur. Again melt half a test-tubeful of sulfur (use the same tube), and continue the heating slowly and carefully till the sulfur boils. Note and record the changes in color and viscosity (tilt the tube from time to time). Pour the boiling sulfur into a beaker of water. Note what happens where the fumes strike the surface of the water, and describe the product (flowers of sulfur). Examine the sulfur which was suddenly cooled below the water, recording its color, hardness, brittleness, and solubility in carbon disulfid. After 24 hours, examine it again and record the changes noted in any of these properties.

Questions: 1. Which of the forms of sulfur that you have studied is unchanged by standing 24 hours, *i.e.* is most stable at the ordinary temperature?

2. Into which form do the others change on standing?

3. Under what conditions is each form stable?

II. **Chemical Properties.** (*a*) Recall the burning of sulfur, describing the flame; give name and properties of the product, and write the equation.

(*b*) Recall the effect of heating together powdered sulfur and zinc dust, describe the action, name the product, and write the equation.

(*c*) Boil about 2 cc. of sulfur in a test-tube (the tube used above may be used here; fasten it in a clamp on the stand) and drop a strip of copper foil into the tube (?). When the action is finished, remove the copper strip and examine the product. Name and describe it, and write the equation.

(*d*) Mix thoroughly 5 g. of iron powder and 3 g. of flowers of sulfur, place the mixture in the same test-tube and heat the rounded part of the tube red hot. As soon as the action is well started, remove the tube from the flame. When the tube is cool, break it and examine the product. Describe the action, and record the name and properties of the product. Equation.

Questions: 4. State two chemical properties of sulfur.

HYDROGEN SULFID

As a laboratory reagent, the chemist finds hydrogen sulfid very useful in analysis to determine what metals are present in a material of unknown composition. All the common metals unite with sulfur to form sulfids. The solubilities of these sulfids differ in such a way that it is possible to separate the metals into groups for final isolation and identification. Some of these sulfids are black, some yellow, some brown, one is pink, another orange, and another white; so these different colors make the identification of these metals easy and certain.

Hydrogen is used as a standard element in several ways. Other elements are compared with it in density, atomic weight, and valence. The nature of the other elements is better understood and more easily remembered if there is a common basis of relationship established. Some interesting relations are shown by a study of the hydrogen compounds, known as hydrids. Hydrogen sulfid is one of the most important members of this class of compounds. Hydrogen always forms a positive ion in solution, so it may be expected to combine readily with non-metallic ions. The hydrids of the halogen group of elements are strong acids, hydrofluoric, H_2F_2, hydrochloric, HCl, hydrobromic, HBr, and hydriodic, HI. It will be noticed that these elements are monovalent. The hydrids of bivalent non-metals, oxygen, sulfur, selenium, and tellurium, are weak acids, excepting that of oxygen (water), which is neutral.

The hydrids of the trivalent non-metals, nitrogen, phosphorus, arsenic, and antimony, have no acidic properties. Ammonia, NH_3, is decidedly alkaline, and the others seem to be neutral. Bismuth, the remaining element of this group, does not form a hydrid, presumably because of its metallic nature.

Methane, CH_4, and silicane, SiH_4, are the only hydrids of tetravalent elements, and do not affect litmus. The only other non-metallic element, boron, forms a hydrid, BH, which is inactive. As a rule, hydrogen does not unite with metals. The most active metals, the alkalies and alkaline earths, however, do form hydrids. This exception may be explained by considering that the hydrogen functions as a negative non-metal when combining with these very active positive elements. The less active metals do not form hydrids.

The chemical properties of the hydrids of selenium and tellurium are remarkably similar to those of hydrogen sulfid. Since these elements are not abundant, and the hydrids unstable, they cannot be used in place of hydrogen sulfid. These elements are little used at the present time.

COMPARISON OF HYDRIDS OF OXYGEN–SULFUR GROUP

Hydrids	Mol. Wt.	Melt. Pt.	Sp. Gr.	Odor	Dissociation Temp.	Reaction to Litmus	Boil. Pt.
H_2O	18.02	0° C.	1.0	none	1800° C.	neutral	100° C.
H_2S	34.09	– 85.5° C.	1.17	"ripe" eggs	400° C.	weak acid	– 61.8° C.
H_2Se	81.22	– 64° C.	2.81	fetid	150° C.	weaker acid	– 42° C.
H_2Te	129.52	– 48° C.	4.48	fetid	0° C.	weakest acid	0° C.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Materials: Ferrous sulfid, litmus paper, solutions of arsenious chlorid, antimony trichlorid, copper sulfate, lead nitrate, lead acetate, chlorin, and bromin, copper foil, paper.

I. **Preparation**. Put the equivalent of about one level teaspoonful of iron sulfid, FeS, in the generator and insert the stopper with the thistle tube (extending to the bottom of the bottle) and the delivery tube as shown in Figure 85. Pour in enough dilute hydrochloric acid to cover the iron sulfid and to close the end of the thistle tube. Note the action and write the equation of the action (valence of the iron is + 2). Fill 2 clean dry jars with the gas by " upward displacement of air " (jar upright), keeping them covered with a perforated cardboard during the filling and with a glass plate afterwards. Make a diagram of the apparatus.

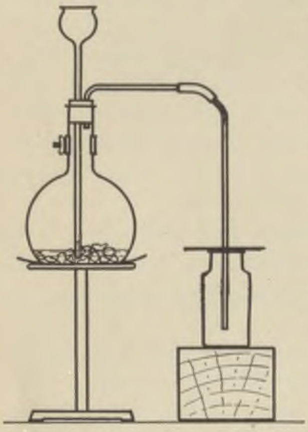


FIG. 85. Hydrogen sulfid from iron sulfid.

II. **Properties**. (*a*) Note the color (?) and odor (?) of the gas, but do not inhale much of it, as it is somewhat poisonous.

(*b*) Light the gas at the end of the delivery tube and note the odor of the burning gas (?). Hold a cold, dry jar over the flame and note whether any product of the combustion is condensed thereon (?). What are the products of the complete combustion of hydrogen sulfid? Write the equation.

(*c*) Extinguish the flame and insert the delivery tube to the bottom of a 15-cm. test-tube three-fourths full of distilled water. Allow the bubbles to rise through the water for several minutes. Then smell the liquid to find out whether any of the gas dissolved (?). Save the solution.

(*d*) Bring a lighted splinter to the mouth of a jar filled with the gas (?) and then insert it into the jar (?). Note whether the gas burns (?) or supports combustion (?). Smell the contents of the jar cautiously (?), and then examine the walls of the jar for two other products of the action (?). Explain the formation of the third product (compare with the incomplete combustion of carbon compounds), and write an equation for the action.

(*e*) Test the solution of the gas from Part (*c*) with both colors of litmus paper (?). Make up another name for this compound to express the fact learned from the litmus test, and in accordance with the rules previously given for naming binary compounds of this class.

(*f*) Tests. The odor may be used as a test when rather large quantities of the gas are present. The best test is the following: dip a piece of filter paper into a solution of lead acetate and hold it in the gas to be tested, the second jar (?). Name the product formed and write the equation.

(*g*) Pour about 2 cc. of each of the following solutions into separate test-tubes, and add to each an equal volume of the solution of the gas: copper sulfate, $CuSO_4$; arsenious chlorid, $AsCl_3$; antimony trichlorid, $SbCl_3$ (?). Describe and name the precipitates formed. State what variety of chemical change is here illustrated. Write the equations of the four actions.

(*h*) Dip a bit of copper foil and a silver coin into the gas solution (?). Name the products and write the equations. State the variety of chemical change illustrated.

(*i*) Add 1 cc. of chlorin water to 5 cc. of the hydrogen sulfid solution (?). Interpret the result, recalling the power of chlorin to take hydrogen from its compounds and form hydrogen chlorid with it. Write the equation for the action. The formation of the product here noted is always a sign that the hydrogen sulfid is acting as a reducing agent. The ease with which it gives hydrogen makes the compound a powerful reducing agent, for this hydrogen can then combine with any negative or non-metallic radical.

(*j*) Give three physical and four chemical properties of the gas and its solution.

SULFUR DIOXID AND SULFUROUS ACID

The term " sulfurous acid " is used in the commercial world to include its anhydrid, sulfur dioxid, and its salts, the sulfites. These substances have a very important bearing upon the public health, as they are more generally used in the preparation of food products than any other preservatives. Their purpose is three-fold : to destroy germs that cause unfavorable chemical changes during the process of manufacture; to bleach the food to a more pleasing color; and to protect the food from decomposition by the entrance of germs.

Before wine is put in the barrel for aging, a cloth dipped in molten sulfur is burned in the barrel to destroy yeast and ferments in the wood which would cause undesirable changes in the wine. The interiors of the barrels are moist and frequently the wine is run in soon after the fumigating process. Since the sulfur dioxid is soluble, considerable sulfurous acid may be found in the wine.

Much attention has been given in the past twenty years to the preparation of foods of **pleasing appearance**. In some cases, the effect has been obtained with a loss in food value of the product, and in others, impurities are introduced that are positively detrimental to health. Dried fruits darken during the drying process, but if they are fumigated with sulfur dioxid, they retain their light color or become yellow. Their attractive color makes them more desirable to the consumer who may not know how this appearance is obtained. Fruits long exposed to the sulfur dioxid may be left quite juicy — even as much as 30 % of moisture — without danger of decay. Manufacturers are thus tempted to use sulfur dioxid freely in order to make a greater profit. Wines, grape juice, and syrups are also bleached by the same method with the inevitable absorption of the sulfur dioxid fumes.

It seems to be the custom to display dried fruits and sausage meats for sale with little attempt to protect them from the organisms always floating about in the air. Such foods are usually treated with sulfur dioxid to make them less susceptible to the attack of molds and fungi.

For some or all of the three above reasons, manufacturers have sought to justify the use of this preservative. Government experts have found, after long investigation and mature study, that, while the use of food containing sulfites on a few occasions may produce no ill effect on bodily health, yet the **continued** use of even small quantities of sulfites will result in serious disturbances to digestion and interference with the functions of the body. A lowered vitality, a lessened efficiency of work, and a shortened life would be the logical result of the continued use of this drug.

During the last few years, in some cases where the use of sulfur dioxid has previously been considered a necessity, foods have been prepared without its use, and have been found to have the desirable qualities of the drugged product, and to keep fully as well. Other means of sterilizing foods have been found for use in the process of manufacture. Sanitary handling and protection from germs have been found to be better preservatives than any drug. Laws forbidding the use of this preservative in foods are yet to be enacted. A vigorous public sentiment must be aroused to compel the legislators and manufacturers to adopt sanitary and healthful methods for the preparation of foods.

There are no delicate tests for the detection of sulfurous acid or its salts. The usual procedure in such cases is to change the substance into something else which can be found by tests even when present in small amounts. Oxidation and reduction actions are quite commonly employed in such changes.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Materials: Sodium bisulfite, calico, pink flower (carnation), apple, hydrogen peroxid, solutions of barium chlorid and lead acetate. One or more of the following: Hamburg steak, sausage, mushrooms, catsups or fruit juices, white wines, fruit and vegetable pulp, dried fruits.

I. **Preparation**. (Two pupils.) Place about 10 g. of sodium bisulfite in the flask, add 5 cc. of water, insert the stopper, and clamp the apparatus securely (Fig. 86). Pour about 10 cc. of concentrated commercial sulfuric acid into the dropping funnel and adjust the latter so that the acid will drip slowly upon the sodium bisulfite. Note the action and write the equation (?). As soon as the air is out of the flask, collect two jars of the gas by upward displacement of air, covering jars promptly. While the second jar is being filled, boil 15 cc. of distilled water for several minutes to remove dissolved air, cool under water, and pour it into a beaker. Fill a test-tube with the gas in the same way and cover it tightly with the thumb. Attach a calcium chlorid tube to the end of the delivery tube and bubble the gas through the water in the beaker for five minutes.

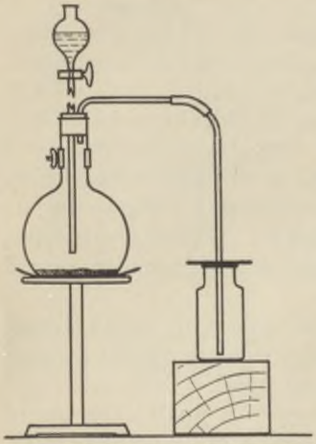


FIG. 86. Sulfur dioxid from a sulfite.

II. **Properties of Sulfur Dioxid**. (*a*) Uncover the test-tube just enough to run in a little water, close it again tightly with the thumb, and shake. Open the mouth of the tube under the water in the pan, note and explain any change in the level of water in the tube.

(*b*) Use one jar of the gas to determine its combustibility and power to support combustion (?).

(*c*) In the second jar place a piece of moistened calico, a pink flower (carnation), and a thin slice of apple (?). Also expose a slice of apple to the air outside the laboratory. Compare the pieces.

III. **Properties of Sulfurous Acid**. (*a*) Test the solution made in Part I with litmus papers (?). Write the equation for the formation of this product, H_2SO_3.

(*b*) Boil 5 cc. of the solution slowly in a dish, noting the odor from time to time (?). State the degree of stability of the substance, and correct the equation of the preceding part so that it will also represent the change occurring in this.

(*c*) Place about 1 cc. of mossy zinc in a 15-cm. tube and cover with dilute hydrochloric acid. When the effervescence (?) is vigorous, lay a piece of moist lead acetate paper over the mouth of the tube for a minute (the paper should remain white). Now pour 2 cc. of the sulfurous acid solution into the tube and replace the lead acetate paper (?). Aided by the energy of the action and probably by the catalytic action of the zinc, hydrogen is a powerful reducing agent; *i.e.* it takes oxygen from many substances. Write three equations: (1) the formation of hydrogen, (2) the reduction of sulfurous acid to hydrogen sulfid, (3) the test for hydrogen sulfid.

(*d*) Divide the rest of the sulfurous acid solution into two parts, and add about 1 cc. of hydrogen peroxid solution to one. Now add about 2 cc. of barium chlorid solution (test for sulfate) to each tube, a little hydrochloric acid (?), and compare results. Hydrogen peroxid, H_2O_2, is very unstable, and readily gives half its oxygen to other substances. State the action of hydrogen peroxid on sulfurous acid. Write the equation for it and the test for a sulfate. State four physical properties of the gas and seven chemical properties of the gas and its solution.

IV. **Sulfurous Acid or Sulfites in Food Products**. Bring one or more of the above stated foods from home, and test each by the reduction method, Part III (*c*), and also by the oxidation method (*d*). Use 30–50 g. of the food (macerated if solid) instead of the sulfurous acid in the tests. If any test is doubtful add a little sodium sulfite and continue the test with a fresh test paper. Compare results (?). Attach test papers to analysis report sheet and fill out for a record.

REPORT OF CHEMICAL ANALYSIS

EXPERIMENT

Partial analysis of ....

Obtained from..

Manufactured by..

Exact name of brand ....

Tested for..

Description of test..

....

....

....

....

Equations of test..

....

....

Found..

....

Quality of sample..

Remarks..

....

....

I hereby certify that the above report is correct to the best of my knowledge.

(Signed)..

Date......................

SULFURIC ACID

This substance is probably of more importance to the general industrial world than any other single substance except water. Practically every industry uses either sulfuric acid or some substance prepared by means of sulfuric acid. Immense quantities of the acid are used every year in making " superphosphate " of lime for fertilizers, refining petroleum oil, manufacture of other acids, manufacture of explosives, and washing soda. It has been stated that the industrial importance of a nation can be measured by the quantity of sulfuric acid it uses.

The uses of sulfuric acid can be grouped around three properties, as follows: 1st, **Dehydrating Agent.** The acid combines vigorously with water to form two hydrates having the formulas $H_2SO_4{\cdot}H_2O$ and $H_2SO_4{\cdot}2\,H_2O$. Since the reaction is very vigorous, it is used to remove water from substances, or even to remove the elements hydrogen and oxygen from substances in proportions to form water from them. A notable example is the preparation of explosives like nitroglycerin and guncotton. When nitric acid acts on glycerin or cotton, water is produced, and this water would cause the action to be incomplete on account of a reverse action between the water and the other product. By adding concentrated sulfuric acid to the nitric acid the reverse action is prevented.

2d. **Acid Properties.** Sulfuric acid is the most important and useful of all the acids because of its cheapness, its high boiling point (338° C.), and because it is quite active. Its activity, of course, is due to the fact that it is quite highly ionized when in solution. It is the most active of all the polybasic acids, though not nearly so active as nitric and hydrochloric acids. Most of the ionization of the sulfuric acid is a separation into a positive hydrogen ion and a negative " acid sulfate " or " bisulfate " ion. To a slight extent, the " bisulfate " ion is further separated into hydrogen ion and sulfate ion. The entire ionization may be represented by one equation, thus:

$$H_2SO_4 \rightleftharpoons H^+ + (HSO_4)^- \rightleftharpoons H^+ + (SO_4)^{--}.$$

The concentrated acid is but very slightly ionized (very poor conductor of electricity), hence the properties are very largely those of the molecules of the acid. As water is added, the degree of ionization is increased, and the properties are more and more those of the ions present. As there are two different negative ions present, $(HSO_4)^-$ and $(SO_4)^{--}$, two kinds of salts can be made from sulfuric acid, the **acid sulfates** or **bisulfates** containing the $(HSO_4)^-$ ion, and the **sulfates** containing the $(SO_4)^{--}$ ion.

3d. **Decomposition by Heat.** When concentrated sulfuric acid is heated, white fumes of its anhydrid, sulfur trioxid, SO_3, are given off at 150° to 180° C., which is considerably below the boiling point of the acid. This decomposition increases as the temperature rises, and at the boiling point, one-third of the vapor consists of sulfur trioxid and water vapor. When the temperature is raised above 400° C., the sulfur trioxid decomposes into sulfur dioxid and oxygen. If some substance is present which can combine with the oxygen, the decomposition of the sulfur trioxid takes place at a lower temperature. For this reason the action of the hot concentrated acid on metals is quite different from that of the cold dilute acid. The metals are first changed to their oxids, and these are then changed to the sulfates by the action of additional sulfuric acid.

$$\begin{array}{r}
H_2SO_4 \longrightarrow H_2O + SO_2 + [\cancel{O}] \\
Cu + [\cancel{O}] \longrightarrow [\cancel{CuO}] \\
[\cancel{CuO}] + H_2SO_4 \longrightarrow CuSO_4 + H_2O \\
\hline
Cu + 2\,H_2SO_4 \longrightarrow CuSO_4 + SO_2 + 2\,H_2O
\end{array}$$

SULFURIC ACID

EXPERIMENT

Purpose: State it on a separate sheet used for the record.
Materials: Starch, sugar, barium chlorid solution, copper turnings.

I. **Dehydration.** (*a*) Examine the pure concentrated sulfuric acid and note its properties (?). (Be very careful not to get any on the skin or clothes, and taste only very dilute acid — a couple of drops in 100 cc. of water.)

(*b*) Pour 10 cc. of the concentrated acid into 10 cc. of distilled water in a beaker with stirring, and touch the beaker to the back of the hand (?). Pour this solution into 20 cc. of distilled water and save it for II (*a*) and (*b*).

(*c*) Pour a little concentrated sulfuric acid on a little powdered starch, sugar, and paper in separate test-tubes, warming gently if necessary to get a decisive result (?). Explain.

II. **Acid Properties.** (*a*) Measure 30 cc. of potassium hydroxid solution (10 %) into an evaporating dish. Rinse the graduate, and fill it to the 25-cc. mark with the dilute sulfuric acid solution prepared in I (*b*). Place a bit of litmus paper in the dish and then add the dilute acid, stirring constantly, till a permanent lavender color of the litmus paper shows that the base has been neutralized. At first the acid may be added in one cubic centimeter portions, but as the end of the action approaches (this will be shown by the litmus paper becoming pink and turning back to blue on stirring), the acid must be added a drop at a time till the end point is reached. A clean glass tube may be used as a dropper, holding a finger over the upper end to control the flow of the acid. Note the number of cubic centimeters of the acid solution used in neutralizing the base (?). If crystallization has not occurred on cooling, rub the inside of the dish with a glass rod (?). When crystals appear, examine them carefully, recording color, taste, effect on litmus, and the shape. Name the product and write the equation for its formation.

(*b*) Measure another 30-cc. portion of the same potassium hydroxid solution into another dish or beaker and add twice as much of the dilute acid solution as you used in II (*a*). Evaporate the liquid to half its volume and allow it to cool and crystallize as before. Examine the product carefully (?), comparing it with that obtained in II (*a*) (?). If not the same, name it and write the equation for its formation.

(*c*) State the test for the sulfate radical (see the experiment of the Purification of Water), and explain why all soluble sulfates give this test. Try several, and write the equations.

(*d*) Review the experiments on hydrogen chlorid and sulfur dioxid. Write the equations for the preparation of these substances. From these examples, state the effect of sulfuric acid on salts.

III. **Oxidizing Properties.** Place a little ball of copper turnings in a test-tube and add 1 cc. of concentrated sulfuric acid, noting any action (?). Heat carefully with a small flame, taking care to see that the acid does not spurt out over any one. Identify the gaseous product of the action as well as that in the tube, and explain how each is formed. Write the equation for the action.

Summary. Summarize the properties of sulfuric acid as to the following particulars: physical state, color, odor, taste, solubility, specific gravity, stability, action on litmus, on water, on carbohydrates, on salts, on barium chlorid solution, on elements like copper.

SULFURIC ACID IN THE INDUSTRIES

Among the many experimental endeavors of the alchemists was the heating of ferrous sulfate (green vitriol). A gas was formed that reacted with water to form a heavy, corrosive liquid which they called "oil of vitriol." This liquid is now called sulfuric acid, one of the most important substances of the modern world.

Basil Valentine studied this acid, for he records its method of preparation and its properties. He formed it in another way by burning sulfur with sodium nitrate in the presence of water. The nitrate oxidized the sulfur to sulfur trioxid which united with water to form the acid.

A Londoner, Ward, extended its preparation by this method on a commercial scale, using huge glass globes. Lead vessels soon replaced the more fragile glass ones.

Sulfur readily oxidizes to form the anhydrid of sulfurous acid, but further oxidation is difficult. If the gases are confined and heated to a higher temperature, the reverse action quickly undoes the desired combination.

Following the earlier methods of its preparation, chemists found that the oxids of nitrogen acted as carriers of oxygen which could be delivered to the sulfur dioxid, raising it to the trioxid. Further study developed the chamber process of making sulfuric acid. At first the oxids of nitrogen were allowed to escape into the air, but this made the process costly, and their bad odor created a considerable nuisance. Later the waste gases from the lead chambers were made to pass up through a tower (Gay-Lussac tower) and were deprived of the oxids of nitrogen by concentrated sulfuric acid which trickled down over stones so as to present a large surface to the gas. This acid was pumped to the top of a similar tower at the beginning of the plant (Glover tower), and in trickling down was deprived of the oxids of nitrogen by hot sulfur dioxid fumes from the sulfur burners. In this way the oxids of nitrogen found their way back into the lead chamber to be used over again. This improvement in the process reduced the cost of the commercial concentrated acid to about one cent a pound. The main disadvantages of the chamber process are the immense lead chambers required, and the fact that the acid produced is diluted with water and must be concentrated for most uses.

In 1831, Peregrine Phillips, a vinegar manufacturer of Bristol, a man of investigative turn of mind, discovered that finely divided platinum hastened the combination of sulfur dioxid and oxygen to such a great degree that at 400° C. the action was practically complete. As the platinum was not used up, it seemed that this method must surely supersede the lead-chamber process. However, when the method was tried with the sulfur dioxid obtained by the combustion of commercial sulfur, or the oxidation of naturally occurring sulfids, the platinum became "poisoned" and lost its catalytic power. One of the German coal-tar dye manufacturing companies, the Badische, spent over a million dollars in experimental work before they learned that the "poisoning" effect was due to certain impurities in the sulfur dioxid, especially arsenic, and how to remove them. It is now cheaper to make the pure concentrated acid by this process than by the chamber process.

At first the acid cost about $25 a pound, and its use was of course very limited. With each reduction in cost new uses have been found, and the quantity of acid manufactured in the world has grown to over 5,000,000 tons a year. This is far greater than the annual production of any other chemical industry except that of the chemical fertilizers. The history of sulfuric acid is a very interesting example of the manner in which man continually strives to better and cheapen processes of manufacture, and of the greatly increased use of an article that always follows every reduction in its cost. It also shows how the discovery of an apparently unimportant fact may be the starting point of a new industry, or may completely revolutionize an old one.

PREPARATION OF SULFURIC ACID BY CONTACT PROCESS

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Apparatus: Chlorid tower with stoppers and L-tubes, two 12-inch combustion tubes fitted with stoppers and L-tubes, 2 acid bottles fitted with stoppers, thistle tubes and L-tubes, U connecting tube, U-tube, beaker.

Materials: Very pure iron pyrites; ignited, long asbestos fiber, pumice stone, sodium chlorid, ice.

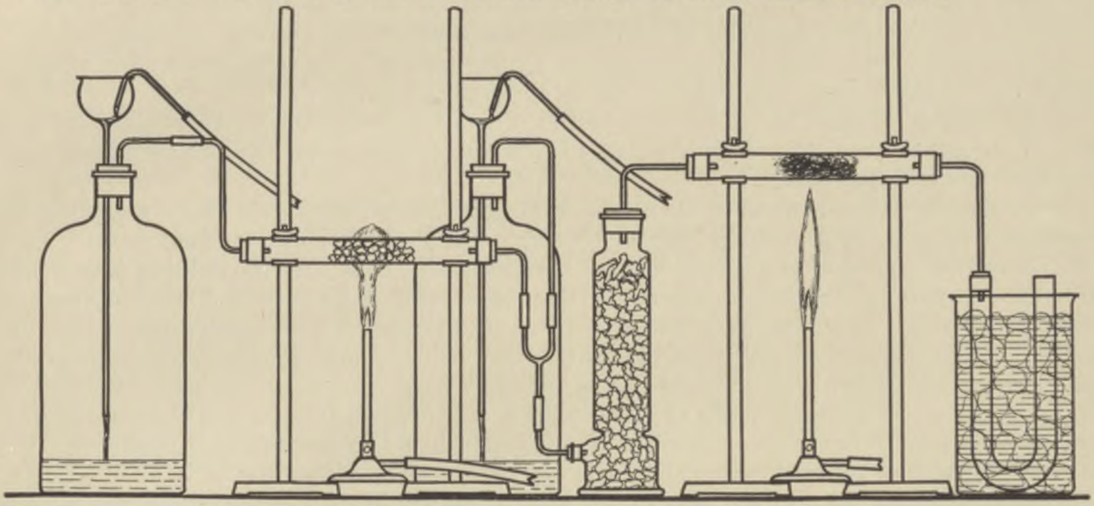


FIG. 87. Sulfuric acid by the contact process.

Fill the chlorid tower with small pieces of pumice stone which have been heated to redness and then soaked in concentrated sulfuric acid. Dissolve some scrap platinum in a very little concentrated aqua regia (3 volumes of hydrochloric acid to 1 of nitric) and dilute the solution of chlorplatinic acid, H_2PtCl_6, to 5 cc. with distilled water. Saturate the asbestos fibers with this solution, and heat them in an evaporating dish in the hood, slowly at first until they are dry, and then hot enough to decompose the chlorplatinic acid into platinum, chlorin, and hydrochloric acid. When the fibers are cool, place them in the middle of one of the combustion tubes and replace the stopper. Fill the middle portion of the other combustion tube with small pieces of very pure iron pyrites.

Assemble the apparatus as shown in Fig. 87, so that water entering the acid bottle on the left will drive air through the combustion tube containing the iron pyrites; and water entering the second acid bottle will force air into the U connecting tube, where it mingles with the gases from the pyrites tube, passes up through the pumice stone in the chlorid tower to be thoroughly dried, thence through the combustion tube containing the platinized asbestos, and finally through the U-tube which is surrounded by a freezing mixture of ice and salt (one-third salt by weight). The inside of all the apparatus following the drying tower must be as dry as possible.

When completely assembled, heat the pyrites tube gradually, and pass a slow stream of air through it. When the odor of sulfur dioxid is distinctly noticeable at the open end of the U-tube, run water into the second acid bottle at about the same rate as in the first. Note the appearance of the mixture of sulfur dioxid and air in the last part of the apparatus (?). Then heat the second combustion tube carefully and note any evidence of chemical action (?).

Describe the product. Add water to the trioxid drop by drop, to change it into sulfuric acid (?). Dilute the acid with water and test it with litmus paper and with barium chlorid solution (?).

Write six equations: (1) the action of chlorin, hydrochloric acid, and platinum giving chlorplatinic acid; (2) the decomposition of the chlorplatinic acid by heat; (3) the oxidation of iron pyrites, FeS_2, giving ferric oxid, Fe_2O_3, as one of two products; (4) the oxidation of the sulfur dioxid; (5) the hydration of sulfur trioxid; (6) the test for the sulfate ion.

DISCOVERIES

In all manufacturing operations there are waste products which are likely to be discarded with little attention because of their seeming worthlessness. Soap makers formerly allowed tons of glycerin to flow from the boiling tanks into the sewer because they did not know its value. Instead of paying thousands of dollars annually for the removal of its garbage, a certain eastern city receives a bonus of $600 per day from a company of thrifty men who collect the garbage and extract products of value from it. A chemist in a government mint had a suspicion that valuable metals were allowed to escape in the process of refining silver and gold. He began an investigation and found that thousands of dollars worth of platinum was going to waste every month.

In 1826 while Balard was studying the composition of the waste liquor after crystallizing the salt out of the water from salt marshes, he obtained a dark brown liquid having an unpleasant odor. He called it " muride " because of its source, but later changed it to " bromin " because of its bad odor. By this study of a waste product a new element was discovered for the use of men. An opportunity to discover the element, bromin, was offered to Justus von Liebig. Some years before Balard's discovery a salt company sent him a jar of liquid containing much bromin, and requested him to examine it. After a casual study of it he believed it to be iodin chlorid. After Balard's discovery Liebig saw the mistake which he had made by relying on preconceived ideas and not on experimental facts. He placed the jar in a special cabinet for storing mistakes — " l'armoire des fautes," as he called it. Balard became famous by his discovery of the element, bromin. His inferences and conclusions were based upon facts that he obtained directly from Nature by experiment. He used what is called the scientific method.

Near Stassfurt, Germany, much " Abraumsalz " (refuse salt) was thrown aside in the process of mining rock salt. Two chemists, Rose and Rammelsberg, made a study of it and found that it contained much potassium and magnesium. To-day these mines are worked for the salts of these metals and the sodium chlorid has become the by-product. The " Stassfurt Salts " constitute one of Germany's richest natural resources. The world's chief supply of bromin and potassium is obtained from these salts.

The salts and the acid of fluorin have long been known to scientific men, and many of the famous ones tried to isolate the element. H. Moissan succeeded in separating it from potassium fluorid by electrolysis in 1886. H. E. Roscoe called this task " one of the most difficult problems in modern chemistry."

" Every man has in himself a continent of undeveloped character. Happy is he who acts the Columbus to his own soul."

EXPERIMENT

Purpose: State it on the separate sheet used for record.
Materials: Potassium bromid, powdered manganese dioxid, carbon disulfid, starch.

I. **Preparation**. Bore a hole through a cork that has been fitted to a 15-cm. tube, and insert the short arm of the L-tube in the hole. Mix several medium-sized potassium bromid crystals, or 1 cc. powdered crystals, with twice its bulk of powdered manganese dioxid and put it in the 15-cm. tube. Adjust the clamp to hold the tube in a slanting position and fill another tube two-thirds full of cold water. Pour 2 cc. of concentrated sulfuric acid into 1 cc. of water and cool it. Then pour it upon the dry mixture in the other tube, insert the cork with L-tube, and hold the tube of cold water so that the long arm of the L-tube will dip **just below** the surface of the water. Heat the mixture **slowly** and **constantly** with a medium flame until the reddish vapor is driven beyond the bend in the L-tube. Do not heat enough to boil the black liquid through the L-tube nor stop heating until all the red vapor is driven over. Then stop heating and remove the tube of water **at once**. The dark liquid in the bottom of the tube is bromin. Describe the action in the generator during the preparation (?). Diagram the apparatus and label the parts. Write the equation and balance it. The molecule of bromin has two atoms.

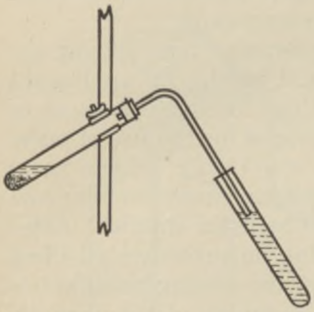


Fig. 88. Bromin from a bromid.

$$? \, KBr + MnO_2 + ? \, H_2SO_4 \longrightarrow KHSO_4 + MnSO_4 + ? + ?.$$

II. **Properties**. (*a*) The colored water above the bromin in the tube is called " bromin water." Pour 2 cc. of it into a 10-cm. tube, add 1 cc. (no more) of carbon disulfid, cover the tube with thumb and shake it. Most of the disulfid settles to the bottom. Compare the colors of disulfid and water, which show the relative solubilities of bromin in the two liquids (?).

(*b*) Wet a stirring rod and roll it in dry, powdered starch. Heat a drop or two of bromin liquid and hold the rod with starch in the vapor of bromin (?). This effect is used as a test for bromin.

(*c*) Test bromin water with both colors of litmus (?). Pour some on a little cochineal solution (?). Note the effect of the bromin on the cork used in the generator (?). Pour any liquid bromin remaining after tests into the stock bottle of bromin water.

(*d*) Tabulate the physical and chemical properties of bromin. Give the degree of solubility.

During storms on the ocean the waves bring considerable quantities of seaweed or kelp to the windward coasts and pile it up on the beach. The inhabitants of the coast collect it, and after drying, burn it to obtain lye from the ash. In 1812 a Frenchman, named Courtois, while using lye from this source in the manufacture of saltpeter, noticed that the copper vats used in the process were rapidly corroded. He traced this effect to an unknown substance in the lye. After separating out the various constituents of lye obtained from kelp he heated the remaining solution in a retort with sulfuric acid. He obtained " a vapour of a superb violet colour " which collected on the cool parts of the retort in crystalline plates. It was a new element, later to be called iodin from the color of its vapor. It is certain that his methodical study solved his practical problem, and brought him fame as a discoverer of a new element.

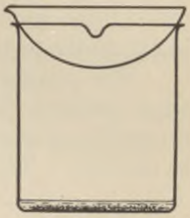


FIG. 89 *a*. Iodin from an iodid.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Materials: Potassium iodid, manganese dioxid, alcohol, carbon disulfid, sodium thiosulfate solution.

I. **Preparation**. Pour 5 cc. of concentrated sulfuric acid into the same quantity of water in the smallest beaker. Pulverize several medium-sized crystals of potassium iodid with twice its bulk of powdered manganese dioxid and put it in the prepared acid. Cover beaker at once with an evaporating dish filled with cold water. Warm the beaker with a small flame. If the violet vapor escapes, decrease the heat and refill the dish with cold water. Describe the action (?). Make a diagram (?). Write correct equation (?). The iodin vapor collects as a solid (sublimes) on the cold dish. Scrape it off on a paper.

Optional Procedure. Make a U-tube of 6 mm. diameter glass tubing of such width that it can be inserted into the mouth of a 250 cc. erlenmeyer flask. Have the U-part about 10 cm. long and the ends turned at right angles for side tubes, Fig. 89 *b*. Put the materials for preparing iodin in the flask and run cold water through the U-tube. Heat the flask slowly enough so that iodin vapor does not escape from the flask. When most of the iodin has sublimed on the cold U-tube remove it carefully and make the property studies.

II. **Properties**. Shake a small crystal with water (?). Add 2 cc. of carbon disulfid and shake (?). Shake a small crystal with 2 cc. of ethyl alcohol (?). Compare solubilities (?). Other solvents by chemical action are solutions of potassium iodid and sodium thiosulfate. Place a crystal of iodin on the palm of the hand for a moment (?). Try removing the stain with the iodid solution (?). Repeat with the thiosulfate solution (?). Test for iodin vapor with moist starch as was done for bromin (?). Recall the iodin test for starch. Put any remaining iodin crystals in the stock bottle.

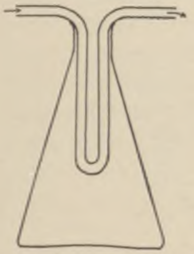


FIG. 89 *b*.

THE HALOGEN GROUP OF ELEMENTS

Fluorin, chlorin, bromin, and iodin form a remarkable family of elements. They are characteristic non-metals, so they unite with hydrogen, forming strong acids; and with metals, forming some very useful salts. All except fluorin are found in sea water. On account of this fact the other three were called **halogens**, meaning salt formers. Fluorin is now included in the term, making it apply to the whole group. The salts of these elements are often called **halids**.

The physical and chemical properties of these elements and their corresponding compounds show decided similarity. They differ in properties in degree only. When these elements are arranged in a certain order, a gradual transition of properties can be seen in passing from one element to another. The four elements form a series.

PROPERTIES OF THE HALOGEN ELEMENTS

Property	Fluorin	Chlorin	Bromin	Iodin
Atomic weight..........	19	35.5	80	127
State of aggregation	gas	heavy gas	liquid	solid
Color of gas............	light yellow	greenish-yellow	brownish-red	violet
Specific gravity.........	1.14 (liq.)	1.55 (liq.)	3.2 (liq.)	5.0 (solid)
Melting point............	– 233°	– 102°	– 7.3°	114.0°
Boiling point............	– 187°	– 33°	60°	183°
Solubility (100 c.c. at 20°)	decomposes water	0.0001 g.	0.032 g.	0.00015 g.
Activity................	greatest	less	still less	least

PROPERTIES OF THE HYDRIDS AND HALIDS

Property	H_2F_2	HCl	HBr	HI
Union with hydrogen takes place...........	In the dark at ordinary temperatures	In the sunlight	At red heat	Incomplete at red heat
Heat of formation in calories..................	38,500	22,000	12,300	– 6000
Stability of the hydrid ...	most stable	Decomposes at 1500°	Decomposes at 800°	Decomposes at 180°
Its melting point........	– 92.5°	– 116°	– 87°	– 57°
Its boiling point.........	19.4°	– 83.7°	– 64.9°	– 34.1°
Its solubility in water....	[1]	42.0%	49.0%	57.0%
Specific gravity, saturated water sol...............	1.15	1.21	1.49	1.7
Boiling point of water solution................	111° (43%)	110° (20%)	126° (47%)	127° (57%)
Heat of formation, potassium salt (cal.)........	110,600	105,700	95,300	80,100
Potassium salt melts.....	885°	790°	750°	705°
Calcium salt melts.......	1330°	780°	760°	740°
Calcium salt solubility...	0.16	42.7	58.8	67.6
Silver salt solubility (100 c.c. at 20°)............	181.8	0.016	0.00084	0.000028

---

[1] Completely miscible.

Exercise

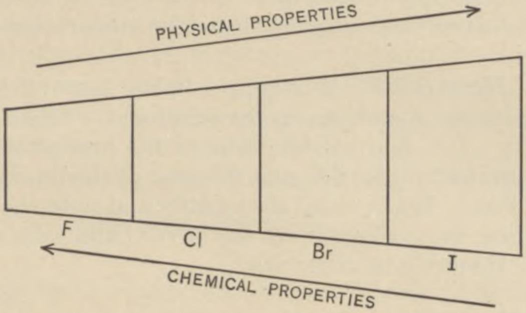
1. Give derivation and literal meaning of the names of the halogen elements.

2. Compare the order of variation of the physical properties with the order of variation of the atomic weights (?).

3. Point out two individual exceptions to this general rule.

4. Contrast the order of variation of the chemical properties (activity of elements and stability and heats of formation of compounds) with the order of variation of atomic weights.

5. Color the spaces in the diagram below to correspond with the colors of the respective gases of the elements. Explain how the diagram summarizes the properties of the halogen group.



EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Materials: Calcium fluorid, sodium chlorid, potassium bromid, potassium iodid, granular manganese dioxid, litmus paper, carbon disulfid. Solutions: silver nitrate, sodium chlorid, potassium bromid, potassium iodid, bromin, chlorin.

I. **Relative Stability of Hydrids.** To 1 cc. of washed calcium fluorid in a 15-cm. test-tube add 2 cc. of concentrated sulfuric acid (?). If no action is perceptible, warm the mixture. Make a record of the results in a tabulated form like that below. Try the odor **cautiously** (?). Hold moist blue litmus paper in the gas at the mouth of the tube (?). Pour a little ammonium hydroxid into a tube, dip the end of a stirring rod in it, and hold the rod in the mouth of the tube (?). Repeat the procedure to this point, using sodium chlorid instead of the fluorid (?). Repeat with potassium bromid and then with potassium iodid (?).

	CaF_2	NaCl	KBr	KI
Is a gas formed?				
Color of it?				
Odor of it?			[1]	[1]
Fuming? (Blow breath across the mouth of the tube)				
Effect on litmus?				
Effect with NH_4OH?				

---

[1] Note the odors of sulfur compounds from the sulfuric acid, mixed with other odors.

Questions: 1. From the effect on litmus, what compound must be formed in each case? Write the equations.

2. These hydrids formed are colorless. Explain the colors in the last two tubes by the instability of the hydrids.

3. Are the hydrids of fluorin and chlorin stable or not? How do you know? Add a few grains of manganese dioxid to the second tube (?). Explain the formation of the colored gas.

4. Add some manganese dioxid to the first tube (?). Is a colored gas formed (?). Explain the results by the relative stabilities of the hydrids.

5. Write the halogen acids in the order of their stability, greatest first.

6. State a general method of preparing an acid. Illustrate by an equation for the preparation of hydrogen fluorid. Why is this method not suitable for the preparation of two of the halogen acids?

II. **Properties of Silver Halid Salts.** In separate 10-cm. test-tubes put 1-cc. portions of solutions of sodium chlorid, potassium bromid, and potassium iodid. Add slowly, drop by drop, 1 cc. of silver nitrate solution to each (?). Record the colors of the precipitates with their formulas in a tabulated form. After the precipitates settle, pour off most of the liquid, and fill each tube half full of ammonium hydroxid solution. Shake them thoroughly and note the degree of solubility shown by each precipitate (?). These two properties of the silver halid salts are used to distinguish one from another. They are all insoluble in nitric acid.

Question: 7. What is the test for a chlorid, a bromid, and an iodid? Write the equations.

III. **Order of Replacement of the Halogen Elements.** Add 1 cc. of carbon disulfid to 1 cc. of potassium iodid solution in a 10-cm. test-tube, and shake. Note whether any free iodin is shown to be present by a purple color (?). Add several cubic centimeters of bromin water and shake again (?).

Question: 8. Which halogen element is now free? Which is combined? Equation.

Repeat the procedure with potassium bromid solution, carbon disulfid, and chlorin water (?). Interpret the results.

Questions.

9. What would be the probable result of adding chlorin water to potassium iodid solution? Verify your answer by experiment.

10. If chlorin water were added to a mixture of potassium bromid and iodid solutions, which halogen would be replaced first? Verify your answer by experiment.

11. From the previous relationships shown, where would fluorin stand in a replacement series? Write the series in proper order.

NITRIC ACID

Without explosives much of our present life and work would be impossible of realization. Quarrying, mining, and excavating are industries making extensive use of explosives. The great tunnels, the Panama Canal, and many other great undertakings would be practically impossible if explosives were not available. Explosives thus have a constructive use which far exceeds in value their use in warfare. All of the explosives in use to-day are derivatives of nitric acid, and immense quantities of the latter are used in their preparation.

The explosiveness of so many nitrogen compounds depends upon the fact that the union between nitrogen and oxygen is never a firm one, and is in some compounds very easy to break. Although not sufficiently unstable to be explosive, the nitric acid molecule breaks up quite easily, and some of the most important properties of nitric acid depend upon this decomposition. Concentrated nitric acid decomposes slowly at the ordinary temperature, as is shown by the fact that the pure colorless acid soon becomes colored yellow by one of the products of its decomposition being dissolved in it. This disruption of the acid molecule is hastened by heating or by the presence of some substance which can react with the products of the decomposition.

Like other ternary acids, nitric acid may be regarded as consisting of water and a non-metallic oxid, its anhydrid.

$$2\,HNO_3 \rightleftarrows H_2O + N_2O_5.$$

Since the water is exceedingly stable, the breaking up of the nitric acid molecule can be regarded as a decomposition of the nitrogen pentoxid. When concentrated nitric acid is decomposed, a brown gas, nitrogen tetroxid (NO_2), is always noticed; while dilute nitric acid always gives a colorless gas, nitric oxid (NO), which has the property of combining with free oxygen to form the brown gas. These are the principal actions occurring under these conditions.

$$2\,HNO_3(H_2O, N_2O_5) \longrightarrow H_2O + 2\,NO_2 + [O] \quad \text{(concentrated acid).}$$

$$2\,HNO_3(H_2O, N_2O_5) \longrightarrow H_2O + 2\,NO + [3\,O] \quad \text{(dilute acid).}$$

The first of these actions takes place even at ordinary temperatures, the oxygen being given off as gas, O_2. Water makes the nitric acid more stable, so that the dilute acid does not decompose unless some substance is present to combine with the oxygen, and the first action is much hastened if such a substance is present. In both cases, nitric acid is an **oxidizing agent**, and the concentrated acid is one of the most powerful oxidizing agents known. Although the dilute acid will give up a larger proportion of its oxygen, the presence of the water makes the oxidation much less vigorous, so that the dilute acid is not able to oxidize some substances that the concentrated acid acts upon very easily. Since the oxidizing property belongs to the molecules, the concentrated acid is a better oxidizer than the dilute acid which has its molecules largely separated into ions.

When nitric acid acts upon most metals, it acts in both capacities, as oxidizing agent and as acid. With the inactive metals, one must regard the nitric acid as oxidizing the metal first, and then dissolving this oxid to form the nitrate of the metal. With the active metals, one can consider that the metal first displaces hydrogen as with other acids, but that this hydrogen is immediately (incompletely, with the most active metals) oxidized into water by the oxygen given off by the nitric acid in equation (1) or (2) above.

References: Martin, *Modern Chemistry and its Wonders*, Chaps. II and III, pp. 24–87.

NITRIC ACID

EXPERIMENT

Purpose: Write it on a separate sheet used for the record.

Apparatus: 250 cc. tubulated retort, condenser clamp.

Materials: Sodium or potassium nitrate, sodium carbonate, ferrous sulfate solution, white woolen yarn, copper, zinc, charcoal, iron.

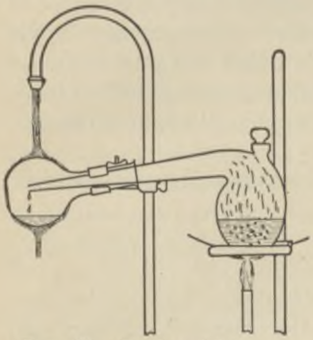


Fig. 90. Nitric acid from a nitrate.

I. **Preparation.** By means of the condenser clamp support a flask under the faucet so that the jet of water will strike the bulb of the flask and cool it by flowing all around it. Put 25 g. of either sodium or potassium nitrate in a 250-cc. tubulated retort and insert the neck of the retort into the mouth of the flask so that the end of the retort is in the middle of the bulb of the flask. The neck of the retort should slope slightly downward. (See Fig. 90.) By means of a funnel pour 15 cc. of concentrated sulfuric acid upon the nitrate crystals and replace the stopper. **Caution:** Sulfuric and nitric acids produce severe wounds on the flesh and destroy clothing, so the greatest care must be used in handling them.

As soon as the sulfuric acid has moistened the entire mass, heat the retort carefully, distilling at as low a temperature as possible until no more nitric acid condenses in the neck of the retort. Meanwhile, proceed with II (*f*). Note the color of the gas in the retort, and explain how it was formed.

When the retort is cool enough to touch, the residue may be extracted with warm water. It is sodium (or potassium) hydrogen sulfate, $NaHSO_4$. Write the equation for the preparation of the nitric acid.

II. **Properties.** (*a*) Note the color of the acid you have prepared. Explain (preceding page). Boil 3 cc. of the nitric acid in a 15-cm. test-tube (shaking prevents spurting) and note the color of the gas given off at first (?), also the color of the gas later when the nitric acid is boiling rapidly (?), and the color of the acid left in the tube (?).

(*b*) Keep the acid at the boiling point and insert a glowing spark on the end of a long stick into the fumes (?). If the spark is extinguished, repeat the operation at a higher temperature or else use a fresh sample of the acid. Note the gas above the spark (?), and also the color given to the uncharred wood by the acid fumes (?). Write the equation for the effect of heat upon concentrated nitric acid.

(*c*) Pour 3 cc. of the acid into the porcelain evaporating dish. Blow the breath gently across the acid (?). Twist an iron wire around a bit of charcoal, ignite the latter, and dip it into the acid (?). Explain the action and write the equation.

(*d*) Pour 1 cc. of the acid upon white woolen yarn (?), rinse with water, and then add ammonium hydroxid (?). The action of nitric acid followed by ammonium hydroxid is used as a test for proteins.

(*e*) Pour 1 cc. of the acid upon a piece of copper (?). This action is used as a test for nitric acid, by itself, or in mixtures. Write the equation (preceding page).

(*f*) **Test for the Nitrate Ion.** (Follow instructions carefully.) Add 2 drops of dilute nitric acid to 3 cc. of ferrous sulfate solution in a test-tube and shake. Incline the tube slightly and pour about 2 cc. of concentrated sulfuric acid in a steady stream down the wall of the tube so that it will sink to the bottom without mixing with the ferrous sulfate solution. Describe the coloration in the layer where the liquids meet (?). Repeat, using a small crystal of some nitrate instead of the nitric acid (?).

(*g*) Summarize the physical and chemical properties of nitric acid, and state the tests for nitric acid and for the nitrate ion.

EXPERIMENT

Purpose: Write it on a separate sheet used for the record.
Material: Copper scrap, sulfur, oxone, candle, asbestos paper.

I. **Preparation.** Put some pieces of copper in the generating bottle, and close it with a 2-hole stopper fitted with a thistle tube, L-, and delivery tubes. Hold the generator several inches from the table (Fig. 91) so that a half-filled beaker of cold water can be placed under it. Arrange to collect 4 or 5 jars of the gas over water. Fill a 15-cm. tube with water, invert over water, and support it by a clamp with its mouth just below the surface of the water. Pour water into the generator until it is 2 cm. deep and the end of the thistle tube is covered. Then add half as much concentrated nitric acid. If the action becomes very vigorous, raise the beaker of water until the bottle is half immersed. The action will then become slower. Note the color of the gas first formed in the generator (?). After collecting two jars of the gas, invert a 10-cm. tube — filled with water — over the outlet of the delivery tube and fill it with gas. Transfer the gas by pouring it **upward** into the 15-cm. tube. Add another tubeful in the same way. Fill the remaining jars with gas. If either acid or copper is exhausted, put more in the generator. Make diagram of apparatus and complete the following equation.

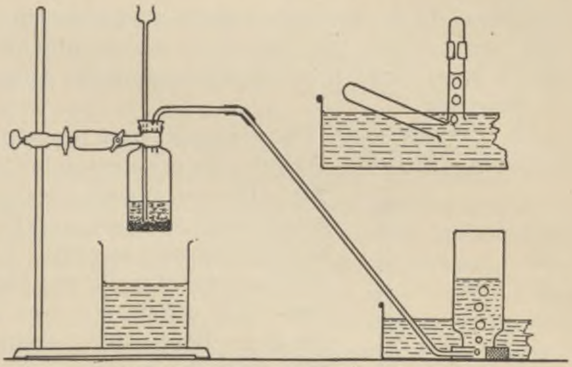


FIG. 91. Nitric oxid from nitric acid.

$$? \, Cu + 8 \, HNO_3 \longrightarrow \, ? \, NO + 3 \, Cu(NO_3)_2 + \, ?.$$

II. **Properties of Nitric Oxid.** (*a*) Note color of gas in the jars (?).

(*b*) Set a jar of the gas, mouth down, in a pan of fresh, cold water and remove the glass plate. Let the jar stand in the water until the end of the period, and judge as to the solubility of the gas by the amount of water which has entered (?).

(*c*) Lower a burning candle into a jar of the gas (?).

(*d*) Line a spoon with asbestos paper, put in some sulfur, and ignite it by directing the burner, flame downwards, on it. Lower the burning sulfur in a jar of the gas (?).

III. **Properties of Nitrogen Tetroxid.** (*a*) Uncover a jar of the gas (?).

(*b*) The action in III (*a*) may be understood by the following: Drop a piece of oxone — size of a pea — in water and invert a 10-cm. tube filled with water over it at once. Pour the collected oxygen into the 15-cm. tube partially filled with the gas. Watch closely and state all that happens (?). The brown gas is NO_2, nitrogen tetroxid (N_2O_4, at low temperatures). Write the equation (?).

Questions: 1. What is the degree of solubility of nitrogen tetroxid?

2. What is the ratio of the volumes of nitric oxid and oxygen that combine? Any gas in the tube not dissolved is due to impurities or error in measurement of gases.

3. How is the combining ratio of gases shown in correctly written equations? See the last equation written.

4. Explain the result of exposing a jar of nitric oxid to air.

5. Why was the gas brown at first in the generator?

6. What became of this brown gas?

7. Does nitric oxid burn or support combustion?

8. Do you think nitrogen tetroxid would burn or support combustion? Give reasons.

9. Summarize in parallel columns the properties of nitric oxid and nitrogen tetroxid.

10. How could these two gases be distinguished from each other?

NITROUS OXID

EXPERIMENT

Purpose: State it on a separate sheet used for the record.
Materials: Ammonium nitrate, sulfur, splinter or candle, asbestos paper.

Five oxids of nitrogen are known, the largest number formed by any of the non-metallic elements. They afford an excellent example of the law of **multiple proportions**. In unit volume (22.4 L. at S. T. P.) of the gaseous oxids of nitrogen, the weight of the nitrogen is in every case a simple multiple of 14 g., the unit or gram-atomic weight of nitrogen; the weight of oxygen, similarly, is in each case a simple multiple of 16 g., the unit or gram-atomic weight of oxygen. These relations make a satisfactory proof of the existence of atoms. The multiples indicate the number of atoms in the molecules.

I. **Preparation**. Put about 10 g. of ammonium nitrate in the flask and clamp it in a nearly horizontal position. Attach the generating bottle as a condensing trap by means of L-tubes (Fig. 92). Use warm water in the jars and pan when collecting the gas. Put a wire gauze under the flask and let one pupil give his entire attention to heating the flask gently and steadily. After the solid has melted keep the flame moving, since overheating in one place will cause trouble. Let the first jarful escape since it is chiefly air from the flask and bottle. Collect four small jars of the gas. **Before the heating is discontinued be sure to remove the delivery tube from the water**, so the water will not come back into the bottle. Test the substance in the trap bottle with cobalt chlorid paper (?). Identify it and write the equation of the decomposition of ammonium nitrate. Make diagram of apparatus.

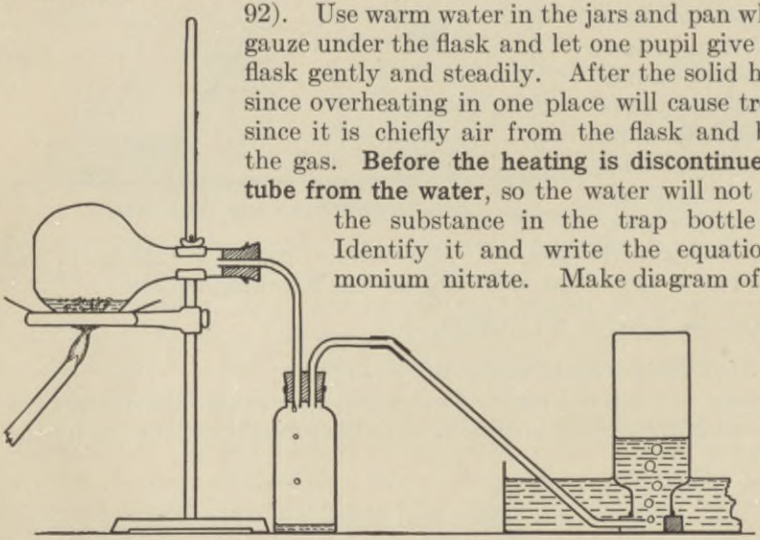


Fig. 92. Nitrous oxid from the decomposition of ammonium nitrate.

II. **Properties**. (*a*) Describe the color and odor of the gas (?).

(*b*) Invert a jar of the gas over cold water, and leave for ten minutes (?). State the degree of solubility of the gas.

(*c*) Insert a glowing splinter in a jar of the gas (?).

(*d*) Line a spoon with asbestos paper, put some sulfur in it, ignite by directing burner flame downward on the sulfur, and when well ignited, put it in a jar of the gas (?).

Questions: 1. Why was the gas collected over warm water?

2. Why does nitrous oxid support combustion when nitric oxid does not? Explain difference by instability.

3. Why does this gas support combustion better than air? Consider the proportions of oxygen.

4. By what properties could this gas be distinguished from nitric oxid? From oxygen?

5. Write equations to show the formation of nitrous and nitric acids from their respective anhydrids, N_2O_3 and N_2O_5.

PHOSPHORUS

Phosphorus is an essential constituent of all living matter. Plants take it up from the soil as soluble salts of phosphoric acid, and build it into their stalks and seeds. Kernels of corn contain approximately 1% of phosphorus. Soils frequently lack sufficient of some or all of three elements, potassium, phosphorus, and nitrogen, to nourish plant life properly. Hence complete fertilizers usually contain at least these three elements in combined form. Animals obtain phosphorus from plants. It is an important constituent of the bones and nerve tissue. It is excreted in the body wastes, particularly by the kidneys, as sodium ammonium phosphate. In this form it goes back to the soil to be used in another cycle of life.

The ancient alchemists sought with great diligence for a substance that would turn baser metals into gold. They tried all manner of experiments with various substances, dissolving, burning, and distilling with never ending zeal. In some such quest a Hamburg alchemist, Brand, distilled a mixture of sand and urine and accidentally obtained this element in a free state. He is credited with the discovery of the element. It derived its name from the peculiar property of emitting a greenish light when exposed to the air. This light is due to the oxidation of the phosphorus. Previous to the discovery of this element other substances were known which gave out light that was not due to burning. The sulfids of barium and calcium, after exposure to a bright light for some time, appear luminous in the dark. The light appears to be absorbed and given out again similar to heat. Such substances are used in paints for clock faces, so that the clock may be useful after dark. Wood from the interior of a well-rotted log is often found to glow brightly in the dark. The ocean water presents a similar aspect at night. These effects are due to the action of minute organisms, but they are not fully understood.

Such substances giving out light were called phosphorescent, from the Greek words, φῶς (phos) — light; and φέρειν (pherein) — to carry. When Brand discovered the new element — it was called Brand's phosphorus to distinguish it from the other phosphorescent materials. In time the term was restricted to this element.

Two allotropic modifications of phosphorus are known, differing remarkably in properties.

Property	Red Phosphorus	White Phosphorus
Color	Chocolate red	White-colorless
Odor	Odorless, tasteless	Garlic-like, ozone
Crystalline form	Rhombohedral	Cubic
Specific gravity	2.2	1.8
Melting point	500° to 600° C.	44° C.
Physiological	Non-poisonous	Poisonous
Solubility in CS_2	Insoluble	1 g. dissolves 10
Exposed to air	No phosphorescence, no oxidation	Phosphorescence and oxidation
Ignition point	240° C.	34° C.

The larger part of phosphorus produced is used by the match industry. By its use much inconvenience in starting a fire has been removed. The poisonous nature of the white variety has brought it into general disfavor in most countries. A compound, phosphorus sesquisulfid (P_4S_3), not having this fault, has taken its place in the manufacture of the ordinary match.

Study Topics: Martin, *Triumphs and Wonders of Modern Chemistry:* "The Phosphorus Group of Elements"; "The Story of a Phosphorus Atom," pp. 308–323.

EXPERIMENT

Purpose: State it on the separate sheet used for record.

Material: White and red phosphorus, carbon disulfid, filter paper, paraffin, glue, potassium chlorate, antimony trisulfid, phosphorus sesquisulfid, zinc oxid, ferric oxid.

Caution! White phosphorus takes fire readily, so it should never be touched with the hands. Experiments with it should be carried out with utmost care. Any residues of phosphorus should not be put in waste receptacles but should be burned in the hood.

I. **White Phosphorus.** (*a*) Note the physical properties of white phosphorus (?).

(*b*) Put a tiny piece of white phosphorus in the evaporating dish. Dry, by touching it with the end of a small roll of filter paper. Watch the action taking place while exposed to the air (?). In a few moments take it to the hood and warm the dish (?). When the combustion has ceased, examine the residue in the dish (?). Heat the residue quite strongly (?). Remember to change the temperature gradually.

(*c*) Put a bit of white phosphorus in a 10-cm. tube and add 2 cc., no more, of carbon disulfid (?). Pour the liquid on a filter paper placed on an iron ring. Do not get any of it on the hands or the clothing. Watch the paper until action begins (?).

Questions: 1. The color of the residue in the dish and its action when heated strongly shows it to be what substance?

2. Explain why it did not burn when previously ignited.

3. Why did the phosphorus from solution ignite more readily than the piece in the dish?

4. State the conditions necessary to change white phosphorus to red. See I (*b*).

II. **Red Phosphorus.** (*a*) Note the physical properties of red phosphorus (?).

(*b*) Test the solubility of a bit of red phosphorus in 2 cc., no more, of carbon disulfid (?).

(*c*) Close one end of a 10-cm. piece of glass tubing, 5 mm. diameter or larger. When the tube is cool, push a little red phosphorus into it near the closed end and heat that portion of the tube slowly and steadily, holding the tube horizontally and **not** having the open end pointed at any person. When the red color of the phosphorus has nearly disappeared, cool the tube and note the color of the deposit in the tube (?). Break the tube near the deposit and rub it with a splinter of wood (?). Bring the deposit near, not into, a flame (?).

Question: 5. State the conditions for changing red phosphorus into white.

III. **Preparation of a Safety Match.** (*a*) **The Match.** Dip 10 prepared match sticks into hot liquid paraffin for a moment. Then dip them in a thin paste made from 3 drops of thin glue, the same bulk of powdered potassium chlorate, and half as much antimony sulfid. Dry them thoroughly and try to ignite them on the rubbing surface when it is also dry.

(*b*) **The Rubbing Surface.** Mix 3 drops of thin glue with half as much red phosphorus, and the same amount of finely powdered glass or silica. Spread the mixture over the top of a match box or a piece of pasteboard and let it dry.

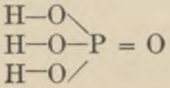
IV. **Preparation of a Friction Match.** Grind about 3 grams of glass to a fine powder in a mortar. **Protect the eyes.** Next grind carefully, by itself, 12 grams of potassium chlorate. Put with the powdered glass and chlorate 3 grams of pulverized phosphorus sesquisulfid (P_4S_3), 3 grams of zinc oxid, and 3 grams of ferric oxid. Mix thoroughly with about 9 grams of glue, adding water until the mixture is a thin paste. Dip paraffined sticks into the paste and dry them. When dry, they will ignite on any moderately rough surface.

The phosphorus sesquisulfid is non-poisonous and is generally used commercially in making ordinary matches.

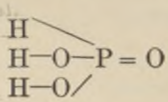
ACIDS OF PHOSPHORUS

From its non-metallic properties, phosphorus might naturally be expected to be a distinctive element in the formation of acids. Three are known whose compositions permit them to be named according to the plan previously given (p. 45) : phosphoric acid, phosphorous acid, and hypophosphorous acid. Although each one has three hydrogen atoms in the molecule, experiment shows that only the first is tribasic (all three hydrogen atoms replaceable by metals). The second is found to be dibasic, and the third monobasic. These facts are best expressed by the following formulas :

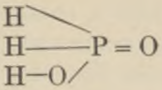
(1) Phosphoric acid, (ortho) H_3PO_4,



(2) Phosphorous acid, H_2PHO_3,



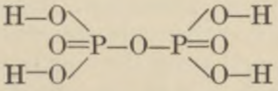
(3) Hypophosphorous acid, HPH_2O_2,



When the hydrogen atoms are united directly to the characteristic non-metallic element in ternary acids instead of being united to the oxygen, they do not dissociate in a water solution to form hydrogen ions, and are not replaceable by metals.

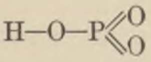
There are two partially dehydrated forms of phosphoric acid that are important. Phosphoric acid is generally referred to as orthophosphoric acid to distinguish it from the following acids. When orthophosphoric acid is heated to 250° C., two molecules of the acid lose one molecule of water and form pyrophosphoric acid.

$$2\,H_3PO_4 \xrightarrow{250^\circ\ C.} H_2O + H_4P_2O_7,$$



If the temperature is raised to 800° C., each molecule of orthophosphoric acid loses a molecule of water and metaphosphoric acid is formed.

$$H_3PO_4 \xrightarrow{800^\circ\ C.} H_2O + HPO_3,$$



The bones of animals not put to other uses are made into fertilizer. Some years after the wholesale destruction of bison on the western plains for their hides, a considerable number of men found profitable employment in gathering the skeletons for the manufacture of fertilizer. The chief plant food in bones is phosphorus in the form of phosphates. A cheaper and more abundant supply of phosphates has been found in the beds of phosphate rock in the southern states. The rock, calcium phosphate, is converted into soluble superphosphate, $CaH_4(PO_4)_2$, by the action of acids. In this form it is ready to begin another life cycle, passing from the inorganic form into the organic in the structures of plants and animals and then, by decay, back to inorganic Nature. Chemical action is the chief cause of the never ending changes that come to notice every day and from year to year. It may be said that the industrial world moves by means of chemical action. But the amount of matter in the universe is unvarying. Matter does not wear out. If the laws of the universe maintain the constancy of matter so rigidly, what may be expected for the guardianship of creation's masterpiece, human kind, that is able to control matter for its purposes? Is it possible that the "three score years and ten" sums up the entire existence of man?

Study Topics: Thorp, *Outlines of Industrial Chemistry:* "Fertilizers," pp. 146–156.
Duncan, *New Knowledge:* "The Mystery of Matter," pp. 15, 16. *Some Chemical Problems of To-day:* "On the Chemical Interpretation of Life," pp. 63–83.
Lassar-Cohn, *Chemistry in Daily Life:* "Fertilizers," pp. 36–45.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Materials: Phosphorus, litmus paper, solutions of silver nitrate, albumen, magnesia mixture, and ammonium molybdate, phosphate rock, orthophosphoric acid, sodium phosphate, asbestos paper.

Reagent	Ortho	Pyro	Meta
$AgNO_3$	Canary-yellow precipitate	White crystalline precipitate	White gelatinous precipitate, darkened by light
Albumen	Nothing	Nothing	Coagulated

I. **Preparation of Phosphoric Acids**. (*a*) Line a combustion spoon with asbestos paper, and burn some phosphorus in a large, clean glass jar, having about 20 cc. of distilled water in it. When the action has ceased, remove the spoon and shake the jar to dissolve the pentoxid (P_2O_5) formed. Determine the nature of the solution with litmus (?). Add one-half the solution to albumen in water (?). Make the other half nearly neutral with dilute ammonium hydroxid, and add silver nitrate solution drop by drop until a precipitate appears (?). Save it. Compare the results of the tests with the table above and determine which kind of phosphoric acid is present (?). Write the equation for its formation and also for the test with silver nitrate. After ten minutes note the effect of the light upon the color of the silver salt (?)

(*b*) The acid formed by the operation in Section (*a*) can be converted into orthophosphoric acid by long boiling with nitric acid as a catalytic agent. Chemical action takes place with water. Write the equation for its formation. Nearly neutralize some orthophosphoric acid with dilute ammonium hydroxid, and then add silver nitrate solution drop by drop until a precipitate forms. Name and describe it and write the equation for its formation.

(*c*) Evaporate 25 cc. of orthophosphoric acid solution to dryness, setting the burner beneath the dish and adjusting the flame so that the tip **does not** quite touch the dish. After the solid melts continue the heating until it is no longer a liquid. Cool the dish slowly and dissolve the residue in cold water. Divide the solution in two parts, and test the first with egg albumen solution (?), and the second with silver nitrate solution after making it nearly neutral, as in the previous operations (?). Save the precipitate in the latter test for ten minutes and then note the effect of light upon its color (?). Compare the results of these tests with the preceding page and determine what acid was formed by heating the ortho acid (?). Write the equation for its formation and for its silver nitrate test. Summarize the methods of preparing the three phosphoric acids.

II. **Other Tests for Orthophosphoric Acid.** (*a*) Dissolve some phosphate (sodium or calcium) in dilute nitric acid. Add 10 cc. of ammonium molybdate solution and allow the mixture to stand for several minutes, then warm it a little (?). A precipitate shows the presence of the phosphate ion (PO_4). Describe the precipitate (?).

(*b*) Add to a solution containing phosphate ions some " magnesia mixture " (ammonium hydroxid, ammonium chlorid, and magnesium sulfate). The precipitate forms slowly, so allow it to stand for at least ten minutes. It may be hastened by rubbing the glass vessel with a stirring rod. Describe the precipitate, $Mg(NH_4)PO_4$, and write the equation (?).

III. **Calcium Superphosphate**, $CaH_4(PO_4)_2$. Put 10 g. of dry powdered phosphate rock containing calcium phosphate (apatite, phosphorite, or bone ash) in an evaporating dish and add 4 cc. each of water and concentrated sulfuric acid. Stir the mixture occasionally while warming it for ten minutes. Add about 10 cc. of water, filter, and test the liquid for an orthophosphate (?). Write the equation for the forming of the superphosphate (?). If more acid had been added, the products would have been calcium sulfate and orthophosphoric acid. Write this equation also (?).

Questions: 1. To what class of salts does the superphosphate belong?

2. Summarize the second method of preparing orthophosphoric acid.

THE NITROGEN GROUPS OF ELEMENTS

A decided relationship of certain elements was noted on two previous occasions, when they were grouped together to make what is known as the halogen and the oxygen-sulfur groups. These two groups were composed entirely of non-metallic elements. Another most interesting group of five elements, nitrogen, phosphorus, arsenic, antimony, and bismuth, — arranged in the order of their atomic weights, — shows the same gradual transition of properties from the first element to the last in the group. The nitrogen group is more remarkable than either of the other groups in that the transition takes place from non-metallic nitrogen to the metal, bismuth.

Property	Nitrogen	Phosphorus	Arsenic	Antimony	Bismuth
Atomic weight......	14	31	75	120	208
Specific gr..........	1.026	1.82	5.7	6.7	9.8
Melting point......	– 214°	44°	480°	629°	270°
State..............	gas	solid	solid	solid	solid
Boiling point.......	– 194°	287°	450°	1440°	1600°
Character..........	Non-metallic	→	→	→	metallic
Valence............	3, 5	3, 5	3, 5	3, 5	3, 5
Hydrid............	NH_3	PH_3	AsH_3	SbH_3	——
Heat of formation...	12,000 cal.	111,600 cal.	– 36,700 cal.	– 81,800 cal.	
Character..........	basic	feebly basic	neutral	neutral	
Chlorids...........	NCl_3	PCl_3	$AsCl_3$	$SbCl_3$	$BiCl_3$
Boiling point.......	71°	78°	130.2°	200°	447°
Stability...........	extremely unstable	→	→	→	stable
Oxids.............	N_2O				
	$(N_2O_2)NO$				Bi_2O_2
	N_2O_3	P_2O_3	As_2O_3	Sb_2O_3	Bi_2O_3
	N_2O_4	P_2O_4		Sb_2O_4	Bi_2O_4
	N_2O_5	P_2O_5	As_2O_5	Sb_2O_5	Bi_2O_5
Hydroxids	—	$(HO)_3P$	$(HO)_3As$	$(HO)_3Sb$	$(HO)_3Bi$
Trivalent........	HONO	—	HOAsO	HOSbO	HOBiO
	weak acid	weak acid	acid or base	acid or base	base
Pentavalent......	—	$(HO)_3PO$	$(HO)_3AsO$	$(HO)_3SbO$	—
	$HONO_2$	$HOPO_2$	$HOAsO_2$	$HOSbO_2$	$HOBiO_2$
	strong acid	weak acid	weak acid	weak acid	very weak acid

Bismuth presents an exception to the expected order of melting points of these elements. Yet its principal use is based on this exception. Because of its low melting point it is used in making low melting point alloys for use in automatic fire sprinklers and safety plugs in boilers.

Some exceptions to general rules seem to be significant. When mixed with other metals, antimony causes the alloy formed to expand slightly as it freezes to a solid. Such an alloy is used to cast into type because it fills the mold completely, making perfect type. Water expands on cooling below 4° C., and this exception has an enormous bearing on the world's economy as it is now constituted. In a similar way the element bismuth is more valuable, having the exception, than if its properties followed the regular order.

THE HYDRIDS OF THE NITROGEN GROUP

EXPERIMENT

Purpose: Write it on a separate sheet used for the record.

Material: White phosphorus, solutions of sodium hydroxid (25%), arsenic trichlorid, antimony trichlorid, and sodium hypochlorite; zinc, aluminum, anhydrous calcium chlorid.

I. **Hydrid of Nitrogen, Ammonia.** (*a*) Recall the preparation of ammonia and its properties. Write the equation.

$$? NH_4Cl + Ca(OH)_2 \rightleftarrows ? + ? + ?.$$

(*b*) **Reduction of a Nitrogen Compound.** Put 5 cc. concentrated sodium hydroxid solution on some aluminum in a test-tube and heat it. When bubbles of hydrogen are formed by the chemical action, add a little nitric acid, and note the odor of the escaping gas (?). Write two equations: one for the formation of hydrogen and sodium aluminate, Na_3AlO_3; the other for the reduction of nitric acid to ammonia by hydrogen.

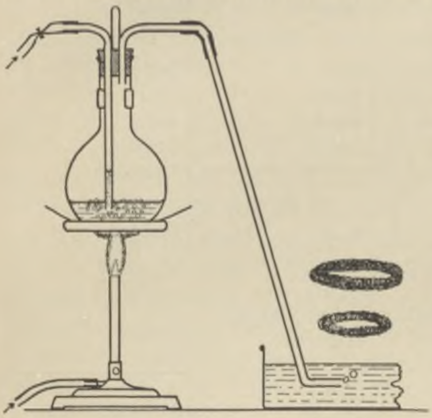


Fig. 93. Preparation of phosphin.

II. **Hydrid of Phosphorus, Phosphin.** Pour sodium hydroxid solution (25 %) into the flask until it is 2 cm. deep. Drop in a 1-cm. piece of white phosphorus and close the flask with a 2-hole stopper carrying an L-tube which extends into the solution in the flask. Another L-tube with attached delivery tube is used to carry the gas from the flask below the surface of water in a pan (Fig. 93). Place the apparatus in the hood. Attach a rubber tube to the long L-tube and pass the laboratory gas through the flask for 2 minutes to sweep the air out. Shut off the gas or close the tube with a clamp and heat the solution with a small flame. Keep the liquid boiling gently until the properties of the gas, phosphin, are known, then remove the burner. Water will flow from the pan into the flask and help stop the action. Make a diagram of the apparatus and complete the equation.

$$? NaOH + ? P + 3 H_2O \longrightarrow ? NaH_2PO_2 + ?.$$

Test the water in the pan with litmus and record the properties of phosphin (?).

Questions: 1. Why was the air swept out of the flask at the beginning?

2. Name the other product formed beside phosphin.

III. **Hydrid of Arsenic, Arsin.** Set up a hydrogen generator (Fig. 94) and start the action by pouring dilute hydrochloric acid on the zinc in the generator. Attach a drying tube, filled with granular anhydrous calcium chlorid, to remove moisture from the gas. Attach a jet tube to the drying tube. If this tube has a platinum tip, the flame will not be colored yellow by the sodium in the glass. After the action has continued for fully 5 minutes ignite the gas issuing from the jet. Wrap a towel loosely about the generator to lessen the danger from a possible explosion. While the hydrogen is burning, add a little of a solution of arsenic trichlorid ($AsCl_3$) and note the effect on the rate of action in the generator and the color of the flame (?). Note the odor of the fumes, but do not breathe them (?). Let the flame touch the inside of the evaporating dish, making a metallic deposit on the cold porcelain (?). Replace the jet tube with a straight piece of tubing and heat it near the

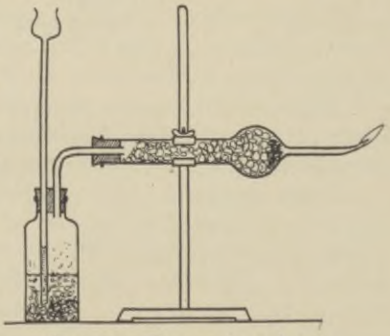


Fig. 94. Preparation of arsin.

center with the burner flame. Note the deposit in the tube (?). This last test is known as Marsh's test for arsenic by which 0.00001 g. of arsenic can be detected. Make a diagram of the apparatus and write the equation of the action of arsenic trichlorid and hydrogen (?). The deposit on the porcelain will dissolve in a solution of sodium hypochlorite. Try it.

IV. **Hydrid of Antimony, Stibin.** Repeat the procedure in Part III with fresh materials, but this time use a solution of antimony trichlorid ($SbCl_3$) instead of the arsenic compound. Make all the tests and write the equation (?). Finally summarize the properties of arsin and stibin in parallel columns under these heads: color, odor when burning, deposit on cold porcelain, deposit in a heated tube, solubility in sodium hypochlorite. (The deposit from stibin is insoluble.)

V. No hydrid of bismuth is known. State reasons why it may not exist (?).

SILICON COMPOUNDS

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Apparatus: Iron crucible.

Materials: Sodium silicate (water-glass), sand, calcium fluorid, sodium carbonate, broken chemical glass, phenolphthalein.

I. **Metasilicic Acid.** Dilute a few cubic centimeters of sodium silicate, Na_2SiO_3 (water-glass), with twice the volume of water and add concentrated hydrochloric acid, a few drops at a time, until 2 cc. have been poured in (?). If no precipitate forms, add a few drops of ammonium hydroxid to hasten it. Describe the precipitate and write the equation for the formation of this acid, disregarding the ammonium hydroxid. Empty the contents of the tube into the evaporating dish and evaporate to dryness. Then add dilute hydrochloric acid and rub the bottom of the dish with a stirring rod (?). The gritty feeling like sand is the "insoluble silicic acid" (silica). This procedure constitutes a test for silicates. Evaporate the contents to dryness again and save the dried residue for the next experiment.

II. **Fluosilicic (Hydrofluosilicic) Acid Test for Silicon.** Put a little sand and calcium fluorid in a test-tube and add about 5 cc. of concentrated sulfuric acid. Warm the mixture and hold a drop of water on a glass rod in the gas formed in the tube (?). The cloudiness of the water is due to the formation of orthosilicic acid, H_4SiO_4. Write three equations: (1) the formation of H_2F_2 by the action of sulfuric acid on the fluorid; (2) the formation of silicon tetrafluorid, SiF_4, by the action of hydrofluoric acid on silica; (3) the formation of orthosilicic acid and fluosilicic acid, H_2SiF_6, by the action of the tetrafluorid and water. The test of a substance for silicon must be carried out in a platinum dish, since glass and porcelain contain silicon. The procedure is also used as a test for fluorin.

III. **Solubility of Glass.** Grind a piece of chemical glass in the mortar. **Protect the eyes.** Moisten the powder with a little water and add a drop of phenolphthalein (?). Explain the result by the hydrolysis of potassium silicate, K_2SiO_3, which is an ingredient of chemical glass. How does this action show the solubility of glass?

IV. **Preparation of Sodium Silicate.** (Optional.) Mix thoroughly one part of clean sand with five parts of sodium carbonate. Heat the mixture strongly in a covered iron crucible. Heat until a clear liquid is obtained. Let the crucible cool, and extract the material with water and filter. Test the filtrate for a silicate (?).

V. **A Silica Garden.** (Instructor.) Fill a large beaker with a solution of sodium silicate (1 : 3 aq., specific gravity 1.1). Drop crystals of the sulfates of copper, iron, nickel, manganese, zinc, and the nitrates of cobalt and cadmium into the solution, and allow it to stand quietly over night (?).

Study Topics: Sadtler, *Chemistry of Familiar Things*, pp. 300–313.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.
Materials: Borax, filter paper, alcohol, turmeric paper, lime-water, food products.

I. **Preparation of Boric Acid**. Heat 40 cc. of water to boiling in a beaker, and dissolve 10 g. of borax in it. Pour 6 cc. of concentrated hydrochloric acid into the solution, stir, and allow the mixture to cool (?). Filter the crystals from the liquid, and wash them with a little water. Complete the equation. Use the product in Part II.

$$Na_2B_4O_7 + ?\,HCl + \ ?\,H_2O \longrightarrow 4\ H_3BO_3 + ?.$$

II. **Tests for Boric Acid**. (*a*) Dissolve a pinch of boric acid crystals in 3 cc. of alcohol in the evaporating dish. Wet a small shred of asbestos in the solution and hold it in the burner flame (?). Ignite the solution in the dish (?). The color of the flame constitutes the test.

(*b*) Dissolve some borax or boric acid in water and add a few drops of hydrochloric acid. Dip two strips of turmeric paper in the solution, and dry by pasting them on the outside of a beaker full of hot water. When dry, note the color (?). Wet one piece of test paper with ammonium hydroxid. Note the color (?). When this paper is dipped into some dilute hydrochloric acid, the original red color returns. Dry and compare (?).

III. **Borax and Boric Acid as Food Preservatives**. Borax is a valuable cleansing agent and has antiseptic properties in common with boric acid. The chief use of the latter depends on this property. Some food manufacturers use borax or boric acid in the preparation of foods, claiming that it is necessary. This necessity could arise only from careless and unsanitary methods of preparation, or to prevent loss from the careless handling or long keeping of the food products. Foods will keep without preservatives if they are prepared in a sanitary manner. The State of Illinois and some other states have forbidden the use of boric acid or its derivatives in the preparation of foods. If used, the National Pure Food Laws require that this fact shall be stated on the label. READ THE LABEL. " The common symptoms observed after long-continued doses of borax or boric acid in food are headaches, sensations of fullness in the head, uneasiness and nausea in the stomach, and disturbances of the digestion and appetite " (Wiley). These preservatives are likely to be found in fancy crackers (saltines) and biscuits, ice cream cones, sausage and chopped meats, canned meats and fish.

The turmeric test, given above, is best. Organic matter must be removed first. In the evaporating dish mix about 10 cc. of lime-water with about 10 g. of the food to be tested, and heat the dish gradually until it is red hot. Heat meats in the hood. When the black carbon is gone and the dish has cooled, add a few drops of hydrochloric acid and enough water to dissolve the ash. Test this solution with turmeric paper as above. Record in the form of report of chemical analysis.

REPORT OF CHEMICAL ANALYSIS

EXPERIMENT

Partial analysis of ....

Obtained from ....

Manufactured by ....

Exact name of brand ....

Tested for ....

Description of test ....

....

....

....

....

Equations of test ....

....

....

Found ....

....

Quality of sample ....

Remarks ....

....

....

I hereby certify that the above report is correct to the best of my knowledge.

(Signed) ....

Date ....

IV. Tests for Metals with Borax Beads. The oxids of certain metals impart characteristic colors to beads of borax, $Na_2B_4O_7$, or microcosmic salt, $NaNH_4HPO_4$, which act as a flux, dissolving the metallic oxid. The latter is best in most cases except for nickel. Make a 3-mm. loop with the platinum wire by bending the wire at the end until it touches the wire further back nearer the handle. Heat the wire red hot, keeping it out of the inner green cone of the flame, and dip it quickly into the powdered borax or microcosmic salt. Heat strongly in the flame until the bead becomes clear. While hot touch the bead to a very small particle of a salt of one of the metals given below and heat again until the color is uniform. If the bead is black, too much of the salt was used. Try again. To remove the bead dip it, while red hot, into water.

Make beads from a salt of each metal to learn the colors. Ask the instructor for an unknown salt to test by the bead method. Record results.

Metal	Colors	
	Oxidizing Flame	Reducing Flame
Cobalt	Blue	Blue
Chromium	Yellow-green	
Iron	Colorless to yellow	Bottle-green
Manganese	Amethyst	
Nickel	Red-brown	
Silicon	Insoluble particles of silica, called a “skeleton” Visible when the bead is red hot	

BLOWPIPE ANALYSIS[1]

This form of analysis is simple and short, but not very conclusive. Some of the tests are not distinct, so careful judgment is necessary to make a correct analysis. In case of doubt the analyst should obtain a sample of a salt of the same composition that the unknown is suspected to be, and compare the results of tests of the two substances. Other tests with the solution of the unknown substance may be made to confirm or correct the opinions already formed. For a complete systematic analysis follow the outline on pages 205–222.

I. **Detection of Base-forming Elements, Metals.** *A*. Moisten the substance to be tested with hydrochloric acid, and hold it in the outer edge of the burner flame by means of a clean iron or platinum wire. The important colors are:

1.	Violet, or through blue glass reddish violet	Potassium
2.	Yellow, brilliant, invisible through blue glass	Sodium
3.	Yellowish green	Barium
4.	Orange to brick-red	Calcium
5.	Scarlet, color very persistent	Strontium
6.	Carmine red, color vanishes quickly	Lithium
7.	Green, moistened with HCl, bluish	Copper
8.	Green only when wet with acid (sulfuric)	Borates

B. Grind a portion of the assay in the mortar with about four times as much sodium carbonate. Put a **small** quantity of the mixture in a shallow hollow in a piece of charcoal and heat with the reducing flame of the blowpipe. Do not mistake ash for coating.

1.	A coating on the charcoal, sometimes very volatile	C.
2.	White coating, sometimes volatile or with odor	G.
3.	Metallic bead without coating on the charcoal	H.
4.	A gray, black, or green residue	I.
5.	A white, luminous residue, fusing with difficulty	K.
6.	A decided coating	M.

C.	1.	A bead, if it is formed it is metallic	D.
	2.	No bead	G.
D.	1.	The bead is brittle when cold	E.
	2.	The bead is malleable when cold	F.
E.	1.	Coating is bluish white and volatile	Antimony
	2.	Coating orange when hot, light yellow cold	Bismuth
F.	1.	Coating yellow, bead quickly formed	Lead
	2.	Coating very slight close to assay, yellowish when hot, white when cold	Tin
G.	1.	Very volatile, abundant fumes, garlic odor	Arsenic
	2.	Odor of ammonia when warmed with a base	Ammonium
	3.	Quickly reduced, volatile. Heated with 12 parts of Na_2CO_3 and 6 parts of charcoal powder in a tube, a sublimate of minute globules of mercury	Mercury
H.	1.	Bead white	Silver
	2.	Bead red	Copper
	3.	Bead yellow	Gold

---

[1] Adapted from *Notes on Blowpipe Analysis*, Knight.

I. A minute portion of the assay heated in the oxidizing flame with a borax bead on platinum wire forms:

1. A blue bead when hot or cold .... Cobalt
2. An amethyst bead, black with excess amount .... Manganese
3. A green bead hot, bluish green when cold .... Copper
4. A reddish brown bead when cold .... Nickel
5. A yellow to green bead when cold .... Iron
   Confirm by testing a solution of the assay with potassium ferrocyanid. A blue precipitate shows iron.
6. A green bead hot, emerald green when cold .... Chromium
7. Insoluble particles moving about in hot bead .... Silicates

K. When the previously heated assay is moistened with a few drops of a solution of cobalt nitrate and re-heated forms:

1. A blue mass .... Aluminum or a phosphate
2. A slightly pink mass, not always pronounced .... Magnesium
3. A green residue .... Zinc

II. **Detection of the Acid-forming Elements, Non-metals.** *A*. Heat a little of the substance in a test-tube with a few drops of concentrated sulfuric acid. A very little of the substance should be taken first, since chlorates form an explosive gas with this acid.

1. A colored gas .... B.
2. A colorless, pungent gas .... D.
3. A colorless, inodorous gas. (It may be poisonous.) .... E.
4. No reaction .... F.

B. 1. The color of the gas, if red or reddish brown .... C.
2. The color of the gas is violet and colors starch. Wet a stirring rod, roll it in powdered starch, and hold it in the gas .... Iodid
3. A yellow explosive gas with odor of chlorin .... Chlorate

C. 1. The gas colors starch orange .... Bromid
2. Gas does not color starch .... Nitrate or Nitrite

D. 1. Gas forms white clouds with ammonia .... Chlorid
2. Gas fumes much and etches the glass of the tube. A drop of water on rod held in gas turns milky .... Fluorid
3. Gas having the odor of " ripe " eggs .... Sulfid
4. Gas having the odor of burning sulfur, SO_2 .... Sulfite
5. Gas having the odor of sulfur dioxid and the liquid becoming yellow with the separation of sulfur .... Thiosulfate

E. 1. Vigorous effervescence with HCl, turns lime-water milky .... Carbonate
2. Inflammable gas, CO, blue flame .... Oxalates
3. Acid blackens .... Organic matter

F. 1. Mix the substance with equal parts of sodium carbonate and borax, and fuse it on charcoal with the blowpipe. Place the fused mass on a clean silver coin and moisten it. If a brown or black stain forms .... Sulfate

2. Moisten with sulfuric acid and hold in the flame by a wire. A bluish green color shows .... Phosphate

If sodium or potassium phosphate is present, a blue mass forms with cobalt nitrate as before stated.

G. Tests for organic acids.

1. Acetates. To some of the substance in a test-tube add 2 cc. of ethyl alcohol and the same volume of concentrated sulfuric acid. Warm the mixture. A pleasant fruity odor shows acetates present.
2. Tartrates. Heat some of the substance in the evaporating dish. As the material blackens, an odor resembling burnt sugar will be noted.

BLOWPIPE ANALYSIS

EXPERIMENT

Write the purpose of this experiment with the proper heading, obtain an unknown substance from the instructor, and begin the tests, following the above plan of analysis. In the following form **record every test immediately after it has been made.** The instructor will state how many analyses are to be made.

ILLUSTRATION

No.	Substance	Reagent	Observations	Conclusion
1	unknown 1	flame	no color	—
2	unknown 1	blowpipe	brown coating	Cadmium
3	unknown 1	H_2SO_4	pungent gas	HCl?
4	unknown 1	Ammonia	white clouds	Chlorid
	Unknown #1 Cadmium chlorid, $CdCl_2$			

The above form of analysis shows something of how metals are separated from their ores. Most of those metals that are detected by reducing the salt on charcoal are separated from the ores in commercial practice by similar methods. The metals are not active, so the compounds are not very stable and can be decomposed by heat. The carbon (coke) combines with oxygen and sets the metal free. Most of the ores not oxids are roasted to convert them to this form before reducing with carbon.

Other metals are too active to permit of the decomposition of their compounds by heat. Electrolysis has become the best plan of separating these metals from their ores. It is interesting to note that very few of the metals admit of detection in more than one of the above methods.

BAKING POWDER

This product originated as a means for the quick aëration of bread. Its most essential ingredient is the sodium bicarbonate, from which comes the carbon dioxid to " lighten " the bread. Some acid principle is also present to liberate the carbon dioxid at the proper time under the combined influence of heat and moisture. The mixture is unstable, and the atmospheric moisture causes its premature chemical action, so it should be kept only in air-tight containers. Starch is always used as an ingredient which tends to keep the mixture dry.

The different kinds of baking powders vary according to the kind of acid agent present. All kinds may be grouped into three classes :

1. **Cream of Tartar.** In these powders, potassium bitartrate, cream of tartar, is used as the acid principle. Some of them also contain tartaric acid. The equation for the main chemical action is

$$\mathrm{HNaCO_3 + HKC_4H_4O_6 \rightleftharpoons KNaC_4H_4O_6 + H_2O + CO_2}\uparrow.$$

2. **Phosphate.** In these powders, an acid phosphate, usually calcium superphosphate, is the acid principle. The equation is

$$\mathrm{2\ HNaCO_3 + H_4Ca(PO_4)_2 \rightleftharpoons HCaPO_4 + HNa_2PO_4 + 2\ H_2O + 2\ CO_2}\uparrow.$$

3. **Alum.** In alum powders, the acid principle is one of the alums, either potassium, sodium, or ammonium alum. Although the alums are normal salts, their water solutions are quite strongly acid on account of the hydrolysis by water into aluminum hydroxid and sulfuric acid. The sulfuric acid then acts upon the sodium bicarbonate. The complete equation is

$$\mathrm{6\ HNaCO_3 + K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24\ H_2O \rightleftharpoons 2\ Al(OH)_3 + 3\ Na_2SO_4 + K_2SO_4 + 24\ H_2O + 6\ CO_2}\uparrow.$$

Some kinds of baking powder are mixtures of two or even all three of these classes. These acid agents vary in activity, and powders containing alum or calcium phosphate require a smaller weight of these agents to liberate the carbon dioxid. This allows more room for a filler of starch or other cheap material. This is the main reason for the variation in price of different powders. The phosphate and alum are also cheaper per pound than cream of tartar.

Much has been said about the harmful products in certain baking powders. There is no conclusive evidence of harmful quality for any one powder, although some states prohibit by law the sale of alum powders. The by-products of the chemical action in every case may be classed as drugs. Injurious results are more likely to follow the excessive use of baking powder bread, irrespective of the brand used, than the occasional use of the so-called worst kind. Bread made from yeast is generally recognized as being much more wholesome.

The powders which do not conform to the standards prescribed by law are mainly those that are deficient in available carbon dioxid. They may have been carelessly made, or have " lost strength " by reason of age. See method of determining the available carbon dioxid on page 159.

References: Report of State Food Commission (free). Leach, *Food Inspection and Analysis*, pp. 332–339.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Materials: Baking powder (2 tablespoonfuls, from home), resorcin, solutions of ammonium chlorid, barium chlorid, ammonium oxalate, ammonium molybdate, silver nitrate, iodin, logwood, and ammonium carbonate.

I. **Tests for Alum Powders.** (*a*) Aluminum. Heat about 2 g. of the powder to redness in an evaporating dish, cool, and add about 5 cc. of hot water to the residue. After stirring awhile, filter, and discard the residue, for any aluminum present will be contained in the filtrate as sodium aluminate. Add enough ammonium chlorid solution to the filtrate to give a distinct odor of ammonia (?). A white, gelatinous precipitate of aluminum hydroxid indicates the presence of aluminum compounds in the baking powder.

(*b*) Ammonia. Place a pinch of the baking powder in a test-tube, add sodium hydroxid solution, and warm carefully. Test the first of the escaping vapors with moist red litmus paper, and also note the odor (?). A change in the litmus paper and the odor of ammonia indicate the presence of ammonium salts in the baking powder.

(*c*) Sulfates. Place about 5 g. of the baking powder in a flask and add about 30 cc. distilled water (?). Shake till the action is finished, then filter, and use the filtrate in Parts I (*c*), II, and III. Use the residue on the paper for Part IV.

To a little of the filtrate add barium chlorid solution. A white precipitate, insoluble in hydrochloric acid, — test it, — indicates sulfates.

Any of these tests, or all of them, indicate an alum powder.

II. **Tests for Phosphate Powders.** (*a*) Calcium. To 5 cc. of the filtrate from Part I (*c*) add a little freshly prepared ammonium oxalate solution, $(NH_4)_2C_2O_4$. A white precipitate, insoluble in acetic acid, — add some, — but soluble in hydrochloric acid, — pour some into the same tube, — indicates calcium as calcium phosphate.

(*b*) Phosphates. To a mixture of equal volumes of sp. gr. 1.20 nitric acid and ammonium molybdate solution, $(NH_4)_2MoO_4$ (total volume about 4 cc.) add a few drops of the filtrate from Part I (*c*). Warm the liquid nearly to boiling (?). A bright yellow precipitate indicates phosphates.

III. **Tests for Tartrate Powders.** (*a*) If in the test for aluminum in Part I (*a*) the powder smells of burnt sugar while being charred, tartrates are indicated. If phosphates are not present, tartrates may be tested for by adding 2 to 3 cc. silver nitrate solution and a little very dilute ammonium hydroxid to 5 cc. of the filtrate from Part I (*c*). Warm gently (?). Tartrates will reduce the silver nitrate to metallic silver, forming a silver mirror on the test-tube.

(*b*) If phosphates were found present in Part II (*b*), the test for tartrates must be made as follows: Evaporate the rest of the filtrate from Part I (*c*) to dryness on the water-bath. Put the residue into a test-tube, add the same amount of dry resorcin and a few drops of concentrated sulfuric acid. Heat gently (?). A red color indicates tartrates.

IV. **Starch.** Put a few drops of a solution of iodin in potassium iodid solution on the filter paper used in Part I (*c*) (?). What is the residue on the paper?

V. **Detection of Alum in Pastry.** Make a paste of the sample with water in a mortar. Add 2 cc. of logwood solution and an equal amount of ammonium carbonate solution. A change of color to lavender or blue, not disappearing with boiling, indicates alum. With no alum, the color varies from pink to red.

VI. **Baking Powder Formulas.** The following formulas give approximate proportions for two kinds of baking powders. Look up the cost per pound of each constituent in a trade catalogue and determine whether it would be economical or not to prepare baking powder for use in the home.

TARTRATE POWDER	COST	PHOSPHATE POWDER	COST
2 lb. cream of tartar 1 lb. sodium bicarbonate 1 lb. corn starch		1.75 lb. calcium acid phosphate 1 lb. sodium bicarbonate 1 lb. corn starch	

Questions: 1. What one ingredient only is needed where **sour** milk is used to make bread?

2. Is it practical and profitable to make baking powder in the home? Give reasons.

3. Could other gases than carbon dioxid be used to leaven bread? Why?

REPORT OF CHEMICAL ANALYSIS

EXPERIMENT

Partial analysis of...

Obtained from..

Manufactured by...

Exact name of brand...

Tested for...

Description of test..

....

....

....

....

Equations of test...

....

....

Found...

....

Quality of sample...

Remarks..

....

....

I hereby certify that the above report is correct to the best of my knowledge.

(Signed)...

Date..................................

EXPERIMENT

Purpose: State it on another sheet used for the record.

Apparatus: Dry cells or storage batteries or D. C. generator, electrolysis apparatus, nickel crucible, electroplating jar.

Materials: Metals: sheet zinc, copper, lead, nickel, mercury, nickel ammonium sulfate, solutions of potassium ferrocyanid, copper sulfate, zinc sulfate, mercuric chlorid, magnesium sulfate, tin chlorid, silver nitrate, cobalt chlorid, nickel chlorid, and lead nitrate.

I. **Production of Energy by Chemical Action (Exothermic).** (*a*) Place a small strip of sheet zinc in dilute hydrochloric acid in a half filled 10-cm. test-tube. Note the evolution of a gas. Identify it by test (?). Feel the tube and note any change of temperature of the liquid (?). The zinc goes into solution forming zinc ions (Zn^{++}), and liberates, besides the gas, energy that seems to be responsible for the chemical action.

(*b*) Place a zinc and also a bright copper strip in dilute hydrochloric acid in the smallest beaker, not letting the strips touch each other. Examine each strip closely and describe results. Touch the upper end of each strip separately with the tongue (?), and then both of them at the same time (?) while the lower ends are apart in the acid. A definite taste is due to the passing of an electric current from one metal to the other. Press the upper ends of the strips together and describe the results on each strip. Allow the strips to remain in contact until it is apparent which strip is being acted on chemically (?). Note any temperature change as before (?). Test small portions of the acid solution for zinc and copper ions (?). A white gelatinous precipitate formed with sodium hydroxid solution and dissolving in an excess of the reagent is the test for zinc. A brown precipitate formed with potassium ferrocyanid solution is the test for copper.

Questions: 1. What metals replace hydrogen from its compounds? (See the activity series of metals, page 256. Write the symbols.)

2. What must be the magnitude of the solution tension of the metals which do not replace hydrogen? The force tending to drive the atoms of a metal into the ionic form is called solution.

3. Why does the acid solution attack one metal and not the other?

4. According to the ionization theory, how does the current pass through the acid solution from one plate to the other?

5. Why do hydrogen bubbles form on the copper strip when the strips are touched?

6. Why is the chemical action hastened when the two strips touch?

7. Why should there be less heat given off when the strips touch than when they do not?

II. **Electromotive Series of Metals (by Replacement).** When a metal is placed in a solution containing ions, some of its atoms pass into solution as positively charged ions, at a rate depending upon the solution tension of that metal, leaving an equivalent negative charge on the metal remaining. This accumulated negative charge is able to attract to itself and neutralize the positive charges on ions of metals having a less solution tension. When these ions have lost their charges they are displaced from solution as the free metal.

Put 5 cc. each of copper sulfate and zinc sulfate solutions in separate tubes. Place a bright strip of lead in each solution and let them remain for 15 minutes. Examine each strip for deposited metal and describe the result (?). In case a negative result is obtained reverse the order of procedure. If the lead does not replace zinc from solution try a strip of zinc in a solution of a lead salt. Always have the metals clean and bright. Write the three metals in the order of replacement, the one replaced following the one replacing it.

Put a strip of copper in a solution of mercuric chlorid and observe results as before (?). Place mercury in order with the list of three.

Proceed in like manner, placing zinc in a solution of magnesium sulfate (?). If the result is negative, put magnesium in zinc sulfate (?). If the time permits, try zinc in a solution of tin chlorid, and as many other combinations as are necessary to establish the position of tin in the series (?). Similarly place silver, cobalt, and nickel (?).

III. **Chemical Action Caused by Electrical Energy.** Any compound that dissociates into ions can be decomposed by electricity. The charged electrodes placed in a solution of such a substance attract the ions of unlike charge and draw them to it. When the ions come into contact with the electrodes, their electric charges are neutralized by the opposite kind of electric charge on the electrode. Whenever this happens chemical changes take place.

(*a*) Fill a Hofmann's electrolysis apparatus or its equivalent with dilute sulfuric acid, and pass through the acid a direct current of electricity (2 or more dry cells or storage cells may be used). Record what is observed at the surface of each electrode. Identify by suitable tests the products formed at each electrode (?). Review page 78.

$$(1)\ H_2SO_4 \rightleftarrows H^+ \cdot H^+ + SO_4^{--}. \qquad (2)\ \underset{\text{(ions)}}{2\overset{+}{H}} + \underset{\text{(charges)}}{2\ominus} \longrightarrow \underset{\text{(molecules)}}{H_2}.$$

$$(3)\ SO_4^{--} + 2\oplus \longrightarrow SO_4. \qquad (4)\ 2SO_4 + 2H_2O \longrightarrow 2H_2SO_4 + O_2.$$

In a similar manner write the equations of the electrolysis of solutions of nitric acid, HNO_3, hydrochloric acid, HCl, sodium sulfate, Na_2SO_4, sodium chlorid, $NaCl$, sodium nitrate, $NaNO_3$, sodium hydroxid, $NaOH$, copper chlorid, $CuCl_2$, copper nitrate, $Cu(NO_3)_2$, copper sulfate, $CuSO_4$, silver nitrate, $AgNO_3$, silver sulfate, Ag_2SO_4.

(*b*) Electrolysis of fused sodium hydroxid. Place about 100 g. of solid sodium hydroxid in a nickel or iron crucible and connect the crucible to the positive terminal of a series of three or four dry cells. Melt the sodium hydroxid and then touch its surface with a wire connected to the negative terminal of the cells (?). Globules of sodium will form. Describe them. Write the equation for the electrolysis of molten sodium hydroxid.

(*c*) Electroplating with nickel. In order that the deposited metal may adhere firmly to the article, the latter must be made smooth and perfectly clean. It should be rubbed with suitable abrasives (powdered emery, pumice, etc.) until all roughness and rust are removed, then dipped into a hot solution of sodium hydroxid, 75 g. per liter, and thoroughly rubbed with a stiff brush to remove the last traces of grease from the article. After rinsing in clean water, it is ready for the plating solution. It should not be touched by the hands after the rinsing. The least trace of grease will prevent deposition of metal or will cause the metal coating to strip off when rubbed. Articles to be replated should have the old coating removed by making it the anode in an old plating solution containing free cyanide (violent poison). Acids may be used if care is taken to prevent injury to metal underneath the plating.

Dissolve 80 g. of nickel ammonium sulfate in 1 L. of distilled water in the tall electroplating jar. Connect the battery so that the nickel metal will be the anode. Never remove the anode unless the solution is to be renewed. An electromotive force of 5 or 6 volts is necessary at the start of the plating. After the article has been covered by a film of nickel the plating is best continued at about 2 volts to insure a tough, firm deposit. The article must never be lifted from the solution before the operation is finished.

Nickel adheres better to copper than to iron, so iron objects may receive a preliminary coating of copper from a solution prepared as follows: Add concentrated ammonium hydroxid to copper sulfate solution until the green precipitate first formed dissolves in an excess of the reagent. Then add potassium cyanid solution until the blue color of the copper solution changes to amber. Use a current of 6 to 8 volts. The nickel coating is never bright, so it should be rubbed with metal polish or burnished.

"HARDNESS" IN WATER

Most carbonate rock is insoluble in water; yet the rain water, with the aid of carbonic acid, annually dissolves many thousand tons of limestone rock. As the rain falls through the atmosphere it dissolves carbon dioxid and forms carbonic acid. The dissolved lime in drinking water furnishes some material for the bones of animals. The various forms of crustacea living in water use it to make their shells. In Nature's processes the dissolved lime performs desirable functions, but in the arts and industries it is troublesome and harmful.

This mineral matter in solution is commonly spoken of as "hardness." If the minerals dissolved are chiefly calcium and magnesium **bicarbonates** the water is said to have "temporary hardness." If the minerals are the **sulfates** of these same metals the water is "permanently hard." These minerals react with soap, used in washing, to form an insoluble soap which not only has no cleansing power, but because of its greasy properties makes the hands grimy, the clothes dingy, and the tubs and lavatories dirty and unsightly. This soap precipitated by the lime represents trouble and waste. The waste amounts to several dollars per year for each family. In a large city like Chicago this waste may amount to thousands of dollars annually, depending on the degree of hardness of the water. If water contains less than 25 parts of lime salts for each million parts of the water the hardness is hardly noticeable. If over 50 parts per million, it is decidedly hard, and over 100 parts per million makes very hard water.

When "temporary" hard water is heated in kettles and boilers the solvent carbonic acid is decomposed and driven off as carbon dioxid. The lime is then deposited in scaly form on the inside of the container. Since the scaly lining prevents the heat from reaching the water so readily, the iron becomes overheated and is more readily corroded by the oxygen outside and the water inside the boiler. There is a loss in heating power depending on the thickness of the scale.

Thickness of scale, inches.	$\frac{1}{64}$,	$\frac{1}{32}$,	$\frac{1}{16}$,	$\frac{1}{8}$,	$\frac{1}{4}$,	$\frac{1}{2}$,	$\frac{3}{4}$,
Loss of heat energy.	2%	4%	9%	18%	38%	60%	90%

The Mississippi River water, which contains 120 parts per million of scale-forming ingredients, if used in a 1000 horse power steam plant for 20 days of 24 hours each, would necessitate the labor of shoveling, scraping, and hammering one ton of scale from the inside of the boilers. Chemical engineers have devised methods of removing most of the dissolved minerals, using heat and chemicals to render the minerals insoluble **before** the water is introduced into the boilers or used in washing. The insoluble lime is filtered off and the water is ready for use. Boiler compound for treating the water inside the boiler is only a partial remedy. The cost of softening water is approximately only one-tenth of the expense that would follow if the minerals were not removed.

Study Topics: Martin, *Triumphs and Wonders of Modern Chemistry:* "The Great Caverns of the Earth," pp. 250–264.
Brownlee, *The Chemistry of Common Things:* "Hard Water," pp. 181–189.
Snell, *Elementary Household Chemistry:* "Hard Water," pp. 121–126.

TREATMENT OF HARD WATER

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Materials: Marble chips, litmus paper, lime-water, solutions of sodium carbonate, calcium sulfate, magnesium sulfate, and soap.

I. **Preparation of Temporary Hard Water.** Dilute about 10 cc. of lime-water with an equal volume of water and test its reaction with litmus paper (?).

Start the generation of carbon dioxid (Fig. 95) by pouring dilute hydrochloric acid on the marble chips in the generator. Pass the gas into the diluted lime-water in the test-tube(?). Recall the test for carbon dioxid. Write the equation for the formation of the precipitate.

Continue to pass the gas into the lime-water until the precipitate, first formed, disappears. Test the solution with litmus (?). The precipitate has been converted into the bicarbonate, $Ca(HCO_3)_2$ by the action of the carbonic acid formed. Write the correct equation (?). Save the solution for further experimental work. In waters naturally hard the bicarbonate of magnesium is usually present also. The chemical action with magnesium is identical with calcium.

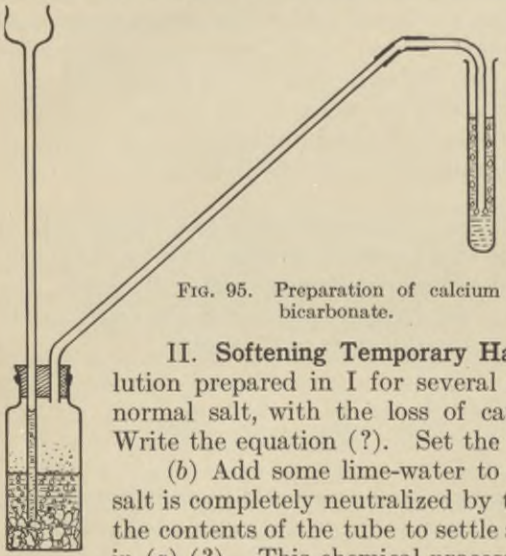


FIG. 95. Preparation of calcium bicarbonate.

II. **Softening Temporary Hard Water.** (*a*) Boil about 5 cc. of the hard water solution prepared in I for several minutes (?). The acid salt is changed back into the normal salt, with the loss of carbonic acid. Recall the unstable nature of this acid. Write the equation (?). Set the tube with contents aside to settle.

(*b*) Add some lime-water to about 5 cc. of the prepared hard water (?). The acid salt is completely neutralized by the calcium hydroxid. Write the equation (?). Allow the contents of the tube to settle and compare the amount of precipitate with that formed in (*a*) (?). This chemical process is known as the " lime process " devised by Clark.

Questions: 1. Why does boiling soften temporary hard water?

2. Why was a larger precipitate formed in process (*b*) than in process (*a*)?

III. **Softening Permanently Hard Water.** (*a*) Add some sodium carbonate solution to 10 cc. of calcium sulfate solution (?). The solubility of calcium carbonate is 0.0018, and that of the sulfate is 0.241. Explain the formation of the precipitate and write the equation(?). This is the " soda " process devised by Porter.

(*b*) Test the effect of this same reagent, sodium carbonate, on 5 cc. of the temporary hard water (?). The bicarbonate of sodium is formed and is soluble. Write the equation. Filter off the precipitate and add to the filtrate 2 cc. of the soap solution. Mix by shaking. Add the same volume of soap solution to 5 cc. of the untreated hard water, and mix as before. Compare the effect of the treated and untreated hard water on soap solution (?).

(*c*) Add sodium hydroxid to a dilute solution of magnesium sulfate (?). Magnesium hydroxid is precipitated. Filter the mixture and write the equation. Test the effect of the filtrate and also the untreated magnesium sulfate solution on equal volumes of soap solution. Compare the amounts of precipitated soap in each case (?).

Questions: 3. Why would the boiling treatment not soften permanently hard water?

4. What is the insoluble compound of calcium? Of magnesium?

5. State the importance of the property of solubility as shown in this experiment.

PLASTER OF PARIS, MORTAR, AND CEMENT

EXPERIMENT

Purpose: State it on a separate sheet used for the record.
Apparatus: Air bath, iron plate, blast lamp or other source of high temperature.
Materials: Powdered gypsum, plaster of Paris, marble chips, quicklime, Portland cement, sand.

I. **Plaster of Paris.** (*a*) Preparation. Spread 10 g. of powdered gypsum evenly on a piece of sheet iron and leave it 20 minutes in an air bath maintained at a temperature of 125° to 130° C. Gypsum is a hydrate having the formula $CaSO_4 \cdot 2\,H_2O$, and may be completely dehydrated at temperatures above 200° C., but the product is then worthless as a plaster. At 125° C. gypsum is changed into a lower hydrate, $(CaSO_4)_2 \cdot H_2O$. Describe the product and write the equation for its formation. Test its hardening qualities by mixing it with a little water (?).

(*b*) Setting. Mix two teaspoonfuls of the commercial plaster of Paris with enough water to make a thick paste, and pack it closely around some small object (previously greased) of which you wish to make an impression. When nearly hard, cut the plaster open with a knife so as to remove the object, but take care not to break the mold. The plaster of Paris has changed back to gypsum again. Describe the final product, write the equation for the hardening, and state what has become of the water used in making the paste.

II. **Mortar.** (*a*) Making the Quicklime. Heat a 5-mm. piece of marble as hot as possible for 15 minutes on a corner of the wire gauze. Compare the product with marble, particularly as to taste (cautiously) and action with water. The product is calcium oxid. Write the equation for the action. All carbonates except those of the alkali metals are decomposed by heat.

(*b*) Slaking the Quicklime. Pour a little boiling water on 25 g. of quicklime, and from time to time add more water as the action proceeds. At the end you should have a smooth, stiff paste. Test the product with litmus paper (?), describe and name it. Also describe the action and write the equation to represent it.

(*c*) Setting. Mix 50 g. of sand with the paste, adding more water if needed. Put half the mortar in a wide-mouth bottle and cork it tightly. Spread the rest of the mortar on a brick or board and let it stand. After a day or two compare the two samples of mortar (?).

(*d*) Theory of Setting. In separate test-tubes add dilute hydrochloric acid to some marble, the mortar exposed to the air, and also to some of the mortar in the bottle (?). If any gas is evolved in either case, hold a film of lime-water in a wire loop in the gas (?). Account for the difference between the two samples of mortar and write the equation for the hardening action.

III. **Cement. Theory of Setting.** The chemical changes taking place when cement hardens are not as well understood as are those involved in the hardening of plaster of Paris and mortar. One of the recent theories advanced is that the cement consists mainly of calcium silicate and calcium aluminate, and that when water is added the latter is hydrolyzed into calcium hydroxid and aluminum hydroxid. The hardening is thought to be due to the crystallization of the calcium hydroxid, binding all components together. The aluminum hydroxid is very gelatinous, so it fills up the pores and makes the hardened product quite impervious to moisture.

$$Ca_3(AlO_3)_2 + 6\,H_2O \longrightarrow 3\,Ca(OH)_2 + 2\,H_3AlO_3.$$

Mix a tablespoonful of Portland cement with two spoonfuls of clean sand (sharp-grained sand, called torpedo sand, gives a stronger product). Add water slowly with continual stirring to make a stiff paste. Make a small ball of a portion of the mixture, place it in a bottle, cork tightly, and let it stand until the next day (?). Mold the rest of the paste into any desired shape and let it stand until the next day (?). If a mold is used, it should be greased first and the paste firmly tamped into it with the pestle. Finish off the top surface by spreading over it a thin paste of cement and water.

Study Topics: Baskerville, *Municipal Chemistry:* "Cement and Concrete," pp. 470–496.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Materials: Solutions of lead nitrate, sodium carbonate, sodium sulfate, potassium dichromate, 25% sodium hydroxid; metals and soluble salts of zinc, tin (stannic), and aluminum.

I. **Lead as a Base Former.** Fill a clean 15-cm. tube half full of lead nitrate solution from the stock bottle, and pour it into 100 cc. of distilled water in a beaker. Put about 5 cc. of this solution into each of four tubes, and add to one a few drops only of a solution of sodium hydroxid (?). Add a few drops of sodium carbonate to another 5-cc. portion (?), and likewise sodium sulfate to the third portion (?), and potassium dichromate to the fourth.

The precipitates formed are insoluble compounds of lead. They show that lead ions, Pb^{++}, were present in the dilute solution of lead nitrate, so the precipitates are said to be " tests for lead." When in the form of the positive ion, the lead — or any other metal — acts as a base former and combines with negative ions. Write correct equations for each precipitate, having its formula last and followed by a word stating its color. The first three reactions are cases of double replacement. In the fourth, water takes part in the action, and potassium nitrate, nitric acid, and lead chromate, $PbCrO_4$, are formed.

II. **Lead as an Acid Former.** To 15 cc. of the diluted lead nitrate solution first prepared, add sodium hydroxid until the precipitate first formed is dissolved in the excess of the reagent. Add to separate portions of this solution a few drops of the test reagents for lead as given above (?).

Questions: 1. Is lead in the positive ion form in the solution last prepared? Give reasons.

2. In what ionic form must the lead be in this solution? Its composition is PbO_2^{--}, the characteristic ion of plumbous acid, H_2PbO_2. This acid is the lead hydroxid precipitated, which — in the presence of the much stronger base, sodium hydroxid, added in excess — reacts as an acid and is written after the manner of acids, hydrogen first. The sodium salt of this acid is formed. Write the equation, using correct valences.

Use potassium hydroxid in place of sodium hydroxid, and repeat the operations given above in Part II only (?).

Use ammonium hydroxid in place of sodium hydroxid in Part II (?). Explain its behavior by its weak basic properties.

III. **Aluminum, Zinc, and Tin as Amphoters.** (*a*) Precipitate the hydroxid of one of the three metals after the manner in Part II, divide the precipitate into two parts, and add sodium hydroxid to one part and hydrochloric acid to the other (?). If the precipitate dissolves in both reagents, it is shown to act as a base, forming a salt by action with hydrochloric acid; and also as an acid, reacting with sodium hydroxid to form a different salt. The sodium salts formed by the hydroxids of these metals acting as acids are Na_3AlO_3, Na_2ZnO_2, and Na_2SnO_3, the aluminate, the zincate, and the stannate, respectively. Test the hydroxid of the other two metals, writing all six equations.

(*b*) Heat a piece of aluminum in a tube with a 25% solution of sodium hydroxid. Test the gas liberated with a blazing splinter (?). If time permits, make the same tests with zinc and tin. Write the equations. The above-named salts of these elements are formed.

IV. **Other Amphoteric Elements.** Examine the formulas of the salts on the stock bottles and write the formulas of all having a metal in the negative radical or ion. The text may be consulted also.

Summarize the evidence of amphoteric properties as shown above, and define the term.

ALUMINUM

Aluminum is the most abundant metal. "Every clay bank is a veritable mine of it." 7.3 % of the earth is composed of this metal. If a cheap enough process for extracting it from clay could be discovered, it would rival iron as the most useful metal. Aluminum was discovered by Woehler in 1827. Deville of Paris extracted it from aluminum chlorid by reducing it with sodium and put it on the market at $90 per pound. A cheaper process of making sodium reduced the price finally to $4 per pound. It was still too expensive for common uses until modern electrical processes finally reduced the price to about 20 cents per pound.

Zinc is the active metal consumed in voltaic cells to produce a current of electricity. Aluminum ranks ahead of zinc in the activity series, so it may be considered a more efficient metal that may, in the future, be substituted for zinc in the electric cell. Its cost per pound is several times that of zinc, but its chemical equivalent (9) is less than a third of the equivalent of zinc (32.7). Its other properties are not so well suited to the present use in the electric cell as those of zinc.

At ordinary temperatures aluminum tarnishes (oxidizes) very little. At higher temperatures, when finely powdered, it acts quite differently. It burns and liberates an astonishing amount of heat energy. Under these conditions it displaces less active metals from combination with oxygen. On account of this property, a mixture of powdered aluminum and iron oxid is used in welding iron and steel by the "thermite" process. Iron is set free in the molten state, showing the great heat developed by this chemical change. This chemical operation is also utilized in reducing the oxids of manganese and chromium, which are difficult to reduce by the usual method of heating with carbon.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Apparatus: Hessian crucible.

Materials: Aluminum powder, iron oxid, magnesium ribbon, clay, potassium carbonate, aluminum sulfate, alizarin paste (25 %), congo red, muslin, sodium carbonate and sulfate, cochineal or indigo solution.

I. **Aluminum in the Thermite Process.** Mix thoroughly 4 g. of aluminum powder with 10 g. of powdered iron oxid (ferric). Put it in a Hessian crucible and support it over a large pan to catch any hot metal thrown out. Put one end of a piece of magnesium ribbon in the mixture and a little magnesium powder around it. This serves as a fuse. Cover the mixture 5 mm. deep with powdered sodium carbonate or calcium fluorid to keep the air, as much as possible, from the hot metal. Ignite the fuse and observe the action (?). After the residue has cooled examine it for metallic beads. Test them for iron with a magnet. Put some of them in a tube with hydrochloric acid and test the liberated gas with a flame (?). Write the equations.

II. **Uses of Aluminum Hydroxid.** (*a*) Clarifying Water. To 5 cc. of aluminum sulfate solution or the same quantity of any soluble aluminum salt add ammonium hydroxid until a precipitate is formed. Collect it on a filter and state its properties (?). Write the equation.

Shake a tubeful of water with a little finely pulverized clay to make it turbid. Pour half of it into another tube and set it aside to settle. To the other half add about 3 cc. of aluminum sulfate solution and enough ammonium hydroxid to precipitate the hydroxid of aluminum. Set the tube aside to settle with the other one. Note which one becomes clear at the top first (?).

Repeat the above procedure, but this time use enough cochineal or indigo solution to color the water faintly in place of the clay (?). The particles of dyestuff are much smaller than those of the clay, but the hydroxid collects them and forms what is known in the dyeing industry as a "lake."

Questions: 1. Considering the properties of aluminum hydroxid, why should it make the clay particles settle more rapidly?

2. How could this compound be used to make cloth take up dyestuff from solution and hold it?

(*b*) Dyeing Cloth. Remove the sizing from three, 4 × 6 cm. pieces of white muslin, by boiling them for two minutes in a 1 % solution of hydrochloric acid. Rinse well in water containing a little ammonium hydroxid to neutralize any traces of acid. Let each pupil put one piece of the cloth in a congo-red bath prepared by dissolving half a gram each of congo red, sodium carbonate, and sodium sulfate in 400 cc. of distilled water in an enameled dish or large evaporating dish. Boil the cloth in the dye for two minutes, then let each pupil remove a piece, wash it well in water, and dry between blotting papers. This dye spots easily with acid, so care should be taken in handling it. It is not very permanent and soon " saddens " in the air. It is used as an illustration of a " direct " or substantive dyestuff for cotton without a mordant.

Dilute some alizarin paste in an enameled dish until it is fairly liquid. Let each pupil put a piece of cloth in the dye and follow the procedure as for the congo red. The last piece of cloth should be soaked in a solution of aluminum sulfate and then placed in dilute ammonium hydroxid. By this treatment the cloth is mordanted. After all other pieces of cloth have been removed from the alizarin dye the mordanted pieces should be boiled for two minutes, then washed and dried.

Questions: 3. Is alizarin a direct dye for cotton?

4. Explain why the mordanted cotton has a much deeper color.

III. **Preparation of Alum from Clay.** Put about 30 g. of dry, pulverized clay in an evaporating dish and stir into it about 20 cc. of concentrated sulfuric acid. Heat the mixture in the hood for 15 minutes, keeping the temperature low enough so that little white fumes come from the acid. After the mixture is cooled pour it into a beaker and mix about 80 cc. of hot water with it. Stir it well, let it settle, then decant the liquid into another beaker. Treat the residue with a smaller quantity of boiling water than before, stir it well, and after settling, decant the liquid into the first portion. Treat the residue, as before, with a still smaller quantity of water and filter all the decanted liquor. Heat the filtrate to boiling and add slowly, while stirring, 8 g. of potassium carbonate. Evaporate the liquid to half of its volume or less, and let the dish remain quietly for a day. If no crystals are found, evaporate more water and set the dish aside again. When crystals are found, remove, and dry them on a filter paper. Bring the crystals to the instructor. Complete the equation

$$2\,HAlSiO_4 + 3\,H_2SO_4 \longrightarrow Al_2(SO_4)_3 + ?\,SiO_3 + ?.$$

This equation shows the nature of the action, but the alum will contain some potassium sulfate.

IV. **Cleaning Silverware by Means of Aluminum.** (At Home.) (*a*) The Problem. The greatest inconvenience in the use of silverware is that it tarnishes so readily, especially with sulfur and certain of its compounds. The tarnish is chiefly the dark colored silver sulfid.

(*b*) The Method. Dissolve a teaspoonful each of sodium carbonate ($Na_2CO_3 \cdot 10\,H_2O$) and sodium chlorid in a quart of water. Heat it to the boiling point and put it in a suitable aluminum vessel. Dip the tarnished silver articles into the solution, letting them touch the metal vessel. The tarnish will be removed in one minute. The so-called coating of oxidized silver on some tableware is silver sulfid, so it will be removed by this process.

(*c*) The Explanation. Referring to the activity series of the metals, it may be noticed that aluminum stands considerably higher than silver in the series. If strips of aluminum and silver are placed in a solution of an acid, base, or salt — sodium chlorid and carbonate, in this case — and are allowed to touch each other, an electrical cell is formed, the aluminum forming the positive plate and the silver the negative. The electricity is produced by the chemical action at the aluminum plate, and is carried by the sodium ions through the solution to the silver plate. From the silver plate the electricity passes through the point of contact back to the aluminum. When the sodium ions reach the silver, they discharge their positive electricity upon it and become electrically neutral atoms which immediately act upon the dissolved silver sulfid surrounding the silver object, replacing the silver which is precipitated upon the silverware. The silver sulfid is only slightly soluble in such a solution, but as fast as that dissolved is acted upon by the sodium, more dissolves, until all of the silver sulfid has been removed.

There is practically no loss of silver by this method. Aluminum goes into solution and is lost. Zinc may be used instead of the aluminum, and is fairly effective, but it becomes coated with zinc carbonate and ceases to give results unless the coating is removed. Hydrochloric acid may be used.

IRON COMPOUNDS

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Apparatus: Hydrogen sulfid generator.

Materials: Iron picture wire, zinc, hydrogen peroxid, chlorin water, solutions of potassium ferricyanid, ferrocyanid, and sulfocyanate, ammonium hydroxid, potassium permanganate and dichromate, 30% ferric chlorid, oxalic acid saturated.

I. **Ferrous Compounds.** (*a*) Preparation. Place a loose wad of iron picture wire in a test-tube, cover it with water, and add about half as much concentrated hydrochloric acid. Keep the action going briskly for five minutes, adding more acid if necessary. Then pour the contents of the tube into 50 cc. of boiling water in a flask. Boil the liquid several minutes longer and then cork it loosely, removing the cork only when pouring out some of the solution. Note the color of the solution — it is characteristic of the ferrous compounds in solution — and write the equation for the formation of ferrous chlorid, $FeCl_2$.

(*b*) Tests. To 3 cc. of the ferrous chlorid solution in a small test-tube add dilute ammonium hydroxid solution until the odor of the ammonia persists after shaking. Describe the precipitate, name it, and write the equation for its formation. Filter and spread out the paper with the precipitate, saving it for Part II (*b*).

Add about 1 cc. of potassium ferrocyanid solution to 3 cc. ferrous chlorid solution in a small test-tube, noting the properties of the precipitate, naming it, and writing the equation. The negative ion of the potassium ferrocyanid, $K_4Fe(CN)_6$, is $Fe(CN)_6$, and the valence is minus four.

Repeat the test, using potassium ferricyanid solution instead of the ferrocyanid. Its formula is $K_3Fe(CN)_6$, and it ionizes thus:

$$K_3Fe(CN)_6 \rightleftarrows 3\ K^+ + (Fe(CN)_6)^{---}.$$

Repeat the test with potassium sulfocyanate instead of the ferrocyanid. The result should be negative, any change being due to the presence of ferric chlorid. Record the results of these tests in a tabular form like that below.

II. **Ferric Compounds.** (*a*) Pour about one-third of the remaining ferrous chlorid solution into a 15 cm. test-tube; add 1 cc. each of concentrated hydrochloric acid and nitric acid. Boil the liquid about two minutes. Note the color of the ferric chlorid solution produced — characteristic of ferric solutions — and complete the following equation:

$$?\ FeCl_2 + ?\ HCl + ?\ [O] \longrightarrow ?\ FeCl_3 + ?.$$

The oxygen is written [O] because it is not actually present as gaseous oxygen, but is used up as soon as the nitric acid is decomposed.

(*b*) Tests. Repeat all the tests of Part I (*b*), using the ferric chlorid solution instead of the ferrous chlorid. If the precipitate caused by ammonium hydroxid is green, the action above was incomplete and should be repeated before making these tests. Name and describe precipitates, and write equations. Compare the precipitate saved from Part I (*b*) with that produced by the ammonium hydroxid and ferric chlorid. Explain the change.

TEST FOR IRON SALTS

IRON SALTS	POTASSIUM FERROCYANID	POTASSIUM FERRICYANID	POTASSIUM SULFOCYANATE	AMMONIUM HYDROXID
Ferrous chlorid				
Ferric chlorid.......				

III. **Reduction of Ferric Compounds.** (*a*) By Hydrogen. Add 1 cc. concentrated hydrochloric acid to 3 cc. ferric chlorid solution; drop in a few small pieces of mossy zinc. Keep the action going vigorously in the tube until the " ferric " color has disappeared, adding more acid or zinc as needed, or warming slightly. Divide the solution into two parts, and test each for ferrous ion, using the two tests most characteristic for that ion. State the results of the tests, and explain the change produced by the hydrogen, giving equation.

(*b*) By Hydrogen Sulfid. Pass hydrogen sulfid gas slowly through 3 cc. of ferric chlorid in a small test-tube till the " ferric " color has disappeared. Test the solution for ferrous ions (?), explain the change produced by the hydrogen sulfid, and name the precipitate formed (?). Remember the instability of hydrogen sulfid.

IV. **Other Oxidizing Agents.** Pour 3 cc. portions of ferrous chlorid solutions into five test-tubes, add a little dilute hydrochloric acid to each, and then add one of the following reagents to each tube: hydrogen peroxid solution, chlorin water, potassium permanganate solution, potassium dichromate solution, and air (blow through a glass tube several minutes). Test the resulting liquids for ferric ions by the best reagent for that ion (see tabulation of tests) (?). State the results and write the equations for the actions. All the metals are left as their chlorids (see valences in the appendix). The action may be explained by assuming that the oxidizing agent and the hydrogen chlorid give water and chlorin. Part of the chlorin combines with the metals of the oxidizing agent, according to their valences, while the rest of the chlorin combines with the ferrous chlorid, changing it into ferric chlorid. The method of writing and balancing the equations is quite similar to the method used in the experiment on chlorin, and the introduction to that experiment should be reviewed. Write the equations for the changes which have taken place.

V. **Blue Prints.** Mix 10 cc. each of 30 % ferric chlorid solution and saturated oxalic acid solution in a darkened room, and divide into three portions in test-tubes. Expose one portion to direct sunlight for several minutes and record the result. Add a few drops of a saturated solution of potassium ferricyanid and record the result.

Add a few drops of the potassium ferricyanid solution to the second unexposed portion (?). Compare the result with that obtained with the exposed portion, and state the change which has been produced in the exposed portion. Oxalic acid is a reducing agent and is able to produce the change which has taken place in the exposed portion. The speed of the action of the oxalic acid is greatly accelerated by either heat or light. Hence use can be made of the fact in photography.

Dip a piece of cotton in the third (unexposed) portion of the original solution, and spread a very light coat of the solution on a piece of blank writing paper in a dark room. When the paper is dry, lay an opaque object on it and expose the paper to the sunlight for a short time. To develop the print, dip the paper into a solution of potassium ferricyanid a few seconds, and then wash the paper thoroughly with water. Mount the paper on your record sheet. State all the chemical changes which have occurred in making the print.

Good blue-print paper can be made by coating a non-porous paper with a 10% ferric ammonium citrate solution and drying in a dark place. After the print is exposed, it is developed in a solution of potassium ferricyanid and fixed by washing in water. The commercial paper is coated with a mixture of these two compounds. It can be prepared in the following manner:

SOLUTION A		SOLUTION B	
Ferric ammonium citrate (green)	19.5 g.	Potassium ferricyanid	19.5 g.
Water	100 cc.	Water	100 cc.
Ammonium hydroxid, added till a decided color is obtained.			

Mix equal volumes of the two solutions, and to ten volumes of the mixture add three volumes of saturated oxalic acid solution. After being exposed, this paper is developed and fixed by washing it with water.

PHOTOGRAPHY

The essential materials for photography have long been known. A lens was discovered in the ruins of ancient Nineveh. The alchemists of the sixteenth century were acquainted with silver chlorid and knew that it darkened in the light. No one can be properly credited with the discovery of photography, because the art, as now known, has been gradually perfected by many workers. In 1727 Schulze copied the letters of a word by the action of the sunlight on silver chlorid, and a century later Daguerre prepared his famous " types."

The art of photography has created a wonderful industry. It is commonly thought of in connection with the making of portraits, but the sciences, arts, and industries now find this art an indispensable aid in their operations and advances. The X-ray enables the physician to treat the hidden causes of disease more intelligently ; the aëroplane scout searches out his enemy on the field of battle ; and the merchant sends a perfect likeness of his goods to distant customers. With its aid one may travel the world over, sitting comfortably at the fireside ; and masterpieces of art and sculpture are copied and brought to every one for appreciation. In courts of law it furnishes impartial and unprejudiced testimony to establish justice between man and man, while astronomers explore the depths of space with it and find millions of stars unseen by human eye.

The knowledge of an art that touches human activities at so many places surely has a place in a liberal education. Many persons understand the simple mechanics of " taking a picture," but few know the chemical reactions which form the basis of photography. The chemical part of the process of making a picture begins with the exposure to light of the sensitive plate. This plate is coated with silver bromid embedded in gelatin. After exposure the eye can detect no change on the plate until it has been placed in a developing solution. This solution acts upon the silver bromid in proportion to the intensity of the light falling on the different parts of the plate, and changes it to metallic silver. In this way the image is formed. The developing solution has little effect on unexposed silver salts. It is able to continue the change that has been started by the light.

Nearly all developers are carbon compounds derived from benzene. Pyrogallic acid is the most important because of its gradual action and great adaptability. Others in common use are hydroquinon, eikonogen, and metol. Besides one or more of the above developers, the solution contains other substances. The action of the developing agent alone is too slow, so an accelerator is needed. Alkalin substances serve this purpose, sodium carbonate being the best in most cases. Sodium hydroxid is sometimes used. The accelerator, however, causes the developing agent to act too rapidly where the light has affected the silver bromid only a little, and makes what is called a " fog " on the plate. A restrainer, potassium bromid, prevents this trouble, so it is generally used in all developing solutions. This reagent forms a double salt, $KAgBr_2$, which is more slowly reduced than the silver bromid. When in use, the developing agent becomes dark colored by oxidation. It then is likely to stain the plate that is being developed. If a substance that is more readily oxidized than the developing agent is put in the solution, it prevents stains and also preserves the developing solution so that it can be prepared and kept ready for use. This preservative is generally sodium sulfite. When exposed to the air, it oxidizes and becomes the sulfate, and has no value in a developing solution. Developing agents that deteriorate rapidly are generally kept in a solution separate from the accelerator. The two solutions are mixed only when they are to be used at once.

When the developing agent has produced an image of sufficient density or thickness, it is removed from the solution, washed with water, and placed in a solution of sodium thiosulfate (hypo). This solution removes all silver salt that the light and the developing agent have not affected. This process is called " fixing " because no further change can take place on the plate, hence the image is permanent. The fixing bath contains some alum, which hardens the coating in the plate and makes it less susceptible to injury. A little acid may be added also to help preserve the solution if it is prepared for future use.

Thus far only the " negative " is produced. On it the lights and shadows are reversed. A " positive " is made by exposing paper coated with silver bromid to light which has passed through the negative. It is then developed and fixed in the same manner as the negative.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.
Apparatus: Photographic trays.
Materials: Solutions of silver nitrate, potassium bromid (12 g. per L.), M.-Q. developer, sodium thiosulfate (saturated), printing paper.

I. **Properties of the Silver Bromid.** Darken the room for this experiment as well as possible.

To 5 cc. of silver nitrate solution add 5 cc. of the potassium bromid solution. Describe the precipitate, name it, and write the equation for its formation.

(*a*) Pour about one-fourth of the precipitate upon a filter paper and place it in the brightest light available. After a few minutes examine it again (?).

(*b*) Developing. Distribute the remaining precipitate into three test-tubes, saving one for the next part. Expose two of the tubes to bright light for two seconds and drop a couple of small crystals of potassium bromid (restrainer) into one of these tubes. When the crystals have dissolved, add 5 cc. of the developer to each tube. Describe and explain the changes occurring in the precipitates, and also note and explain any difference in the speeds of the changes in the two tubes. Save the contents of the tubes for the next part.

(*c*) Fixing. If the precipitate in the third tube has become dark on account of the room being too light, discard it and prepare a small fresh portion for use here. Add 5 cc. of the saturated sodium thiosulfate solution to this tube, and like amounts to each of the tubes used in Part I (*b*) (?). Explain any differences observed.

The commercial plates, films, and printing papers are always accompanied by full instructions for the timing of the exposures, developing, fixing, rinsing, etc. The success of the amateur will depend upon the faithfulness with which he follows the instructions. The development of plates or films is hardly practical as a laboratory exercise, unless a dark-room is available. Printing, however, may be carried out in an ordinary laboratory if it is provided with good shades.

II. **Printing.** (*a*) Place a small piece of printing paper in the light, observing it from time to time (?). Trim and mount neatly a 2 × 4-cm. piece of the paper and label it " exposed."

(*b*) Expose a piece of printing paper to strong light through a negative for a time interval directed by the instructor. Examine it in a dim light (?). Immerse the paper in the developing solution (?). Rinse and lay it on the table, and examine after a few minutes (?). Mount a piece of the paper as in (*a*) and label it " exposed and developed."

(*c*) Immerse an unexposed piece of printing paper in the developing solution about as long as the exposed piece was left in, rinse it, and immerse in the sodium thiosulfate solution for 10 minutes. Describe the result. Rinse thoroughly. Mount as before and label " developed and fixed, but not exposed."

(*d*) Expose another piece through the negative and develop it as in Part II (*b*), then fix it as in Part II (*c*), and rinse thoroughly. Describe it. Mount as before and label.

Questions: 1. Why is there no visible image on an exposed photographic plate or paper before development?

2. Explain how the image is brought out by the developer.

3. Why will the image on a developed plate or print become obscured if it is not fixed?

4. Explain how the image is fixed.

5. What is the condition of the silver salt on the **negative**: (*a*) in the shadows, (*b*) in the strong lights (high lights), after exposure, after development, and after fixing?

6. What is the condition of the silver salt on the **print**: (*a*) in the shadows, (*b*) in the high lights, after exposure, after development, and after fixing?

SILVERING GLASS FOR A MIRROR

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Materials: Paraffin, heavy, calendered paper, shellac varnish or asphaltum paint, paper clips or pins, zinc, plate glass, and the following solutions: silver nitrate, Rochelle salt (3 g. per L.), chromic acid cleaning solution.

I. **Cleaning the Glass.** If a mirror is to be resilvered, the paint or varnish may be removed with turpentine and ether, and the old silver with nitric acid. Cut a piece of heavy, calendered paper that is 8 cm. larger each way than the glass to be silvered. Fold up the corners of the paper to make a tray a little larger than the glass plate and fasten the corners with pins or paper fasteners. Pour some melted paraffin into the paper tray and by inclining it bring the bottom and the sides halfway up in contact with the paraffin. Pour back at once any excess paraffin.

Thoroughly clean a graduate, two small beakers, funnel, and the glass plate with a chromic acid cleaning solution, and finally rinse well with distilled water. During the rinsing and afterwards, **handle the glass plate by the edges only.**

II. **Preparing the Solutions.** The quantities here given are sufficient to silver 36 sq. in. of glass. Measure out into a beaker 20 cc. of the solution of Rochelle salt, add exactly 3.5 cc. of the reagent silver nitrate solution, and boil the mixture slowly for about 5 minutes. Then filter and set aside to cool.

Dilute 10 cc. of the reagent silver nitrate solution with an equal volume of distilled water. Measure out exactly 11 cc. of this diluted solution into a beaker and add to it slowly a diluted solution of ammonium hydroxid (1 cc. of reagent ammonium hydroxid to 10 cc. water) until the silver nitrate is quite brown. Continue to add the ammonium hydroxid, drop by drop, until the brown color just clears away, then mix it with the remainder of the diluted silver nitrate and filter the brown mixed solutions.

III. **Silvering the Glass.** Place the cleaned glass plate in a clean vessel and cover it with hot distilled water for a few minutes. Transfer the glass plate from the hot water to the paraffined tray with the face to be silvered up. Mix the two prepared solutions and pour at once on the plate in the tray. Let it stand quietly for 15 minutes. The heat from the plate causes the silver to be deposited more rapidly on the surface of the glass than elsewhere in the solution. Remove the mirror from the tray, wash, and let it dry. When dry, coat the silver deposit with shellac varnish, or with a protective paint made as follows: mix 10 cc. of turpentine with 6 cc. of white lead (in oil), add 13 cc. of Damar varnish (4 g. Damar gum in 11 cc. turpentine), and then add this to 100 cc. of asphaltum varnish made by dissolving enough asphaltum in 90 cc. of turpentine to make it like varnish in consistency.

IV. **Making a Silver Stickpin.** (Optional.) Old jewelry or a silver coin may be used as the source of silver. If it is desired to find the per cent of silver present, weigh the object carefully, and also the silver obtained at the end of the experiment, using special care not to lose any of the silver during the experiment.

Place the object in an evaporating dish and cover it with dilute nitric acid (1 acid : 1 water), warming very carefully if necessary to start the action. When in solution, add enough dilute hydrochloric acid to completely precipitate the silver as the chlorid, filter, and wash thoroughly with distilled water. To convert the chlorid to silver add a few pieces of mossy zinc, cover with dilute sulfuric acid, and let it stand until the precipitate is changed to a gray powder (silver) with no trace of the white silver chlorid. Remove any zinc remaining, and wash repeatedly by decantation to remove the sulfuric acid (test wash-water with barium chlorid solution). Fuse the gray mass on charcoal with the blowpipe. Insert a pinhead in the molten mass while it cools.

THE FINENESS OF GOLD

EXPERIMENT

Purpose: State it on a separate sheet used for the record.
Apparatus: Touchstone and set of testing needles, crucible, balance.
Materials: Gold article to be tested, gold ore, ferrous sulfate, ashless filter paper.

I. **Acid Tests.** Pure gold is too soft to give good service in jewelry, so it is commonly alloyed with copper to make it harder. Its fineness is reckoned in " carats," which means the number of parts of pure gold in 24 parts taken. The fineness may be determined approximately by testing it with the following concentrated acids, which will not noticeably affect the gold alloys for which each is prepared, but will leave a brown spot on alloys of a lower grade of fineness.

22 K. gold
30 cc. of C. P. nitric acid
4 drops of C. P. hydrochloric acid

20 K. gold
30 cc. of C. P. nitric acid
3 drops of C. P. hydrochloric acid

18 K. gold
60 cc. of C. P. nitric acid
2 drops of C. P. hydrochloric acid

14 K. gold
Use C. P. nitric acid

II. **Touchstone Method.** It is not always easy for one without experience to test gold by the above methods. If standard alloys of known fineness can be used as a basis for comparison, the amateur can test gold with reasonable sureness. These alloys are needle-shape, and are used with a black, Scotch, gold-testing stone.

Rub some part of the article to be tested on the stone, making a streak. Select two needles that are considered to be nearest the fineness of the article and make streaks with the standard alloys parallel with and one on either side of the other streak. With a glass rod draw a streak of aqua regia, as given above for 22 K. gold, across the streaks of gold. The lowest grade of alloy will be acted upon first and most rapidly. Other needles are then used until one is found whose properties agree with that of the gold in the article. This needle indicates the degree of fineness. Plated articles may be known by the high test of fineness and the cost. If the exterior of some part of the object can be removed so that the underlying metal may be rubbed on the stone, whether it is plated or pure gold can easily be told.

III. **Chemical Assay of Gold Ore.** The result of an assay of gold ore is reported in ounces of gold per ton of ore. Only a very small amount is assayed, so the ounces per ton must be computed. The milligrams of gold in 30 grams is the same as the number of ounces (troy) in 2000 lbs., so this weight, 30 grams, is called the **assay ton** (A. T.).

Weigh out carefully 10 g. of the finely pulverized ore ($\frac{1}{3}$ A. T.) in a porcelain evaporating dish. Put the dish on a water bath, add aqua regia until the ore is well covered, and cover the dish at once with a watch glass. The free chlorin reacts with the gold to form soluble gold chlorid, $AuCl_3$. Write the equation. After a half hour add about 3 cc. of distilled water, filter off the liquid, and wash the residue well with water, collecting the washings with the filtrate. Evaporate the filtrate to dryness over the water bath. Moisten the residue with water and again evaporate to dryness. Repeat this operation until all the acid is gone. Then add enough water to dissolve the gold chlorid.

Put several green crystals of ferrous sulfate in the solution and digest on a water bath for several hours. The gold will be thrown out of solution as a purple powder. The ferrous salt changes to the ferric form, both the sulfate and chlorid being formed. Write the equation. Filter off the gold on a paper of known ash, wash the residue on the paper very thoroughly with distilled water, place the paper in a clean weighed porcelain crucible, dry it in an air bath at 105° C. Then set the covered crucible on the triangle and heat until the paper is thoroughly charred. Remove the cover and continue the heating until all the carbon of the paper has burned off. Weigh the crucible again, subtract the weight of the ash of the paper unless it is smaller than could be weighed on your balance, and state the number of ounces of gold per ton of ore.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Apparatus: Erlenmeyer flask, air bath, balance, graduated flasks 100 and 250 cc., pipettes 50, 25, and 10 cc., burette, 250 cc. Kjeldahl flask, connecting bulb tube, condenser, 250 cc. receiving flask, extraction apparatus, water-bath or electric hot plate.

Materials: Solutions: Haines', Millon's, iodin, phenolphthalein, cochineal, N/10 and conc. sodium hydroxid, sulfuric acid N/10, sodium sulfate, lead subacetate, copper sulfate 1 %. Alcohol, ether or petroleum ether, benzine, sugar, potassium sulfate, soda-lime, zinc, copper sulfate crystals, cheese-cloth, S. & S. extraction shells, cobalt chlorid test paper, quantitative filter paper. Fresh liver, lean beef, bread, apple or carrot, peas or beans (dry), gelatin or casein or cheese.

TESTS

I. **Water**. Place some of the food in a test-tube and hold the latter horizontally over a flame, so as to dry the substance but not decompose it. If drops of liquid condense at the mouth of the tube, apply the cobalt chlorid paper test (p. 32).

II. **Carbohydrates**. Glucose. If a solid, make a water extract from some of the finely divided food, and make the Haines' test (see p. 94).

Sucrose. If glucose is found present in the water extract, it must be removed by boiling with excess of Haines' solution. After filtering, acidify the solution and boil it ten minutes with a few drops of hydrochloric acid in excess to invert the sucrose. Neutralize the acid, and repeat the Haines' test.

Starch. See the iodin test on page 96.

Cellulose. This is a tough, tasteless, fibrous substance found to greater or less extent in all vegetable tissues, and unaffected by repeated boiling with dilute acids or alkalis.

Glycogen. This is often called animal starch and is found in lean meat and liver. Boil 10 g. of finely minced liver (the fresher the better) for 20 minutes with 40 cc. of water. Filter through cheese-cloth and then through paper. Concentrate the filtrate to about 10 cc. by gentle boiling. Cool, and precipitate the glycogen by adding 20 cc. of 95% alcohol. Filter, wash the precipitate with a little alcohol, and then redissolve the precipitate in water. Add a few drops of iodin solution to a portion of the solution (?). Glycogen gives a red color with iodin. Test another portion with Haines' solution (?). Boil the remainder with a few drops of dilute hydrochloric acid, cool, neutralize, and again test with Haines' solution (?). Summarize the properties of glycogen.

Dextrin. Toast a piece of bread to convert some of the starch into dextrin. Scrape off the brown layer, shake thoroughly with water, and filter. To 5 cc. of the filtrate add alcohol to precipitate the dextrin (?). Test another portion of the filtrate with iodin solution (?). Most dextrins are colored red by iodin. Test another portion with Haines' solution (?). Boil the remainder with a few drops of hydrochloric acid, cool, neutralize, and test again with Haines' solution (?). Summarize the properties of dextrin.

Pectin. This is the jellying component of fruits and vegetables. Reduce a carrot or an apple to a fine pulp, cover with water, and boil gently for 10 minutes. Strain the juice through 4 thicknesses of cheese-cloth. To a little of the filtrate add a few drops of alcohol to precipitate the pectin (?). To another portion add a few drops of acetic acid and cool (?). To the remainder of the solution add enough sugar to make a thick syrup and cool (?). In case of a failure to secure the jelly in any test, evaporate off some of the water and repeat. Summarize the properties of pectin.

III. **Proteins**. There are many different proteins, widely scattered throughout the animal and vegetable world. Many are very important food components, being used by the body in building up tissues.

Burning Test. Burn a small piece of meat, or some dried peas or beans. The penetrating, disagreeable odor is due to the decomposition products of the proteins.

Ammonia Test. Thoroughly mix a little gelatin or casein or cheese with powdered soda-lime and heat in a dry test-tube. Note the odor (?), and test the gas given off with moist litmus paper (?).

Xanthoproteic Test. Note the color change produced by concentrated nitric acid on bread, boiled egg white, dried peas or beans, bone or horn (?). Wash off the acid and add ammonium hydroxid (?). Pure gelatin will not give this test.

Millon's Test. Use any of the proteins already mentioned, except gelatin; or use some finely chopped lean beef washed until the fibers are white (save the first wash-water). Boil a little of the substance with 5 cc. of Millon's reagent (see p. 255), and note the red precipitate collecting at the top of the liquid (?).

Biuret Test. Use the wash-water saved above, or a solution of egg white. Add an equal volume of concentrated sodium hydroxid solution to the solution to be tested and a drop or two of 1% copper sulfate solution (?). A blue to violet coloration indicates albumins, while a pink or red color indicates other proteins.

IV. **Fats and Oils.** Place a small piece of the substance on a piece of unsized paper and hold it high enough above a flame to avoid charring the paper. The formation of a grease spot on the paper indicates fat.

From one to three teaspoonfuls of the substance well ground or macerated and free from water are thoroughly stirred with an equal volume of ether or benzine (keep away from flames) for 10 minutes. Filter and place the liquid in a good draft of air until the odor of the ether or benzine has disappeared. The oil or fat remains in the dish.

V. **Mineral Matter.** Place some of the substance in an old teaspoon and heat it in a good draft until no more fumes are given off, and then as hot as possible to burn off the carbon. The white ash remaining consists of salts of the various metals present in the food.

COMPOSITION OF DIFFERENT FOODS

Test as many foods as are available and as time permits for the components named above, recording the results in a tabular form with the following heading, when possible using words which will indicate the approximate proportions present.

Food-stuffs	Water	Carbohydrates							Protein	Fat	Ash
		Glucose	Sucrose	Starch	Cellulose	Glycogen	Dextrin	Pectin			

QUANTITATIVE DETERMINATION OF SOME OF THE COMPONENTS

I. **Water.** Weigh accurately about 5 g. of the food in a tared (weighed) dish, and place in an air bath at 105° C. Weigh again when it seems dry, then put it back in the air bath and repeat until constant weight is obtained. Record all weights and calculate the per cent of water. Use the residue in the determination of mineral matter in V.

II. **Glucose**. Weigh out from 5 to 50 g. of the food accurately, using the smaller amount if much glucose is thought to be present. If a solid, the food must be divided finely and the glucose extracted by repeated treatments with small amounts of hot water, combining the filtrates in a beaker. Add from 2 to 5 cc. of lead subacetate solution to precipitate coloring matter if the filtrate is not clear, and filter into a graduated flask whose capacity is slightly larger than the volume of liquid. Rinse the beaker, pour rinsings through the filter, and add to the liquid in the flask. Make up to the mark with water. Mix thoroughly, and by means of a pipette draw off $\frac{1}{8}$ to $\frac{1}{4}$ of the solution (25, 50, or 75 cc. of the liquid) into a beaker; add enough concentrated sodium sulfate solution to precipitate the excess of lead, and filter into a 100-cc. graduated flask. Rinse the beaker into the filter paper, and add the rinsings to the flask. Fill up to the mark with water. If lead subacetate solution was not used, the sodium sulfate is omitted, and the portion of the solution may be run into the 100-cc. flask directly instead of into the beaker.

Fill a clean, dry burette with the glucose solution made above, and by means of a pipette run exactly 10 cc. of standardized Haines' solution into a 250-cc. Erlenmeyer flask. Add 40 cc. of water to the flask, and bring it to a boil over a wire gauze (2 or 3 small pieces of ignited pumice stone will stop bumping). Add 2-cc. portions of the glucose solution, boiling 3 minutes after each addition, until the Haines' solution becomes green. Then add 1-cc. portions until the color is dark red, and finally add 0.2-cc. portions until the precipitate is brick red, or better until the blue line seen just under the surface on looking through the liquid towards a window changes to a yellow or colorless line. Record the initial and final readings of the burette. Repeat the operations, running in at one time very nearly the full amount of glucose solution required the first time. The method is not very accurate if much less than 10 cc. for more than 25 cc. of the glucose solution is required for the reduction of the 10 cc. of Haines' solution. In such case, it is well to repeat the preparation of the final glucose solution, varying the proportion of the first glucose solution and water so as to bring the results within these limits.

Knowing that the volume of the glucose solution used in the titration contains 0.05 g. of glucose, calculate the weight of glucose contained in the final glucose solution (?), and then the weight of glucose in the original solution (?). This is the weight of glucose extracted from the food. Calculate the per cent present in the food.

III. **Proteins by the Modified Gunning Method.** Place exactly 0.5 g. of the finely divided food-stuff in a clean, dry 250-cc. Kjeldahl flask. Mix 10 g. of powdered potassium sulfate with 0.5 g. of powdered copper sulfate crystals, add this to the food and mix them thoroughly. Pour in 25 cc. of concentrated sulfuric acid, and heat gently in a hood with a small flame until all foaming stops. Incline the flask somewhat to avoid loss of material by spattering. Keep the hands from under the flask. If it should break, the hot acid might cause very serious burns. When the foaming has ceased, slip a gauze under the flask and increase the heat until the gauze is bright red where it comes in contact with the bottom of the flask, causing the acid to boil gently. Continue the heating 90 minutes after the liquid becomes clear. The addition of about 1 g. of ferrous sulfate crystals will prevent bumping of the liquid during boiling. By this digestion of the food with acid, the nitrogen compounds are decomposed, and all of the nitrogen is changed into ammonium sulfate.

Allow the flask to cool, and then pour the liquid into 50 cc. of water in a 500-cc. boiling flask. Use three 50-cc. portions of water to rinse the first flask, adding the rinsings to the liquid in the second. Attach a connecting bulb tube to the flask and set up the apparatus with a condenser and receiving flask as shown in Fig. 96.

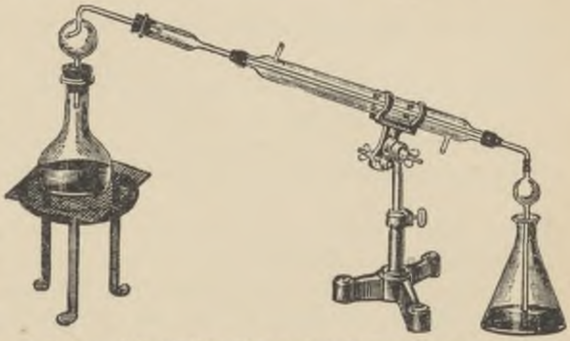


FIG. 96. Distillation of ammonia formed from proteins.

With a pipette measure exactly 50 cc. of N/10 sulfuric acid solution into the receiving flask, and adjust the condenser so that the lower end of it dips just below the surface of the acid solution. Disconnect the distilling flask and add a few pieces of zinc to prevent bumping, 2 drops of phenolphthalein solution as indicator, and then enough saturated sodium hydroxid solution to make the contents strongly alkalin. Reconnect the apparatus immediately, and distil off about 150 cc. of the liquid. The ammonia set free by the alkali passes over with the steam, and neutralizes part of the N/10 sulfuric acid.

Titrate the excess of N/10 sulfuric acid with N/10 sodium hydroxid, using cochineal solution as indicator. The volume of N/10 sulfuric acid neutralized by the ammonia is found by subtracting the volume of the sodium hydroxid solution from 50. Each cubic centimeter of the N/10 sulfuric acid solution is equivalent to 0.0014 g. of nitrogen. So calculate the weight of nitrogen in the food, and multiply this by 6.25 to get the weight of proteins (proteins contain 16 % nitrogen). Calculate the per cent of proteins in the food.

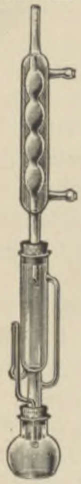


FIG. 97. Soxhlet extraction apparatus.


FIG. 98. Wiley extraction apparatus.

IV. **Fat**. Weigh accurately 2 or 3 g. of the finely divided sample in a tared extraction shell, and then dry it thoroughly in an air bath at 100° C. Place the shell in a Soxhlet (Fig. 97) or Wiley (Fig. 98) extractor. Pour 100 cc. of water- and fat-free ether or petroleum ether in the extraction flask, and start the water through the condenser. Boil the ether by an electric hot plate or a large bath of hot water. Remember that ether is very inflammable, and its fumes must not come near a flame. The ether vapors rise into the condenser, liquefy, and fall upon the food-stuff from which the ether extracts fat, and then returns to the extraction flask with the fat. After two hours remove the shell and dry it to constant weight. Calculate the per cent of fat lost.

V. **Mineral Matter**. Heat the residue left from the water determination until a whitish ash is left, taking care not to lose any of it. Weigh carefully, and calculate the per cent of mineral matter.

VI. **Soluble Matter in Tea**. Place a known weight (about 2 g.) of powdered tea in a weighed quantitative filter paper in a funnel, and pour through it successive portions of boiling water until the filtrate is no longer colored. Dry the paper and residue at 100° C. in an air bath, and weigh. The loss of weight represents the soluble matter. Calculate its per cent. In a good tea this should not be less than 40 %. A result decidedly lower indicates the presence of spent or exhausted tea leaves.

Study Topic: Sadtler, *Chemistry of Familiar Things:* "Foods," pp. 178–224.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Apparatus: Separatory funnel, sand-bath, tablespoon, iron nail, alkalimeter.

Materials: Foods, etc., to be tested should be brought from home, chloroform, ferric alum crystals, ferric chlorid solution, milk, Halphen's reagent, Columbian spirits, ether, amyl alcohol, boric acid (dilute solution), cheese-cloth, gasoline, solutions of sulfanilic acid and a naphthylamine, zinc chlorid, potassium dichromate, bromin water.

Definition of Adulteration. If the quality of a food is lowered by the removal of a part of it or by the addition of substitutes, if imitations are sold in the name of another article or substances are mixed with foods to conceal inferiority, or if the food contains any poisonous ingredient or is decomposed so as to be unfit for food, the food is considered adulterated and its sale is forbidden by law.

I. **Food Preservatives.** Benzoic Acid, Salicylic Acid, and Saccharin. These preservatives are used with fruit preparations, jams, jellies, preserves, cider, wines, canned vegetables, ketchup, and similar foods. If the food is not liquid or entirely soluble in water, macerate it in a mortar with water containing a little sodium hydroxid and strain through four thicknesses of cheese-cloth. Use the filtrate. Make a 50-cc. sample (about 3 tablespoonfuls) acid with a little sulfuric acid and mix it with 40 cc. of chloroform in a separatory funnel (Fig. 99) by a vigorous rotary motion. Do not shake. The chloroform dissolves only the preservatives. (Ether or petroleum ether may be used instead of the chloroform. Keep away from flames. After the settling, separate the solvent and use as directed for the chloroform.) After settling let the heavy chloroform run into a beaker, and clean the funnel. Pour half the solvent back into the funnel.

Test this portion for salicylic acid by adding an equal volume of water and a piece of ferric alum a little larger than a pinhead. Shake the mixture thoroughly and let it settle (?). If salicylic acid is present, the water layer will become purple. Save the used chloroform in a provided bottle and at a convenient time purify it by distillation.

Fig. 99. Separatory funnel.

To test for benzoic acid, make the remaining portion of chloroform neutral with dilute ammonium hydroxid and evaporate half of it on a water bath. Pour the remainder into a watch glass, and place on it another watch glass of the same size with the convex side up. Make a 1-cm. hole in the center of a filter paper and insert the paper between the two glasses. Heat the watch glasses in this position on a sand bath (?). If benzoic acid is present, it will sublime in minute, flat, needle-shaped crystals on the upper watch glass. Pour 5 cc. of dilute ammonium hydroxid on the sublimate in the watch glass, heat it on a water bath until the liquid is neutral to litmus, and add a drop or two of neutral ferric chlorid solution (?). A flesh-colored precipitate confirms benzoic acid.

Taste a minute quantity of any residue left in the lower watch glass after evaporation (?). If it tastes sweet, saccharin was present.

See the tests for boric acid and borax, page 128, sulfurous acid, page 104, formaldehyde, page 90.

II. **Butter.** Oleomargarine and renovated butter are frequently represented as butter.

(*a*) Spoon Test. Melt a lump of the so-called butter in a tablespoon. Real butter will boil rather quickly and produce much foam. Oleo or renovated butter sputters noisily and makes little foam.

(*b*) Waterhouse Test. Put about 50 cc. of sweet milk in a wide-mouth bottle and set it in hot water. When the milk is hot, put in a spoonful of butter and stir it with a wooden splinter until the butter melts. Then place the bottle in ice water and continue to stir (?). Real or renovated butter becomes solid in small granules in the milk, but the oleo attaches to the stick in one lump.

III. **Olive Oil.** This is the most costly of the edible oils, and cotton-seed oil the cheapest, so it is the common adulterant.

Heat 5 cc. of the oil in a test-tube loosely stoppered with cotton with the same volume of Halphen's reagent (p. 254) in a boiling bath of saturated brine for 10 minutes (?). A reddish color shows the presence of cotton-seed oil in proportion to the intensity of color.

IV. **Coloring Matter.** See tests for coal-tar colors, page 183.

(*a*) Mineral. Copper. Copper salts are sometimes used in imported canned peas, beans, spinach, etc., to intensify the green color. Macerate four tablespoonfuls of the product in a mortar with water to make a thin paste. Put the paste in a beaker, add 5 cc. concentrated hydrochloric acid, stir well, and then put a bright iron nail in the paste. Set the beaker in the boiling water of a water bath for 20 minutes and stir occasionally with a wooden stick. Examine the nail (?). If copper is present, it will be found as a coating on the nail.

(*b*) Vegetable Colors. Caramel. (Fradiss Test.) Heat 2 tablespoonfuls of the product to dryness over a water bath. Stir the dry residue with 25 cc. of warm Columbian spirits and filter. Caramel colors the alcohol brown. Add amyl alcohol or chloroform to the solution. A brown flocculent precipitate, slowly settling, shows caramel was present.

Turmeric. Extract the dried residue with alcohol as for caramel. It colors the alcohol yellow. Dip a filter paper in the solution and then dip it for a moment in a dilute solution of boric acid (or of borax made acid with hydrochloric acid). Dry it on the outside of a beaker of hot water. If turmeric was present, the paper will be cherry red, turning olive green on adding ammonium hydroxid. This is the converse of the test for boric acid, page 128.

V. **Flavoring Extracts.** (*a*) Vanilla. This extract may be adulterated with coumarin, the extract of the tonka bean, or it may be artificial vanilla. True vanilla contains much resin, which will be precipitated in a few minutes if 10 cc. of the extract is mixed with an equal volume of lead acetate solution (189.5 g. per L.) (?). The absence of a precipitate shows it to be artificial.

Leach's Test for Coumarin. Shake the filtrate from the resin with 10 cc. of ether in a separatory funnel and separate the two liquids. Shake the filtrate twice more with 10-cc. portions of ether and combine the ether extracts. Evaporate the ether and add a few drops of warm water to the residue and then add a little iodin in potassium iodid solution (?). If coumarin is present, a brown precipitate will form, breaking up into dark green flecks on stirring.

(*b*) " Strength " of Lemon Extract. Put 5 cc. of the extract into a 15-cm. tube and fill it with water. The degree of cloudiness or milkiness indicates the " strength " of the extract.

VI. **Vinegar.** It may be made from cider, wine, malt, spirit, glucose, molasses, and wood products. The last three are generally adulterations for the other kinds.

Evaporate 25 cc. or more of the vinegar on a water bath. As it becomes dry test the odor of the residue carefully. Its character indicates the source. Cider — baked apples; wine — vinous; malt — sour beer; molasses and glucose — burnt sugar; spirit and wood — little odor. Caramel and apple jelly are frequently added to imitate cider vinegar. Test for caramel as previously given (?). Test the residue on a platinum wire in a burner flame (?). (The color imparted will, if the vinegar has been made from pure cider exclusively, consist altogether of the pale lilac color of a potash salt without any of the yellow sodium flame being visible. In all vinegars other than of pure cider the yellow sodium flame will predominate. — Leach.)

VII. **Spices.** Spices are more commonly adulterated than any other food product, owing to the fact that they are retailed in a finely ground condition. Microscopical methods give the only satisfactory way of detecting adulterants in all spices. Some spices are colored. Test mustard for turmeric. Put a half teaspoonful of cloves, mustard, or cayenne pepper into 50 cc. of water and boil it for 3 minutes, dilute the mixture with 4 parts water, and test for starch with iodin solution. Starch is a very common adulterant. It is not contained in these three spices naturally.

VIII. **Flour Bleached by Oxids of Nitrogen.** Many brands of flour are bleached by means of nitrogen tetroxid, because people have come to believe, mistakenly, that the flour that makes the whitest bread is the best. A very small proportion of the gas is absorbed by the flour, and many eminent physicians assert that the long-continued use of flour bleached in this fashion is detrimental to the health.

A simple test for a bleached flour can be made by shaking a teaspoonful of the flour with about 75 cc. of gasoline. The gasoline will remain white if the flour is bleached, but is colored yellow by an unbleached flour. To test for the oxids of nitrogen, place on a smoothed surface of the flour a few drops of a solution of equal volumes of sulfanilic acid and alpha-naphthylamine hydrochlorid (see p. 255). Flour bleached by nitrogen tetroxid turns the unabsorbed liquid pink or crimson.

IX. **Determination of Available Carbon Dioxid in Baking Powder.** Baking powder is used to liberate carbon dioxid at the proper time to aërate bread and cake. Its sole value depends upon this one point. The Illinois Food Laws require that baking powders shall have at least 10 % of carbon dioxid available under the conditions used in cooking. Powder is under standard usually for one of two reasons: 1st, since it is unstable and slowly loses carbon dioxid on contact with atmospheric moisture, it may be below standard if over six months old; 2d, there may have been a deficiency of the bicarbonate or acid principle ingredient when the powder was made. For the sake of time and convenience, special apparatus, Schroedter's alkalimeter (new form), is best. Handle it with care since it is fragile.

Have the apparatus clean and dry, and weigh it carefully (?), then put about 1 g. of the powder in it through the glass-stoppered opening in the side. Weigh it again carefully (?). The difference is the actual weight of the powder taken. Fill the trap bulb (Fig. 100) half full of concentrated sulfuric acid, and put 10 cc. of fresh-boiled distilled water in the other bulb, having the stop-cock closed. Have the apparatus dry on the outside and weigh the whole apparatus carefully (?). Let the water run slowly drop by drop upon the powder. When effervescence has ceased, set the apparatus into slowly boiling water in the water bath for a few minutes. Hold it in place with a clamp. Attach a drying tube filled with calcium chlorid at the top of stop-cock bulb, and attach an aspirator to the trap bulb. Draw air slowly through the drying tube to remove moisture and then through the apparatus (2 bubbles per second) for 5 minutes to expel all carbon dioxid. Dry the outside of the apparatus, and when cool, weigh again (?). The difference in the last two weighings is the loss of carbon dioxid. Find what per cent this difference of weight is of the powder taken (?).

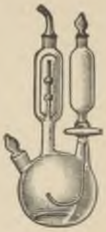


FIG. 100. Schroedter's alkalimeter.

X. **Tests of Headache Powders for Harmful Drugs.** The drugs commonly used are acetanilid, antipyrin, and phenacetin. They are all poisonous and habit forming, and are taken generally by persons not in the best of health. They do not remove the cause of a headache, only cover it up, leaving the person in a worse condition than before. They depress the heart action and frequently cause death. These drugs should **not be taken except by the orders of a competent physician.** Acetanilid is the one most commonly used.

(*a*) Strobel Test. Put a pinch of the powder in a test-tube with the same quantity of zinc chlorid and heat the mixture slowly. Hold a pine splinter in the white fumes that appear (?). The mixture will become black finally, and the splinter will be stained yellow if acetanilid is present.

(*b*) Isonitril Test. Add 5 cc. of a 10 % solution of sodium hydroxid to a half gram of the powder and heat the mixture. Add a few drops of chloroform and set the tube in the hood for 10 minutes (?). A very offensive odor of phenylcarbamin can be detected if acetanilid or phenacetin is present.

(*c*) Antipyrin. Dissolve a half-gram sample in water and add ferric chlorid solution (?). A deep red coloration shows the presence of antipyrin.

XI. **Tests for Alkaloids.** The alkaloids are basic nitrogenous substances, used as medicines. They are very poisonous and are habit-forming drugs. They should not be used except as prescribed by a competent physician. Many of these alkaloids are narcotics. They are used in soothing syrups for children. Many cases are on record of children being put to sleep by doses of paregoric, from which they never awakened.

(*a*) Morphin and Laudanum. Treat a sample of the product with concentrated sulfuric acid and set it aside for 15 hours. Then add nitric acid (?). A bluish violet color changing to blood red shows morphin to be present. This test will detect as little as 0.01 mg. of morphin.

(*b*) Strychnin. Mix a little concentrated sulfuric acid in an evaporating dish with a small sample of the substance to be tested. Sprinkle a little finely powdered potassium dichromate over the liquid (?). An intense violet color, gradually becoming bright red and then yellow, shows the presence of strychnin.

(*c*) Quinin. Make a solution of the sample with water or alcohol and add bromin or chlorin water (?). An emerald, green coloration shows the presence of quinin. Dilute solutions of quinin show light blue fluorescence.

(*d*) Caffein. Test coffee and tea for this alkaloid by first extracting it by boiling in water for 15 minutes. Add nitric acid to the solution after filtering, and evaporate to dryness (?). If caffein is present, there will be a yellow residue becoming intensely violet when treated with ammonia.

Study Topics: Baskerville, *Municipal Chemistry:* "Food Adulteration," pp. 119–135; "Food Inspection," pp. 136–146; "Habit Forming Agents," pp. 175–192.

REPORT OF CHEMICAL ANALYSIS

EXPERIMENT

Partial analysis of ....

Obtained from ....

Manufactured by ....

Exact name of brand ....

Tested for ....

Description of test ....

....

....

....

....

Equations of test ....

....

....

Found ....

....

Quality of sample ....

Remarks ....

....

....

I hereby certify that the above report is correct to the best of my knowledge.

(Signed) ....

Date ....

REPORT OF CHEMICAL ANALYSIS

EXPERIMENT

Partial analysis of ....

Obtained from ....

Manufactured by ....

Exact name of brand ....

Tested for ....

Description of test ....

....

....

....

....

Equations of test ....

....

....

Found ....

....

Quality of sample ....

Remarks ....

....

....

I hereby certify that the above report is correct to the best of my knowledge.

(Signed) ....

Date ....

MILK

Milk is one of the most important foods because it contains all the food components needed by the body, and in about the proper proportion. This, together with its ease of digestion, makes milk especially valuable for invalids and infants. It is a favorable medium for the growth of microorganisms, so that the greatest care is necessary to keep it from contamination. Impure milk is often the means of spreading diseases. The number and kind of microörganisms in milk can be found only by a person trained in bacteriological methods. Milk may be fraudulently adulterated by skimming, watering, coloring with dyes to produce the appearance of richness, and by the addition of preservatives. The preservatives generally used are harmful, formaldehyde being very dangerous to infants. They came into use in the endeavor to keep milk sweet in spite of careless handling and delay in marketing. Most of the milk supplied to large cities at present is pasteurized, so that preservatives are infrequently used now. Ignorant or unscrupulous dealers in small towns where the laws are not rigidly enforced often add preservatives to milk. The coloring matter sometimes used is not harmful in itself, but its purpose is to deceive the consumer, so it is objectionable.

Dirt may be found in milk, and its presence is directly due to careless handling, particularly in the milking operations. The number of bacteria present is certain to be large in this case. The consumer has reason to object if dirt is present in milk.

The nutritive value and the manner of preparing milk for market can be determined by an examination of its physical qualities, particularly the specific gravity, the amount of fat, and dissolved solids. The specific gravity of milk depends upon the amounts of fat and solids-not-fat. The specific gravity of good milk at 60° F. (15.5° C.) should be between the limits 1.027 and 1.033, usually 1.031. The fat is lighter than water, while the solids-not-fat are heavier than water. If the specific gravity is low, it is due to either (1) a high per cent of fat, or (2) a low per cent of solids-not-fat. The second case may be due to " watering the milk." If the specific gravity is high, it is due to either (1) a low per cent of fat, or (2) a high per cent of solids-not-fat. Both of these conditions may be due to " skimming the milk." It may be seen from these facts that clever skimming and watering, with the addition of coloring matter, may deceive all but the expert. In this case, however, the total solids and solids-not-fat will be below the average, and the amount of fat in comparison to solids-not-fat very small. When such milk is allowed to stand, it gives very little cream and a rich-looking skimmed milk.

The U. S. Department of Agriculture has set a standard of 3.25 % for milk-fat and 8.5 % for solids-not-fat. The mineral residue or ash should be about 0.70 %. If over 0.90 %, mineral matter has been added to try to cover up dishonest manipulation of the milk. Many states, including Illinois, require only 3 % of fat in milk. Some cows give milk low in fat or solids-not-fat, or both, but such milk must be mixed with a richer milk before being sold.

Study Topic: Baskerville, *Municipal Chemistry*: " Milk," pp. 90–118.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.
Apparatus: Babcock milk tester, Quevenne lactometer, Richmond's milk scale, Mohr's burette.
Materials: 1 pint fresh milk (from home), tenth normal sodium hydroxid solution.

I. **Visible Dirt, Bottle and Cap, Flavor and Odor.** Judge and score your sample of milk according to the directions on the reverse side of the score sheet.

II. **Fat by the Babcock Method.** Always mix the milk thoroughly before removing a sample from the bottle. By means of a special pipette run exactly 17.6 cc. of milk into each of two test bottles, and add slowly 17.5 cc. of sulfuric acid (Sp. Gr. 1.83 at 15° C.), letting the acid wash the milk down into the bottle. Holding the bottle by the neck over a sink (in case of accidents), mix the contents by a rotary motion. Do not get curds into the neck of the bottle. The contents will become dark brown and hot. Label the bottles with initials and place them in the centrifugal machine in diametrically opposite holders so that the rotating part will be balanced. Operate the machine for five minutes. Set the bottles in a pan of hot water and add hot water up to but not into the necks of the bottles. Whirl them for two minutes. Set the bottles in hot water again and add enough hot water to bring the bottom of the fat column above the zero of the scale on the neck of the bottle. Whirl them again for one minute. Read and record the per cent of fat as it is shown on the scale etched on the neck of the bottle, reading the extreme top and bottom of the fat column. The numbered divisions are per cents, so graduated from the quantity of milk put in the bottle. If white specks show in the fat column, a little more acid should be used; if black specks, use a little less. Empty the bottles as soon as the amount of fat is noted, and clean the pipette and bottles with warm water, soap, and sodium carbonate.

III. **Specific Gravity by the Lactometer.** Cool the milk to 60° F. (15.5° C.), or make temperature corrections by Richmond's slide rule, or by the table on page 245. Shake the milk thoroughly, then fill a dry cylinder deep enough to float a Quevenne lactometer in the milk. Take the reading on the scale as indicated by the surface of the milk and record the temperature of the milk. Attach the lactometer reading to the figures 1.0 to express the specific gravity. Clean the lactometer and cylinder carefully with warm water, soap, and sodium carbonate.

IV. **Solids by Richmond's Slide Rule.** The official method of determining the total solids consists of evaporating a known weight of milk (about 2 cc.) to dryness on a water-bath, and computing the per cent from the weight of the residue. It may be found with a fair degree of accuracy from the table, page 244, by using the per cent of fat and specific gravity previously found. The mathematical relationship between these factors has been worked out more conveniently on Richmond's milk scale (Fig. 101), which is operated as follows: Set the observed lactometer reading on the sliding scale opposite the observed temperature on the upper scale, and the lactometer reading, corrected for temperature, will be found opposite the 60° F. mark. Call this reading (A). Set the arrow at the right-hand end of the slider at the reading (on the upper scale) for the per cent of fat as already found, and then find (A) on the lower edge of the slider. This reading will locate the mark (on the lowest scale) indicating the per cent of total solids with a fair degree of accuracy. Total solids minus per cent of fat gives per cent of solids-not-fat.

V. **Acidity and Bacteria.** Follow the directions on the reverse side of the score sheet.

VI. **Butter Fat in Ice-cream by the Babcock Method.**[1] (Optional.) Warm the sample to 40° C. and mix thoroughly by pouring and stirring. If fat has separated, a few grains of powdered sodium hydroxid will aid in emulsifying it. Weigh carefully 9 g. of the cream into a 10 % Babcock milk bottle, using a pipette with a rather wide opening.

Pour 9 cc. of glacial acetic acid into a small beaker, add 9 cc. of concentrated sulfuric acid (commercial), stir, and cool. Add 12 to 15 cc. of this mixture to the ice-cream, and mix with a rotary motion. Set the bottle in a water-bath, and heat with occasional shaking until the mixture turns a dark chocolate color. Remove the bottle, let it stand 10 minutes, and then whirl it 10 minutes in the centrifuge. Shake again, with a whirling motion. Fill the bottle to the neck with hot water, adding a little of the acid mixture with the water if a precipitate forms. Whirl three minutes, and then add enough boiling hot water to bring the fat column up into the graduated portion of the neck of the bottle. Whirl two minutes. Place the bottle in water at 55° C. for ten minutes, and then read from the bottom of the fat column to the extreme top of the meniscus. Twice this value will be the per cent of fat in the ice-cream.

VII. **Preservatives.** (Optional.) If the mixture in the Babcock bottle turns purple instead of brown, the milk contains formaldehyde. The test may also be made as on page 90. Evaporate 75 cc. of the milk to dryness and then heat the residue in the hood until it is white again. Test a little of the ash for borax or boric acid as on page 128, and test the remainder for sodium bicarbonate as on page 133.

---

[1] Utt, *The Journal of Industrial and Engineering Chemistry*, Vol. VII (Sept. 1915), p. 773.

EXPERIMENT

Milk from.......................................Co. Cost............Date............

(See directions on next page.)

Item	Perfect Score	Score Allowed	Observations
1. Visible dirt (after standing 1 hr.)	10	....	....
2. Bottle and cap...............	5	....	Cap............................
			Bottle.........................
3. Flavor and odor...............	25	....	Flavor ....
			Odor...........................
4. Fat..........................	10	....	Per cent found ....
5. Solids-not-fat (total solids — fat)	10	....	Sp. Gr...........Temp..........
			Corrected Sp. Gr.................
			Total Solids.....................
			Solids-not-fat ....
6. Acidity	5	....	Per cent found ....
7. Bacteria	35	....	No. per cc. ....
Total.....................	100	....	

Inspector..

Fig. 101. Richmond milk scale.

DIRECTIONS FOR SCORING MILK

1. **Visible Dirt — Perfect Score, 10.** (Examine at home.) After the milk has stood quietly for at least **one hour**, lift the bottle high enough so that the bottom can be seen easily. Then slowly tilt the top of the bottle away from the observer, watching the bottle meanwhile for small particles of moving sediment resting on the bottom. State the amount and kind of dirt under "Observations," and decide on the score accordingly.

2. **Bottle and Cap — Perfect Score, 5.** (Observe before opening.) Bottles should be of clear glass and have no attached metallic parts. Caps should be paraffined and so placed as to give adequate protection against contamination. Describe conditions and score accordingly.

3. **Flavor and Odor — Perfect Score, 25.** After opening the bottle, mix the milk thoroughly by pouring it into a clean bottle and back again several times. Then warm a sample slightly and note odor and flavor. Deductions should be made for disagreeable and foreign odors and flavors. The senses of taste and smell are not well trained, so care should be exercised in judging. Describe the conditions found.

4. **Fat in Milk — Perfect Score, 10.**

4.0% and over	10.0 points	3.6%	9.2 points	3.2%	6 points
3.9%	9.8 points	3.5%	9.0 points	3.1%	5 points
3.8%	9.6 points	3.4%	8.0 points	3.0%	4 points
3.7%	9.4 points	3.3%	7.0 points		

NOTE.— When the per cent of fat is less than the local legal minimum, the score shall be 0. Illinois minimum, 3% fat.

5. **Solids-not-fat — Perfect Score, 10.**

8.7% and over	10 points	8.3%	6 points	7.9%	2 points
8.6%	9 points	8.2%	5 points	7.8%	1 point
8.5%	8 points	8.1%	4 points	7.7%	0 point
8.4%	7 points	8.0%	3 points		

NOTE. — If the per cent of fat is equal to or above the local legal minimum, but the per cent of total solids is found to be below the legal minimum, the score of solids-not-fat shall be 0. Illinois legal minimum, 11.5% total solids.

6. **Acidity — Perfect Score, 5.** Measure out 25 cc. of milk with a pipette, add two drops of phenolphthalein solution, and add slowly from a Mohr's burette a tenth normal solution of sodium hydroxid until a faint pink color is produced. The volume of sodium hydroxid solution used multiplied by 0.897 and divided by the volume of milk used is the per cent of lactic acid present.

Less than 0.20%	5 points	0.22% to 0.23%	2 points
0.20% to 0.21%	4 points	0.23% to 0.24%	1 point
0.21% to 0.22%	3 points	0.24% and over	0 point

7. **Bacteria per Cubic Centimeter — Perfect Score, 35.** The approximate number of bacteria per cubic centimeter in raw, unpasteurized milk and cream may be found according to the method given in *An Easy Test for Bacteria in Milk and Cream* (Barthel's Reductase Test Improved), published by Hugo Jone, P. O. Box 37, Station A, Brooklyn, N. Y. If the number of bacteria cannot be determined by either the standard plating method or the Reductase test, the score should be called perfect.

	POINTS		POINTS		POINTS
Less than 400	35.0	20000–25000	26	90000– 95000	12
400– 700	34.5	25000–30000	25	95000–100000	11
700– 1000	34.0	30000–35000	24	100000–110000	10
1000– 2000	33.5	35000–40000	23	110000–120000	9
2000– 3000	33.0	40000–45000	22	120000–130000	8
3000– 4000	32.5	45000–50000	21	130000–140000	7
4000– 5000	32.0	50000–55000	20	140000–150000	6
5000– 6000	31.5	55000–60000	19	150000–160000	5
6000– 7000	31.0	60000–65000	18	160000–170000	4
7000– 8000	30.5	65000–70000	17	170000–180000	3
8000– 9000	30.0	70000–75000	16	180000–190000	2
9000–10000	29.0	75000–80000	15	190000–200000	1
10000–15000	28.0	80000–85000	14	200000 and over	0
15000–20000	27.0	85000–90000	13		

NOTE. — When the number of bacteria per cubic centimeter exceeds the local legal limit, the score shall be 0.

INSPECTION OF CREAM

EXPERIMENT

Cream from.......................................Co.   Cost............Date.............

Item	Perfect Score	Score Allowed	Observations
1. Visible dirt (after standing 1 hr.)..	10	....	....
2. Bottle and cap.................	5	....	Cap........................... Bottle.........................
3. Flavor and odor................	25	....	Flavor........................ Odor..........................
4. Fat...........................	20	....	Per cent found..................
5. Acidity.......................	5	....	Per cent found..................
6. Bacteria......................	35	....	No. per cc......................
Total.....................	100	....	

Inspector..

Fat in Cream by the Babcock Method. The test bottle for cream is first weighed, 9 cc. of the well-mixed sample of cream are introduced by means of a rather wide-mouthed pipette, and the bottle weighed again to centigrams. Five to six cc. of water are added to the cream in the bottle, mixed thoroughly, and then the regular amount of sulfuric acid used in the Babcock milk test (17.5 cc.) is measured in, and the test continued in the manner employed for milk. Multiply the reading of the fat by 18 and divide the product by the weight of the cream taken to find the per cent of fat.

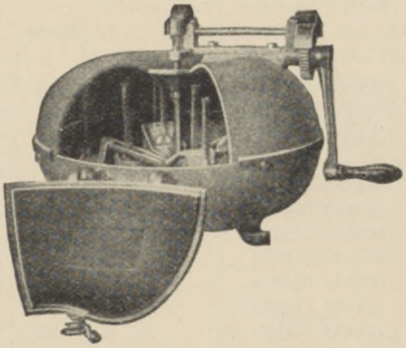


Fig. 102. Babcock milk tester.

DIRECTIONS FOR SCORING CREAM

1. **Visible Dirt — Perfect Score, 10.** (Examine at home.) After the cream has stood quietly for at least **one hour**, lift the bottle high enough so that the bottom can be seen easily. Then slowly tilt the top of the bottle away from the observer, watching the bottom meanwhile for small particles of moving sediment resting on the bottom. State the amount and kind of dirt under "Observations," and decide on the score accordingly.

2. **Bottle and Cap — Perfect Score, 5.** (Observe before opening.) Bottles should be of clear glass and have no attached metallic parts. Caps should be paraffined and so placed as to give adequate protection against contamination. Describe conditions and score accordingly.

3. **Flavor and Odor — Perfect Score, 25.** After opening the bottle, mix the cream thoroughly by pouring it into a clean bottle and back again several times. Then warm a sample slightly and note odor and flavor. Deductions should be made for disagreeable and foreign odors and flavors. The senses of taste and smell are not well trained, so care should be exercised in judging. Describe the conditions found.

4. **Fat in Cream — Perfect Score, 20.** If 20 per cent of fat or above, score perfect. Deduct 1 point for each one-half per cent fat below 20 %. If the per cent of fat is below the legal minimum, the score should be 0. Illinois legal minimum, 18 %.

5. **Acidity — Perfect Score, 5.** Measure out 25 cc. of cream with a pipette, add two drops of phenolphthalein solution, and add slowly from a Mohr's burette a tenth normal solution of sodium hydroxid until a faint pink color is produced. The volume of sodium hydroxid solution used multiplied by 0.897 and divided by the volume of cream used is the per cent of lactic acid present.

Less than 0.20%	5 points	0.22% to 0.23%	2 points
0.20% to 0.21%	4 points	0.23% to 0.24%	1 point
0.21% to 0.22%	3 points	0.24% and over	0 point

6. **Bacteria per Cubic Centimeter — Perfect Score, 35.** The approximate number of bacteria per cubic centimeter in raw, unpasteurized milk and cream may be found according to the method given in *An Easy Test for Bacteria in Milk and Cream* (Barthel's Reductase Test Improved), published by Hugo Jone, P. O. Box 37, Station A, Brooklyn, N. Y. If the number of bacteria cannot be determined by either the standard plating method or the Reductase test, the score should be called perfect.

	Points		Points		Points
Less than 400	35.0	20000–25000	26	90000– 95000	12
400– 700	34.5	25000–30000	25	95000–100000	11
700– 1000	34.0	30000–35000	24	100000–110000	10
1000– 2000	33.5	35000–40000	23	110000–120000	9
2000– 3000	33.0	40000–45000	22	120000–130000	8
3000– 4000	32.5	45000–50000	21	130000–140000	7
4000– 5000	32.0	50000–55000	20	140000–150000	6
5000– 6000	31.5	55000–60000	19	150000–160000	5
6000– 7000	31.0	60000–65000	18	160000–170000	4
7000– 8000	30.5	65000–70000	17	170000–180000	3
8000– 9000	30.0	70000–75000	16	180000–190000	2
9000–10000	29.0	75000–80000	15	190000–200000	1
10000–15000	28.0	80000–85000	14	200000 and over	0
15000–20000	27.0	85000–90000	13		

NOTE. — When the number of bacteria per cubic centimeter exceeds the local legal limit, the score shall be 0.

SCORE CARD FOR DAIRY FARMS

Farm of.......................................Location..............................Date..............

EQUIPMENT	SCORE	
	Perfect	Allowed
COWS, 16		
Comfort:		
Bedding	1	—
Temp. of stable	1	—
Food	2	—
Water:		
Clean	1	—
Fresh	1	—
Light (4 sq. ft. per cow)	4	—
Ventilation, automatic (adjustable windows)	3	—
Air space per cow, 600–1000 cu. ft. (Less than 600, 2; less than 500, 1; less than 400, 0)	3	—
STABLES, 6		
Location of stable:		
Well drained	1	—
Free from contaminating surroundings	1	—
Construction of stable:		
Tight, sound floor and proper gutter	2	—
Smooth, tight walls and ceiling	1	—
Stall, tie, and manger	1	—
UTENSILS, 14		
Construction of utensils	1	—
Water for cleaning (clean, convenient, and sufficient)	2	—
Facilities for steam (Hot water enough to immerse utensils, 2)	4	—
Small-top milk-pail	3	—
Milk cooler	2	—
Clean milking suits	2	—
HANDLING THE MILK, 4		
Location of milk room:		
Free from contaminating surroundings	1	—
Convenient	1	—
Construction of milk room:		
Floors, walls, and ceiling	1	—
Light, ventilation, and screens	1	—
Total	40	

METHODS	SCORE	
	Perfect	Allowed
COWS AND STABLES, 16		
Cleanliness of stables:		
Floor	2	—
Walls	1	—
Ceiling and ledges	1	—
Mangers and partitions	1	—
Windows	1	—
Stable air	6	—
Barnyard (clean and well drained)	2	—
Manure removed daily to field or pit (Manure stored less than 50 ft. from stable, 0)	2	—
UTENSILS AND MILKING, 24		
Care and cleanliness:		
Thoroughly cleansed	6	—
Scalding utensils	6	—
Inverting utensils in pure air and sun	3	—
Cleanliness of milking:		
Clean, dry hands	3	—
Udders washed and dried (Udders cleaned with moist cloth, 4)	6	—
HANDLING MILK, 20		
Cleanliness of attendant	2	—
Milk of each cow removed at once from stable	2	—
Cleanliness of milk room	3	—
Prompt cooling	2	—
Efficient cooling: below 50° F. (51–55° F., 4; 56–60° F., 2)	5	—
Storage: below 50° F. (51–55° F., 2; 56–60° F., 1)	3	—
Transportation: iced (For jacket or wet blanket allow 2; dry blanket or covered wagon, 1)	3	—
Total	60	

Score for equipment.... + score for methods ......= .... final score.

Source of water supply.........................General conditions of farm.........................

Remarks:

...., *Inspector.*

SANITARY INSPECTION OF CITY MILK PLANTS

Owner or manager...Trade name...................

City.........................Street and No...........................State.......................

Number of wagons........................Gallons sold daily { Milk.............................. / Cream..............................

Permit or License No..........................Date of inspection..........................

Equipment	Score	
	Perfect	Allowed
Building, 16		
Location: free from contaminating surroundings	2	—
Arrangement, 6		
Separate receiving room	1	—
Separate handling room	2	—
Separate wash room	1	—
Separate sales room	1	—
Separate boiler room	1	—
Construction, 8		
Floors tight, sound, cleanable	1	—
Walls tight, smooth, cleanable	1	—
Ceilings smooth, tight, cleanable	1	—
Provisions for light	1	—
Provisions for pure air	1	—
Screens	2	—
Minimum of shafting, pulleys, hangers, exposed pipes, etc	1	—
Apparatus, 20		
Boiler	2	—
Hot-water heater	1	—
Milk cooler	2	—
Refrigerator	2	—
Appliances for cleansing utensils and bottles	2	—
Racks, etc., for utensils and bottles after cleaning	1	—
Sterilizer for utensils and bottles	2	—
Bottling and capping machine	1	—
Wash bowl, soap, and towel for attendants	2	—
Protection during delivery	2	—
Condition of apparatus (make deduction for inaccessible parts, open seams, rusty ware, decayed or battered tables or sink, milk-carrying pipes with rough interiors and lack of frequent hand couplings, and for badly worn and poorly repaired material)	4	—
Laboratory and Equipment		
Water Supply, 2		
Clean, fresh	1	—
Convenient and abundant	1	—
Total	40	—

Methods	Score	
	Perfect	Allowed
Building, 15		
Cleanliness:		
Floors	3	—
Walls	1	—
Ceilings	2	—
Doors and windows	1	—
Shafting, pulleys, hangers, pipes	1	—
Freedom from odors	2	—
Freedom from flies and other insects	3	—
Drainage	2	—
Apparatus, 16		
Cleanliness:		
Thoroughly washed and rinsed	6	—
Sterilized in live steam, thirty minutes (Thoroughly scalded after washing with water over 200° F. or live steam, 3)	5	—
Bottle caps sterilized	3	—
Protected from dirt	2	—
Handling Milk, 16		
Received below 50° F. (50–55° F., 4; 55–60° F., 3)	5	—
Rapidity of handling in plant	3	—
Freedom from undue exposure to air in the plant	2	—
Capping bottles by machine	1	—
Bottle top and cap protected by covering	2	—
Storage 45° F. or below (45–50° F., 2; 50–55° F., 1)	3	—
Inspection, 9		
Bacteriological work	4	—
Inspection of dairies supplying milk (Once in a year, 1; twice a year, 2; three times a year, 3; four times a year, 4)	5	—
Miscellaneous, 4		
Cleanliness of attendants (General appearance, hands, etc., 1; clean, washable clothing, 1)	2	—
Cleanliness of delivery outfit	2	—
Total	60	—

Score for equipment...............+ score for methods.................= total score................

Note. — If the conditions in any particular are so exceptionally bad as to be inadequately expressed by a score of "0," the inspector can make a deduction from the total score.

.... *Inspector.*

SCORE CARD FOR DAIRY FARMS

Farm of.................................Location..............................Date................

EQUIPMENT	SCORE	
	Perfect	Allowed
COWS, 16		
Comfort:		
Bedding........................	1	—
Temp. of stable................	1	—
Food...........................	2	—
Water:		
Clean..........................	1	—
Fresh..........................	1	—
Light (4 sq. ft. per cow)...........	4	—
Ventilation, automatic (adjustable windows)	3	—
Air space per cow, 600–1000 cu. ft. (Less than 600, 2; less than 500, 1; less than 400, 0)	3	—
STABLES, 6		
Location of stable:		
Well drained...................	1	—
Free from contaminating surroundings........................	1	—
Construction of stable:		
Tight, sound floor and proper gutter........................	2	—
Smooth, tight walls and ceiling...	1	—
Stall, tie, and manger...........	1	—
UTENSILS, 14		
Construction of utensils...........	1	—
Water for cleaning (clean, convenient, and sufficient)...........	2	—
Facilities for steam................ (Hot water enough to immerse utensils, 2)	4	—
Small-top milk-pail...............	3	—
Milk cooler.......................	2	—
Clean milking suits...............	2	—
HANDLING THE MILK, 4		
Location of milk room:		
Free from contaminating surroundings........................	1	—
Convenient.....................	1	—
Construction of milk room:		
Floors, walls, and ceiling.........	1	—
Light, ventilation, and screens....	1	—
Total.........................	40	

METHODS	SCORE	
	Perfect	Allowed
COWS AND STABLES, 16		
Cleanliness of stables:		
Floor..........................	2	—
Walls..........................	1	—
Ceiling and ledges..............	1	—
Mangers and partitions..........	1	—
Windows.......................	1	—
Stable air........................	6	—
Barnyard (clean and well drained)	2	—
Manure removed daily to field or pit (Manure stored less than 50 ft. from stable, 0)	2	—
UTENSILS AND MILKING, 24		
Care and cleanliness:		
Thoroughly cleansed............	6	—
Scalding utensils................	6	—
Inverting utensils in pure air and sun..........................	3	—
Cleanliness of milking:		
Clean, dry hands...............	3	—
Udders washed and dried........ (Udders cleaned with moist cloth, 4)	6	—
HANDLING MILK, 20		
Cleanliness of attendant...........	2	—
Milk of each cow removed at once from stable...................	2	—
Cleanliness of milk room...........	3	—
Prompt cooling...................	2	—
Efficient cooling: below 50° F...... (51–55° F., 4; 56–60° F., 2)	5	—
Storage; below 50° F.............. (51–55° F., 2; 56–60° F., 1)	3	—
Transportation: iced.............. (For jacket or wet blanket, allow 2; dry blanket or covered wagon, 1)	3	—
Total.........................	60	

Score for equipment .... + score for methods...... = ... final score.

Source of water supply.........................General conditions of farm..........................

Remarks:

...................................*Inspector.*

SANITARY INSPECTION OF CITY MILK PLANTS

Owner or manager...Trade name...................

City.........................Street and No.........................State.......................

Number of wagons.......................Gallons sold daily { Milk................................ / Cream................................

Permit or License No.........................Date of inspection..........................

Equipment	Score	
	Perfect	Allowed
Building, 16		
Location: free from contaminating surroundings	2	—
Arrangement, 6		
Separate receiving room	1	—
Separate handling room	2	—
Separate wash room	1	—
Separate sales room	1	—
Separate boiler room	1	—
Construction, 8		
Floors tight, sound, cleanable	1	—
Walls tight, smooth, cleanable	1	—
Ceilings smooth, tight, cleanable	1	—
Provisions for light	1	—
Provisions for pure air	1	—
Screens	2	—
Minimum of shafting, pulleys, hangers, exposed pipes, etc	1	—
Apparatus, 20		
Boiler	2	—
Hot-water heater	1	—
Milk cooler	2	—
Refrigerator	2	—
Appliances for cleansing utensils and bottles	2	—
Racks, etc., for utensils and bottles after cleaning	1	—
Sterilizer for utensils and bottles	2	—
Bottling and capping machine	1	—
Wash bowl, soap, and towel for attendants	2	—
Protection during delivery	2	—
Condition of apparatus (make deduction for inaccessible parts, open seams, rusty ware, decayed or battered tables or sink, milk-carrying pipes with rough interiors and lack of frequent hand couplings, and for badly worn and poorly repaired material)	4	—
Laboratory and Equipment		
Water Supply, 2		
Clean, fresh	1	—
Convenient and abundant	1	—
Total	40	—

Methods	Score	
	Perfect	Allowed
Building, 15		
Cleanliness:		
Floors	3	—
Walls	1	—
Ceilings	2	—
Doors and windows	1	—
Shafting, pulleys, hangers, pipes	1	—
Freedom from odors	2	—
Freedom from flies and other insects	3	—
Drainage	2	—
Apparatus, 16		
Cleanliness:		
Thoroughly washed and rinsed	6	—
Sterilized in live steam, thirty minutes (Thoroughly scalded after washing with water over 200° F. or live steam, 3)	5	—
Bottle caps sterilized	3	—
Protected from dirt	2	—
Handling Milk, 16		
Received below 50° F. (50–55° F., 4 ; 55–60° F., 3)	5	—
Rapidity of handling in plant	3	—
Freedom from undue exposure to air in the plant	2	—
Capping bottles by machine	1	—
Bottle top and cap protected by covering	2	—
Storage 45° F. or below (45–50° F., 2 ; 50–55° F., 1)	3	—
Inspection, 9		
Bacteriological work	4	—
Inspection of dairies supplying milk (Once in a year, 1 ; twice a year, 2; three times a year, 3; four times a year, 4)	5	—
Miscellaneous, 4		
Cleanliness of attendants (General appearance, hands, etc., 1; clean, washable clothing, 1)	2	—
Cleanliness of delivery outfit	2	—
Total	60	—

Score for equipment...............+ score for methods.................= total score................

Note. — If the conditions in any particular are so exceptionally bad as to be inadequately expressed by a score of "0," the inspector can make a deduction from the total score.

.... *Inspector.*

REPORT OF CHEMICAL ANALYSIS

EXPERIMENT

Partial analysis of..

Obtained from..

Manufactured by..

Exact name of brand..

Tested for..

Description of test..

....

....

....

....

Equations of test..

....

....

Found..

....

Quality of sample..

Remarks..

....

....

I hereby certify that the above report is correct to the best of my knowledge.

(Signed)....................................

Date................................

REPORT OF CHEMICAL ANALYSIS

EXPERIMENT

Partial analysis of..

Obtained from..

Manufactured by..

Exact name of brand..

Tested for..

Description of test..

....

....

....

....

Equations of test..

....

....

Found..

....

Quality of sample..

Remarks..

....

....

I hereby certify that the above report is correct to the best of my knowledge.

(Signed)..

Date..............................

EXPERIMENT

Purpose: State it on a separate sheet used for the record.
Materials: White woolen cloth.

Artificial colors are used to give foods a more attractive appearance. If the coloring material is harmless, the use of it may be legitimate. Vegetable colors are in the harmless class. Coal-tar is the source of a very large number of colors, many of which are known to be harmful when eaten with foods. This is particularly true if the dyestuff has been prepared by the use of such metals as mercury, lead, copper, tin, arsenic, antimony, zinc, chromium, and barium. Abuses have grown with the coloring of foods, the chief one being to conceal inferiority or adulteration. The laws and food inspectors heartily condemn this practice. Since it is difficult for the average purchaser to learn what colors are harmful, the laws require manufacturers to declare on the label the presence of artificial color. This brings the problem to the consumer to decide, whether the practice shall continue. Good foods, suitably prepared, cannot be made more wholesome by coloring, so artificially colored foods are open to suspicion.

Coal-tar colors are likely to be found in ice cream, gelatin, dessert powders, colored sugars, Maraschino cherries, flavoring extracts, jams, jellies, catsup, wine, soft drinks, preserves, macaroni, vinegar, and canned vegetables.

I. **Detection of Coal-Tar Colors**. Practically all coal-tar dyes suitable for food colors are substantive dyes which are readily taken up by wool cloth. This property is used to detect them.

If the sample is a liquid, use 30 to 60 cc., depending on the intensity of the color in the sample. If it is a solid, take from 10 to 25 grams and mix it with enough water to make it fairly liquid. Add about 3 cc. of dilute hydrochloric acid, stir well, and bring the mixture to the boiling point. Put in the mixture an 8-cm. square of white woolen cloth (nun's veiling or albatross cloth) that has been boiled for 5 minutes in a 1 % solution of sodium hydroxid and rinsed to free the cloth from grease of any kind. Boil the cloth in the mixture until it is well colored. If the cloth is not colored after 10 minutes' boiling, the absence of coal-tar dye is fairly certain.

Remove the cloth and wash it in water, then boil the cloth for a few minutes in about 50 cc. of water to which 17 cc. of dilute hydrochloric acid has been added. Rinse the cloth and place it in a bath of 50 cc. of water and 4 cc. of dilute ammonium hydroxid. This reagent will dissolve and remove most of the color. The cloth may now be wrung out and discarded.

Make the colored solution slightly acid with 2 % hydrochloric acid and boil a 2 × 3-cm. piece of similar cloth in it. The depth of color in the second piece of cloth indicates the amount of the dye present. Mount this piece of cloth neatly on the record sheet and label it.

II. **Special Tests**. (*a*) Coal-tar Dye in Milk. Add 5 cc. concentrated hydrochloric acid to an equal volume of milk, and mix thoroughly. A pink color shows azo-orange dye.

(*b*) Coal-tar Dye in Butter, Low's Test. Melt a small spoonful of butter in a test-tube, taking care not to scorch it. Add Low's reagent in volume equal to the melted butter and shake the mixture. Heat almost to boiling and let the tube stand quietly until the acid settles to the bottom. The acid will be wine red if azo colors are present. Pure butter gives only a faint blue tinge to the acid.

(*c*) Coal-tar Dye in Lemon Extract, Lythgoe's Test. To a little of the sample add three drops of concentrated hydrochloric acid. Partial or complete decoloration or a pink color indicates that the dye is present.

Ten cubic centimeters may be evaporated to dryness, the residue dissolved in water, and tested by the double dye process, Part I.

Study Topics: Hesse, *Coal-Tar Colors Used in Food Products. The Use and Effects of Coal-Tar Dyes in Foods.* Bulletin No. 147, U. S. Dept. of Agriculture.
Leach, *Food Inspection and Analysis*, pp. 780–814.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Materials: Thionine or wool blue or Bismarck brown, Congo red, primuline, beta-naphthol, resorcin, toluene diamine, indigo powder or paste, fuchsine, methyl violet, malachite green, methylene blue, sodium hydroxid sticks, tannic acid, sodium carbonate, sodium chlorid, sodium nitrite, quicklime, zinc dust, tartar emetic, 30-cm. lengths of white cotton and wool yarn.

The dyeing of fabrics has become a great industry. The dyes prepared from coal tar have displaced most of those of vegetable origin formerly in use because the derived colors are more brilliant and permanent. The processes of dyeing require careful chemical control and depend on the properties of the fiber and the particular dye used.

I. **Direct Dyes**. (*a*) For Wool. The cotton should be thoroughly wet before dyeing. Boil for 5 minutes in an evaporating dish a piece each of white wool and cotton yarn in 50 cc. of a solution of thionine, or wool blue, or Bismarck brown (1 g. per L.), to which a drop of acetic acid has been added. Remove the yarn, wash well, and let it dry. Compare the intensity of color of each kind of yarn and account for the difference (?). Double each piece of yarn twice and fasten it to the record sheet by a strip of gummed paper. Label suitably each piece mounted. Mount similarly original samples of the yarn as well as all other pieces dyed. Lighter tints of the same color can be made by various methods such as diluting the dye, lowering the temperature, and shortening the length of time in the bath. Vary one of these conditions at a time and dye other pieces of wool.

(*b*) For Cotton. Boil a piece of cotton yarn in 50 cc. of Congo red solution (2 g. per L.) for 5 minutes after adding a very small crystal of sodium carbonate. Wash, dry, and mount the yarn. Compare the intensity of this color and any of the previous dyes on cotton (?).

Questions: 1. What is the meaning of the term " direct dye "?

2. What difference in properties has been shown in the dyes and the kinds of fibers used?

(*c*) Developing a Dye. Boil three pieces of cotton yarn for 5 minutes in 50 cc. of primuline solution (1 g. per L.) after adding a pinch of sodium chlorid. Remove the yarn, wash and dry one piece. Put the other two into the **cold** diazotizing bath containing 1 g. of sodium nitrite per L., to which 3 cc. of concentrated hydrochloric acid is added just before dipping the yarn into it. Let the yarn remain in bath 10 minutes, then remove and at once put one piece each into the following developers: (1) beta-naphthol and sodium hydroxid, each 0.5 g. per L.; (2) resorcin 0.5 g., sodium hydroxid 1 g. per L. Stir the yarn and watch the color change (?). The diazotizing reagent is nitrous acid. Write the equation for its formation from the reagents used (?). Another developer is toluene diamine, 1 g. per L., with sodium carbonate, 2 g. per L.

(*d*) Insoluble Dye. Indigo is reduced to the colorless calcium salt. In this form it is soluble and is used in the dye bath. On exposure to air it oxidizes to the colored form. Mix 10 g. of zinc dust with 150 cc. of warm water. Add 10 g. of indigo powder or 25 g. of paste, and then 10 g. of quicklime while stirring. It should stand covered for 6 hours or more with an occasional stirring. Dilute the solution with 500 cc. of distilled water and let each pupil saturate thoroughly 2 pieces of cotton yarn in it. Wring out and dry. Watch the color form (?). Dip one piece in the bath a second time and dry (?).

II. **Mordant Dyes**. (*a*) Boil a separate piece of cotton yarn in 50 cc. each of the following solutions, each 1 g. per L.: fuchsine, methyl violet, malachite green, and methylene blue. Wash, dry, and mount the samples. Mordant four other pieces by boiling them for 5 minutes in a solution of tannic acid, 1 g. per L. Transfer them to a cold bath of tartar emetic, 1 g. per L., to " fix " for 5 minutes, then put one in each of the above dyes and boil for 5 minutes. Wash, dry, and mount the yarn.

Study Topics: Martin, *Modern Chemistry and Its Wonders*: "The Romance of Coal Tar," pp. 262–290.

TEXTILES

Clothing stands next to food among the necessities of the human race. The textile industry is concerned with the preparation of the fibers, spinning the thread, and weaving the cloth. Much chemical treatment is necessary in preparing the fiber and in bleaching and dyeing the cloth.

In recent years many processes, such as mercerizing cotton, have been developed by which the properties of the fibers are remarkably changed. Many attempts are being made to change cotton, the cheapest fiber, into substitutes for the more costly linen, wool, or silk. Textile manufacturers have succeeded in altering the appearance of cotton to such an extent that the inexperienced buyer is often deceived in buying clothing. The fibers can always be identified, however, by a chemical or microscopical examination.

Commercial terms are often deceiving. The term " all wool " is applied to fabrics that may be only 80 % wool, " strictly all wool" to 90 % wool fabrics, while it takes the term " absolutely all wool " to mean really **all wool.**

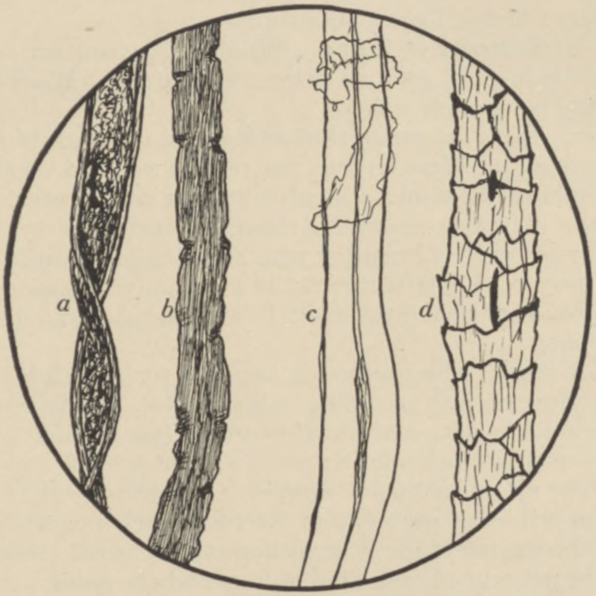


Fig. 103. Textile fibers : *a*, cotton ; *b*, linen ; *c*, silk ; *d*, wool.

Study Topics: Sadtler, *Chemistry of Familiar Things:* " Textiles," pp. 281–285.
Baskerville, *Municipal Chemistry:* " Textile Materials and Their Service to Man," pp. 374–395.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Apparatus: Microscope, balance.

Materials: Cloth to be tested (preferably from home), and the following solutions: concentrated zinc chlorid, lead acetate, chlorin, nitric acid (fuming), iodin in potassium iodid, Schweitzer's reagent, picric acid, hydrochloric acid (1 %), sodium hydroxid (2 %), fuchsin in alcohol (1 %), nickel hydroxid.

Preliminary: Examine fibers of cotton, wool, silk, and flax with a microscope.

I. **Qualitative.** Make tests according to the scheme of analysis on the following page on one or more samples of cloth. By these tests the different kinds of fibers present can be determined. Mount neatly a 4 × 4-cm. piece of the fabric tested on the record sheet, give the name of the cloth, and the use made of it at home. Divide the record sheet into three vertical columns and give them these headings respectively: Tests, Results, Conclusions.

II. **Quantitative.** (*a*) **Removal of Sizing.** Weigh a 100 cm. sq. portion carefully to decigrams, and boil it 10 minutes in a 1 % solution of hydrochloric acid. Wash and dry it well and weigh it as fiber only. Calculate per cent of sizing.

(*b*) **Wool and Cotton.** After the removal of sizing, boil the piece of cloth for 10 minutes in a 2 % solution of sodium hydroxid. Wash and dry the residue well and weigh as cotton. The loss in weight represents the weight of the wool. Calculate the per cent of each.

(*c*) **Cotton and Linen.** After the removal of sizing, wet a small piece of the cloth in a 1 % alcoholic solution of fuchsin, wash well and moisten with ammonium hydroxid. The flax fibers will be rose-red and the cotton uncolored. With the aid of a magnifying glass or low-power microscope count the number of colored and uncolored threads in a given space and determine the approximate per cent of each kind of fiber.

(*d*) **Silk, Wool, and Cotton.** This method is accurate for all fabrics except plush. Remove sizing from a 100 cm. sq. piece of cloth according to Section (*a*). Remove the silk by a two-minute treatment of the cloth in a cold solution of nickel hydroxid (see p. 255). After this treatment boil the residue in a 1 % solution of hydrochloric acid for two minutes, wash and dry it well, and weigh as wool and cotton only. The nickel solution is necessary to separate silk from wool. If silk only is present with the cotton, or if it is not necessary to determine each separately, the wool as well as the silk may be removed by the treatment given in Section (*b*). The final residue dried and weighed is the cotton. Calculate the per cent of each kind of fiber and the sizing.

REPORT OF CHEMICAL ANALYSIS

EXPERIMENT

Partial analysis of..

Obtained from..

Manufactured by..

Exact name of brand..

Tested for..

Description of test..

....

....

....

....

Equations of test..

....

....

Found..

....

Quality of sample..

Remarks..

....

....

I hereby certify that the above report is correct to the best of my knowledge.

(Signed)..

Date..................................

ANALYSIS OF MIXED FIBERS (PINCHON)

<table>
<tr><td colspan="10">Boiled with 10% KOH for 20 minutes</td></tr>
<tr><td colspan="3">Completely dissolves</td><td colspan="4">Remains undissolved</td><td colspan="3">Partially dissolved</td></tr>
<tr><td colspan="10">Another sample boiled 5 minutes in $ZnCl_2$ solution</td></tr>
<tr><td>All dissolves</td><td>Part dissolves</td><td>None dissolves</td><td colspan="4">None dissolves</td><td colspan="2">Part dissolves</td><td>None dissolves</td></tr>
<tr><td>The alkalin solution does not blacken on addition of lead acetate</td><td>The dissolved portion is not blackened by lead acetate, but the insoluble portion is blackened</td><td>The mass is blackened by lead acetate</td><td colspan="4">Chlorin water or NH_3 colors fibers</td><td colspan="2">On addition of lead acetate is</td><td>HNO_3 colors a portion yellow, the residue remaining white</td></tr>
<tr><td></td><td></td><td></td><td>Red-brown</td><td colspan="3">Not at all</td><td>Blackened</td><td>Not blackened</td><td></td></tr>
<tr><td></td><td></td><td></td><td>The fiber is colored red by fuming nitric acid</td><td colspan="3">Alcoholic fuchsin colors fibers</td><td>KOH partly dissolves the fibers insoluble in $ZnCl_2$; the remaining fibers are soluble in Schweitzer's reagent</td><td>Picric acid colors a portion yellow, the residue remaining white</td><td></td></tr>
<tr><td></td><td></td><td></td><td></td><td colspan="2">Permanently</td><td>Color removed by washing</td><td></td><td></td><td></td></tr>
<tr><td></td><td></td><td></td><td></td><td colspan="2">KOH stains it yellow Iodin and H_2SO_4 colors it</td><td>No yellow color with KOH</td><td></td><td></td><td></td></tr>
<tr><td></td><td></td><td></td><td></td><td>Yellow</td><td>Blue</td><td></td><td></td><td></td><td></td></tr>
<tr><td><i>Silk</i></td><td><i>Silk and Wool</i></td><td><i>Wool</i></td><td><i>New Zealand Flax</i></td><td><i>Hemp</i></td><td><i>Flax</i></td><td><i>Cotton</i></td><td><i>Wool, Silk, and Cotton</i></td><td><i>Silk and Cotton</i></td><td><i>Flax and Cotton</i></td></tr>
</table>

TREATMENT OF STAINS

If the nature of the stain is known, and is not too old, and not given the wrong treatment, it is possible to remove it successfully. Three general methods of treatment are as follows:

I. **Warm-water Method**. Sugar, glue, blood; fruit, tea, and coffee stains if fresh. All except blood stains are removed more quickly with boiling water.

II. **Special Solvent Method**. Grease, oil, and wax. Water has practically no effect in such cases, but these substances dissolve readily in chloroform, carbon tetrachlorid, ether, gasoline, naphtha, benzine, kerosene, and benzol. The first three are rather expensive. If the stains are paint, varnish, or tar, they are first treated with turpentine and then with the above solvents. All of the solvents except the first two are inflammable, so they should be used only in daytime and outdoors.

III. **Bleaching Method**. Ink is a typical example of these refractory stains that require chemical treatment. It will be recalled that in the dyeing experiments some dyes were fast only when a mordant was used. Ink usually contains an iron salt as a mordant. Oxalic and hydrochloric acids and bleaching powder are the reagents used. After this treatment the cloth should be washed thoroughly with water or sponged with dilute ammonia to remove traces of the reagents, which will rot the fibers. A dilute solution of " hypo " (sodium thiosulfate) will remove traces of the bleaching powder.

IV. **Manner of Applying Reagents.** If water is to be used, the cloth is placed over a convenient vessel and water poured **around the stain first**, then on it. Other solvents are applied in the same manner, but if in small quantity, then blotting paper or other absorbent material should be placed under the fabric treated. This will absorb the solvent as it dissolves the grease. If the stain spreads or traces are left, use gelatinized benzene or benzolized magnesia.

V. **Special Reagents. Gelatinized Benzene** (not benzine). Mix about 10 cc. of a concentrated solution of castile soap with 100 cc. of benzene (common name, benzole), add a little ammonia to the mixture, and shake it well. Use as any other solvent.

Benzolized Magnesia. Mix calcined magnesia (magnesium oxid) with enough benzene to form a paste. Spread this over the stain and rub it in well. Brush it off and repeat the process until the stain is removed.

Javelle Water. Since bleaching powder deteriorates rapidly when exposed to air, it may be converted into Javelle water, which keeps longer. Dissolve one pound of washing soda (sodium carbonate) in a quart of hot water. Add to it one-fourth pound of bleaching powder (chlorid of lime) stirred up in a quart of hot water. After mixing, let the undissolved particles settle and pour off the clear solution of sodium hypochlorite into bottles for future use. The sediment may be used for scouring and disinfecting.

Ink Eradicator. A two-solution preparation is sold by most stationers for removing ink. It can be prepared cheaply according to the following formula:

SOLUTION 1	
Citric acid	6 grams
Water	50 cc.
Borax (sat. sol.)	12 cc.

SOLUTION 2	
Chlorid of lime	18 grams
Water	50 cc.
Borax (sat. sol.)	12 cc.

Mix the chlorid of lime and water and shake it. Let the mixture settle in a well-stoppered container for a week, or filter; add the borax solution. Keep solutions in separate bottles. To remove ink, apply Solution 1 first and dry with a blotter. Then apply Solution 2. Fresh ink yields to one treatment. If used on clothing, be sure to sponge the place well with dilute ammonia as a final treatment.

An old method of removing rust stains is to use salt and lemon juice. These two substances react to form small quantities of hydrochloric acid. This reacts with iron rust to form soluble iron chlorid, which may be washed out with water.

REMOVAL OF STAINS

Kind of Stain	White Linen and Cotton Goods	Colored Cotton Goods
Sugar, glue, blood......	Wash with warm water; if not entirely removed, use soap and ammonia	
Grease................	Naphtha, benzene, or gelatinized benzene. Follow with warm soap water and ammonia	
Paint and varnish......	Turpentine, followed by the method for grease	
Tar and pitch.........	Rub with sweet oil, then apply the paint treatment	
Iron rust and gall ink...	Warm oxalic or dilute hydrochloric acids. Salt and lemon	Glycerin and soft soap
Fruit and wine, tea and coffee...............	Hot water. Javelle water	Soap and ammonia
Colored ink...........	Fresh stains, alcohol. Javelle water	Alcohol. Javelle water if the color is not affected
Acids.................	Ammonia neutralizes the acid and restores the color. Wash well with water	
Wax stains............	Scrape off exterior wax and dissolve the remainder with benzene	
Grass.................	Sponge with alcohol or ammonia	
Mildew...............	Soft soap mixed with chalk and salt. Chlorid of lime. Sunlight	

Woolen and silk goods are treated by the same methods as for cotton goods, with a few exceptions. Greater care should be exercised.

WOOLEN AND SILK GOODS

Kind of Stain	Colored Woolen Goods	White Woolen Goods	Silk Fabrics
Grease................	Soap and ammonia	Soap and ammonia	Ether, special benzene preparations
Iron and gall ink stains..	Citric or dilute hydrochloric acid if the colors of the cloth are not affected, otherwise use a mixture of glycerin and soft soap		

French Dry-cleaning Fluid. Make a soap jelly of two ounces of finely shaved naphtha or castile soap in one pint of distilled water. Add to it one-fourth ounce each of honey, gum camphor, borax, turpentine, saltpeter, one-fourth pint of alcohol, and one-half ounce of sulfuric ether. Mix all ingredients well and add a pint of hot water. Bottle the solution for future use. It will not harm any fabric.

Reagent for Mildew Stain. Add to the juice of a tomato and a lemon, one-half ounce each of salt, borax, and starch. Mix with a soap jelly. Apply the reagent hot and leave the fabric in the sun.

Paint and Varnish Removers. These mixtures are sold at retail paint stores. They can be obtained more cheaply by purchasing and mixing the ingredients.

FORMULA 1	FORMULA 2	FORMULA 3
200 cc. of anilin	100 cc. of amyl acetate	100 cc. of methyl alcohol
200 cc. of benzol	100 cc. of anilin	100 cc. of benzol
10 cc. of nitro-benzol	200 cc. of benzol	

Reference Text: Woolman and McGowan, *Textiles.*

TREATMENT OF STAINS

EXPERIMENT

Put the more common stains on various kinds of cloth, and treat the stain according to the above methods. Mount the pieces of cloth neatly, after trimming them to uniform size. Label each, stating the kind of treatment applied. Such a record will be of value when one wishes to know the result before applying the treatment. The inexperienced person should always treat the same stain on a separate piece of the same goods before attempting to remove the stain from a garment.

ANALYSIS OF COAL

The comparative values of different coals can be determined by analysis to find the per cent of moisture, volatile matter, fixed carbon (not volatilized when heated without contact with air), and ash. Sometimes the per cent of sulfur is determined, because coals containing much sulfur are very undesirable for many industries.

The water in coal may vary from less than one to over 20 %. The mineral matter left as ash varies from 1 to 6 % in most good anthracite and bituminous coals.

Bituminous coals have a greater heating power than anthracite, but they require especially designed furnaces to secure this great heating power, on account of the volatile matter which distills off when the fresh coal is put on the fire. Much heat is necessarily used in changing this part of the coal into a gas. If insufficient air is admitted to the combustion chamber, this gas escapes unburned, and the efficiency of the coal is considerably reduced. The volatile substances in the coal are more or less decomposed by the intense heat of the combustion chamber, and a great deal of free carbon escapes into the air with the unburned gases to pollute the air with smoke and soot.

EXPERIMENT

Purpose:...
Apparatus: Crucible.
Materials: Coal (preferably from home).

I. **Water.** Reduce several pieces of coal to particles the size of a pea, and mix thoroughly. Reduce a teaspoonful of this coal to a very fine powder. Weigh accurately a clean, dry crucible and cover. Place about a cubic centimeter of the powdered coal in the crucible and weigh again. Record weights below. Place the uncovered crucible in an air-bath and heat it for one hour at 105° C. When cool, weigh the crucible (including cover), and record. Calculate the per cent of moisture.

II. **Volatile Matter.** Cover the crucible and contents from Part I, set it on a triangle, and heat directly with the burner until no more smoke is given off. Do not heat or cool the crucible too suddenly, but during most of the heating a large hot flame should be used so that the entire crucible will be red hot. Let the crucible stand covered until cool, weigh, and calculate the per cent.

III. **Fixed Carbon.** Uncover the crucible, incline it on the triangle, and heat to a full red heat until **all** the carbon has burned off. Burn off any deposit on the cover in the same way. Stir the ash with a clean platinum wire to see whether there is any unburned carbon beneath. Be careful to remove none of the ash. Cool, weigh (with the cover), and calculate the per cent of fixed carbon.

IV. **Ash.** The residue from Part III is the non-volatile mineral matter of the coal, or the ash. Calculate its per cent.

RECORD OF RESULTS

1. Weight of crucible and cover.......................................g.
2. Weight of crucible, cover, and coal..................................g.
3. Weight of coal...g.
4. Weight of crucible, cover, and residue from I..........................g.
5. Weight of water in coal..g.
6. Weight of crucible, cover, and residue from II.........................g.
7. Weight of volatile matter in coal....................................g.
8. Weight of crucible, cover, and ash..................................g.
9. Weight of fixed carbon in coal......................................g.
10. Weight of ash..g.

Percentages:

Per cent of water (5 ÷ 3 × 100)......................................%
Per cent of volatile matter (7 ÷ 3 × 100)..............................%
Per cent of fixed carbon (9 ÷ 3 × 100)................................%
Per cent of ash (10 ÷ 3 × 100)......................................%

COAL

The industrial world obtains its motive power largely from coal. On account of the cost, anthracite coal is practically limited in use to domestic heating. Although good bituminous coal has a greater heating power than anthracite, it is very difficult to prevent loss of heating value through the escape of some of the gases distilled from it by the heat of the furnace. The escape of these gases unburned is further objectionable on account of the smoke and soot. Even with the best of modern smoke consumers, and with the most efficient steam engines, only about 27% of the heat energy of the coal is converted into mechanical energy. From the present increasing consumption of coal, it seems probable that the supply of steam coal will be exhausted within the next 50 years. Water power can be further developed in the United States to the extent of $25,000,000, the amount now going to waste each year. Petroleum is furnishing much power by the use of its more easily vaporized oils in the gas engine. This form of fuel is adapted to use in automobiles, but its supply is limited, so that we cannot hope to make use of it as a substitute for coal.

There are immense deposits of lignite in many parts of the country, but the difficulty of using it in present-day power plants is so great that practically none of it is mined. Yet this despised lignite seems to be the hope of the future. By subjecting lignite to dry distillation, considerable gas is produced which can be used in a gas engine, whose efficiency is greater than the steam engine. The gas can be piped around to different establishments, but it will probably be more economical to use the gas in the central power house where it is generated, driving gas engines, which in turn drive great dynamos. The electricity from these dynamos can be transmitted over a large area of country. Actual tests on such plants have shown that low-grade lignite coals used in this way will produce more power than the very best steaming coals do in the most modern steam plants.

Study Topics: Mitchell, "A New Source of Power": *Nat. Geog. Mag.*, Nov. 1910, pp. 935–944.
Baskerville, *Municipal Chemistry:* "Illuminating Gas," pp. 300–321.

DESTRUCTIVE DISTILLATION OF COAL

EXPERIMENT

I. **Destructive Distillation.** (Two students or teacher.) Fill the test-tube (preferably a hard-glass tube) three-fourths full of small pieces of soft coal. Insert the stopper and arrange the apparatus as shown in Figure 104. Heat the tube carefully and uniformly its entire length, raising its temperature gradually until no more gas is given off. Collect two small bottles of the gas by displacement of water. Several times during the heating test the stream of gas issuing from the tube for hydrogen sulfid with a piece of paper moistened in lead acetate solution (?), and also for ammonia by means of moistened red litmus paper (?).

II. **Solid, Liquid, and Gaseous Products.** (*a*) **Gas.** Use one bottle of the gas to determine the physical properties of the gas, and the other to determine its relation to combustion (?).

(*b*) **Liquid.** Note the properties of the matter in the generating bottle (?), and name it (?).

(*c*) **Solid.** When the residue in the test-tube is cold, examine it carefully, noting its physical properties (?), and its relation to combustion (?). Name it.

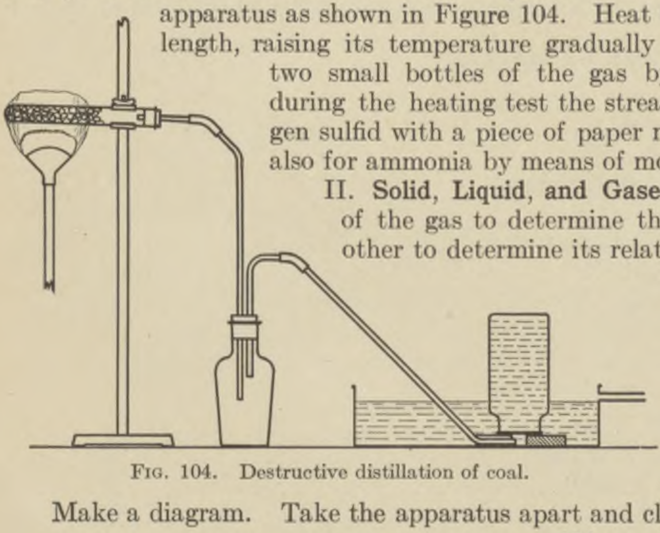


FIG. 104. Destructive distillation of coal.

Make a diagram. Take the apparatus apart and clean it with kerosene.

SOAP MAKING

Soap has become a necessity of a civilized age. Many kinds are made to serve the various needs of a diversified world. All soaps may be grouped into two classes, the toilet and the laundry soap. Since the former class is for personal use, the materials used in making it should be of good quality, clean, and pure. When the soap is made, it should have nothing in it that would be injurious to the skin, such as " free alkali " (free hydroxyl ions).

Cheaper fat stock is used for the laundry soap, but no one desires soap so cheap as that made from the fat of animals that have died of disease, even if the fat has been thoroughly sterilized. Very cheap soaps naturally have some suspicion attached regarding the source of the fat stock used. In every branch of manufacturing, unscrupulous persons may be found who plan for their own profit regardless of the rights and the welfare of others. Many laws have been enacted to hold such persons in check. But laws alone will fail to do this. There must be a vigorous sentiment among citizens, and a strong stand made for the " square deal " and the enforcement of law.

Every purchaser of manufactured goods should be on his guard against deceptions and frauds. He should be able to match the greater cunning of modern rogues by a better trained judgment of quality. For example, a certain brand of soap may seem a better bargain because of the greater size of the cakes. A study of its composition may reveal ingredients in the soap that are worthless for cleaning, and in quantity to more than match the larger size of cake. " Fillers " are mixed with the soap during the " crutching " process. Some increase the effectiveness of the soap while others simply give bulk without value. Sodium and potassium carbonates and borax in soaps soften the hardness of the water and prevent a waste of the soap. Resin may be substituted for fat to a certain per cent in making a good soap, but plaster of paris, barium sulfate, sodium silicate, clay, chalk, and talc give bulk without value. Sugar is added to assist in making soap transparent, but the sugar, being so soluble, causes the soap to dissolve rapidly so a cake of it will soon be used up. The amount of water in different soaps varies considerably. It may range from 10 % to as high as 70 %. Consequently it is a matter of economy to know how much water is being purchased at the price of soap.

The cake of soap that is in use in railway stations, hotels, and other public places has been found guilty of harboring disease germs. These germs may not exist a great length of time on the soap, but they may easily survive long enough to be acquired by the next person using the soap. For this reason the cake of soap in public places has been largely replaced by a soap that has considerable water and other solvents in sufficient quantity to make it a liquid. In this form it can be dispensed from sanitary receptacles as wanted, and without the contamination of what remains.

ANALYSIS OF SOAP

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Apparatus: Balance, burette, platinum wire, hot-air oven, and thermometer.

Materials: Denatured alcohol, phenolphthalein, turmeric paper, acetic anhydrid, methyl orange, normal solution of sulfuric acid, and Haine's solution for glucose, sodium carbonate solution.

I. **Per Cent of Water**. Into a clean, dry, previously weighed evaporating dish, shave finely exactly 10 grams of soap from the interior of the cake. Put the dish with shaved soap in a hot oven to dry for several hours at 110° C. Further work on this experiment is best deferred until the next laboratory day. When the soap has been dried to constant weight, compute the per cent of water. Record every weighing, and show all computations.

II. **Per Cent of Alkali**. (*a*) Free Alkali. This alkali is the sodium hydroxid that has been used in excess of the fat or has not been combined with it. A good soap should have no more than 0.5 % of free alkali, and toilet soaps no more than a trace. It dissolves in alcohol. Put 20 to 25 cc. of neutral denatured alcohol on the dried soap from Part I, and heat the mixture over a wire gauze with the flame not touching the gauze, until the liquid is hot, not boiling. **Keep the bottle of alcohol away from flames.** Pour the hot soap solution through a filter, catching the filtrate in a clean, dry beaker, rinse out the dish with a little more warm alcohol, and pour it on the residue in the filter.

The free alkali is in the alcoholic filtrate. Add two drops of phenolphthalein, and stir it. A red color shows the presence of the free alkali. Add, a few drops at a time, a normal solution of sulfuric acid from a burette (see directions for operating, p. 5), and stir the solution. Continue until the red color of the indicator just disappears. Each cubic centimeter of the normal acid used is equivalent to 0.04 gram of sodium hydroxid. Compute the per cent of free alkali found in the 10 grams of soap taken at the beginning. Show all computations (?).

(*b*) Alkali as Carbonates, Borates, or as Silicates. The above salts of sodium are frequently mixed with laundry soaps to soften hard water and to assist in cleansing. They form hydroxyl ions (OH^-) by hydrolysis. Wash the residue from the filter into a beaker with about 30 cc. of distilled water. Heat the mixture just to the boiling point and stir well. If there is much residue not dissolved in water, it should be filtered off, dried, and weighed as insoluble filler in soap. Add a few drops of methyl orange to the water solution of the residue and titrate the alkali with normal sulfuric acid as before. Compute the alkali as sodium hydroxid (?).

III. **Silicates as a Filler in Soap.** Add a little hydrochloric acid to the titrated solution in (*b*), divide it into two equal parts. Evaporate one part to dryness. Heat the dish carefully as it becomes dry. If the residue blackens, heat until all the carbon is burned away. Cool the dish, add warm water and hydrochloric acid. Stir the liquid with a stirring rod (?). A gritty substance in the bottom of the dish shows that a silicate is present. Test this insoluble material for silica by the borax bead (?). The hydrochloric acid converts the sodium silicate into silicic acid, which in turn is decomposed by the heat into water and silica (sand). Write the equation (?).

IV. **Borax as a Filler in Soap.** Dip a strip of turmeric paper halfway into the other half of the acid solution in Part III, and dry it on the outside of a half-filled beaker of hot water (?). A pink color of the paper, changing to olive-green in ammonium hydroxid and pink again in hydrochloric acid, shows that borax is present.

V. **Carbonates as a Filler in Soap.** Shave some of the soap finely and add a few drops of hydrochloric acid to it (?). Effervescence indicates that a carbonate is present.

VI. **Rosin as a Filler in Soap.** This substance is used as a substitute for fat or oil. A small amount of it is no detriment, but in excess it discolors clothing and makes it harsh.

Acidify a water solution of the soap with sulfuric acid. Filter off the precipitate and dissolve it in acetic anhydrid, $(CH_3CO)_2O$. Add a few drops of this solution to a cold mixture of 1 cc. each of concentrated sulfuric acid and water (?). A violet color shows that rosin was used in making the soap.

VII. **Sugar as a Filler in Soap.** Sugar is used only in the attempt to make a soap transparent. Only such soaps need be tested. The sugar is very soluble, so the soap dissolves rapidly and is soon used up.

Boil the filtrate obtained in Part VI for several minutes. Add a solution of sodium carbonate until no further effervescence is noted. Make the test for glucose with Haine's solution.

REPORT OF CHEMICAL ANALYSIS

EXPERIMENT

Partial analysis of..

Obtained from..

Manufactured by...

Exact name of brand..

Tested for...

Description of test...

....

....

....

....

Equations of test...

....

....

Found...

....

Quality of sample..

Remarks..

....

....

I hereby certify that the above report is correct to the best of my knowledge.

(Signed)...

Date...................................

ANALYSIS OF SOILS

More valuable than all the mines of precious metals of the earth is the upper six inches of land surface that is called the soil. All the higher forms of life are dependent on it for food and existence. If it were removed, the earth would be a desert. It is the product of weathering and erosive agents that have been laboring for millions of years to make the earth habitable. The question of foremost concern in every civilized country is how to maintain the fertility of the soil. In the study of this question attention is naturally directed to the condition and the constituents of the soil.

The farmer tests a soil by planting crops in it and noting how well they grow. Hall, the great English soil expert, says: " Mechanical analysis is considered the most important of the various determinations made in the course of soil analysis, and it is believed that for agricultural purposes the size of the soil particles is more significant than their actual composition." They influence the water-holding power of the soil, the amount of food that can be dissolved for plant use, the ease of movement of air and water, the growth of organisms in the soil, and the crop-producing power; in short, the size of the soil particles affects all the soil activities. Of the three main components of a soil, sand, silt, and clay, the last has the smallest particles. During a rain a clay soil " runs together " and becomes sticky and soggy. After drying, it becomes hard if it is not stirred well. It requires much labor to cultivate, yet it needs much cultivation to keep its texture suitable for plant growth. The successful farmer of such a soil strives by careful management to flocculate the soil, *i.e.* unite the small particles into larger particles. Adding lime to a soil helps much in flocculating it. A loam soil, one containing nearly equal proportions of sand and clay, is more desirable than a sandy or a clay soil.

A chemical analysis of soil requires the services of an expert chemist. It shows how great a store of plant food is present, but it may not all be in available form. Soils are rarely uniform in composition. Samples taken from the different parts of a field may show great differences in content. For these reasons the chemical analysis of soils is not often made.

By the decay of vegetable matter in soils, acids are formed which tend to make the soil " sour." Most crops will not make maximum growth in such soils until they have been " sweetened " by neutralizing the acid. The cheapest materials to use are the oxid, hydroxid, and carbonate of calcium. They also improve the texture of the soil, particularly clay, by making it coarser grained (flocculating). Air and water can circulate more readily to the advantage of the growing crop. They also help to form soluble potassium compounds from the insoluble compounds in the soil. Three materials, known as lime, are in common use to " sweeten " soils.

1. Ground limestone, chiefly calcium carbonate. The acid reacts with the carbonate and in this way is used up.

$$\text{Acid} + \text{calcium carbonate} \longrightarrow \text{calcium salt} + \text{water} + \text{carbon dioxid.}$$

This is the safest material to use, for the amateur is less likely to make mistakes by its use. One to two tons per acre of high-grade limestone is applied every five years.

2. Burnt lime, or quicklime, mainly calcium oxid. This oxid is slaked by water and forms a strong base, calcium hydroxid, which neutralizes the acid. Only one ton per acre is used in five or six years, since it acts the most rapidly of all the forms of lime. The inexperienced person may cause the loss of ammonia from the soil by using too much or by not using the proper procedure in applying it. (See the method of preparing ammonia.)

3. Marl. A naturally formed earth composed largely of calcium carbonate derived from accretion by plants and disintegration of shells. It is similar to limestone. Two tons per acre are used in five years.

Lime should be applied before those crops in rotation which are most likely to be benefited by it, such as wheat, alfalfa, and corn. There will still be sufficient in the soil for the following crops that have less need of it. It should be applied after the ground is plowed and then mixed well with the soil by harrowing. Lime should always be applied separately, never mixed with other fertilizers.

ANALYSIS OF SOILS

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

I. **Testing Soil for Acidity.** (*a*) Observation Test. Sorrel and moss thrive in a sour soil, so abundant growth of these weeds indicates an acid soil.

(*b*) Litmus Test. The test is more carefully made as follows: Place strips of red and blue litmus paper well apart on the bottom of a beaker or glass tumbler. Put a disk of filter paper over the litmus papers and cover it with an inch or more of well-moistened soil. Press the soil down firmly on the paper. If, after three hours or more, both papers are red, lime is needed on the soil. If both are blue, no lime is needed.

II. **Vegetation Analysis.** Even the amateur farmer can obtain some idea of the fertility of soil from the crops growing in it. If the vegetation is rank and dark green in color, it indicates a good soil particularly rich in nitrogen. If the vegetation is scanty and light green in color, it lacks nitrogen and probably other plant foods. Leguminous crops, clover, beans, peas, etc., grow fairly well in a soil deficient in nitrogen if the supply of phosphates and potassium compounds is sufficient. These plants are able to take nitrogen from the atmosphere by the help of bacteria.

If the plants have strong, healthy stalks, but are slow in ripening and with a poor yield, then phosphates are not available in sufficient quantity. If the stalk is weak and the kernels of grain poorly filled, or the fruit small and misshapen, a deficiency of potassium compounds is indicated.

III. **Mechanical Analysis.** Put 4 tablespoonfuls of soil in a quart fruit jar, fill three-fourths full of water, shake the mixture, and allow it to stand uncovered until the next day. Then put on the cover and shake the jar thoroughly. Let the mixture settle for one minute and pour the rily water from the sediment into another vessel. Let this stand for one hour and decant the water as before. Evaporate the water from the clay in the last portion, dry the sand (first sediment) and the silt (second sediment), and compute the per cent of sand, silt, and clay in the soil (?). Find the volumes of the three portions in a graduate (or weigh them). Record all figures.

IV. **Pore Space, or Per Cent of Air in Soil.** Put a measured volume of dry soil into a beaker or glass tumbler. Pour in water from a full graduate until it rises just to the surface of the soil. Record as follows:

(*a*) Volume of soil................................cc.
(*b*) Volume of water added........................cc.
Per cent of air space $\frac{(b \times 100)}{a}$.................%

Questions: 1. What was displaced from the soil by the water?
2. How can soil " hold water "?
3. Is the pore space of any given soil constant? Why?
4. Why is an undrained soil not suitable for growing most crops?
5. What is the value of pore space in a soil?

V. **Flocculating Clay Soil.** Put a tablespoonful of clay soil in each of two bottles. Nearly fill each with water and shake thoroughly. Add a little milk of lime to one bottle and note in which case the clay settles most rapidly (?). Explain the effect observed.

VI. **Conditions Affecting Soil Texture.** Mix 200 cc. of clay soil with sufficient water to enable it to keep its shape when molded. Divide it into two portions and make each into a ball. Let one dry in a warm room. Place the other where it will be frozen. After freezing, let it dry. Make up two more balls of clay, one containing half its volume of leaf mold (humus), and the other a tablespoonful of milk of lime. After several days break apart the balls of clay and note how readily each crumbles to pieces (?). Record in order, the most " mellow " first.

REPORT OF CHEMICAL ANALYSIS

EXPERIMENT

Partial analysis of..

Obtained from..

Manufactured by..

Exact name of brand..

Tested for..

Description of test..

....

....

....

....

Equations of test..

....

....

Found..

....

Quality of sample..

Remarks..

....

....

I hereby certify that the above report is correct to the best of my knowledge.

(Signed)..

Date..............................

INSECTICIDES AND FUNGICIDES

Plants, shrubs, and trees are frequently injured by insects and fungus growths. Chemical preparations used to destroy the former are called insecticides, and for the latter fungicides. The insecticides are divided into two classes: poisons used to destroy leaf-eating insects, and contact remedies sprayed on the bodies of juice-sucking insects, plant lice, to kill them. The spray must be diluted sufficiently to avoid killing the plant as well as its enemies.

INSECTICIDES

Insects	Poisons	Insects	Contact Remedies
Potato bugs	Paris green	San José scale	Lime-sulfur
Squash bugs	Paris green	Plant lice	Kerosene emulsion or nicotine
Codling moth	Paris green		
Asparagus beetle	Lead arsenate		

FUNGICIDES

Fungi	Fungicides
Apple bitter rot	Bordeaux mixture
Potato blight	Bordeaux mixture
Grape black rot	Bordeaux mixture
Peach leaf curl	Lime-sulfur
Peach yellows	Lime-sulfur

Spraying is plant insurance. An orchard in Nebraska reports as follows:

Apples marketed from an acre of sprayed trees,	275 bu.	$117.70
Apples marketed from an acre of unsprayed trees,	175 bu.	41.15
Difference		$ 76.55
Cost of spraying		12.00
Net profit per acre by spraying		$ 64.55

For successful work one must know what he is spraying for, what to use, and when and how to use it. For detailed instructions see books on horticulture and government publications.

References: Farmers' Bulletin, No. 243. *Fungicides and Their Use in Preventing Diseases of Fruits.*
Farmers' Bulletin, No. 127. *Important Insecticides.*
Farmers' Bulletin, No. 227. *Lime Sulfur, and Salt Wash.*
Farmers' Bulletin, No. 99. *Insect Enemies of Shade Trees.*
Farmers' Bulletin, No. 320. *Potato Spraying.*
Farmers' Bulletin, No. 267. *Apple Bitter Rot.*
Farmers' Bulletin, No. 283. *Spraying for Apple Diseases and the Codling Moth in the Ozarks.*

INSECTICIDES AND FUNGICIDES

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

Apparatus: Enameled pan.

Materials: Calcium oxid, sulfur flowers, sodium carbonate, arsenious oxid, copper sulfate, lead acetate, sodium arsenate crystals, cheese cloth.

I. **Lime-sulfur**. This insecticide was first used for the San José scale, and originated from a similar lime-sulfur-salt preparation used as a sheep wash. When used to kill scale insects, the mixture must be applied during the winter or early spring before the buds burst, since it " burns " the foliage too severely. A very dilute solution of it is sometimes used to kill fungus growths and to stop the blighting of summer foliage.

Preparation. Slake 36 g. of calcium oxid in 100 cc. of water by heating in an enameled-ware vessel. Mix 80 g. of sulfur flowers with 200 cc. of water and add the paste gradually while stirring. Boil slowly until about one-third of the volume is lost by evaporation of water. After cooling and settling, decant the liquid from the sediment. This liquid is the concentrated lime-sulfur which is sold at retail. It is diluted to suit the requirements of the treatment wanted.

The slaked lime reacts with the sulfur to form calcium sulfid, CaS_4, calcium thiosulfate, CaS_2O_3, and water. Write the correct equation (?). The calcium thiosulfate loses one atom of sulfur to change the calcium sulfid to CaS_5.

II. **Paris Green**. This insecticide is a standard poison for all leaf-eating insects, such as the potato beetle. It " burns " the foliage severely if too much is applied.

Dissolve 9 g. of dry sodium carbonate or 24 g. of the dekahydrate crystals in 80 cc. of water. Add 16 g. of arsenious oxid slowly to this and boil the mixture until it is practically in solution. Dissolve 20 g. of copper sulfate crystals in 80 cc. of water, heating until it is warm. Pour the first solution into the copper sulfate solution. Stir and add about 15 cc. of acetic acid, the last few cubic centimeters slowly. Put in no more acetic acid after the green compound begins to form. If it does not form, keep the mixture warm over a low flame, and after standing for a time stir and add more acetic acid drop by drop. After digesting for an hour, filter and wash the precipitate with a little water. Paris green is considered to be acetoarsenite of copper, $Cu(C_2H_3O_2)_2 \cdot 3\,Cu(AsO_2)_2$. It may be mixed with considerable flour and sifted on the wet foliage, but it is usually sprinkled or sprayed as a suspension in water.

III. **Bordeaux Mixture**. This mixture is especially valuable as a fungicide. Lime is slaked and mixed with copper sulfate, $CuSO_4 \cdot 5\,H_2O$, to form copper hydroxid, calcium sulfate, and water. Write the equation and calculate the amount of dry calcium oxid that is needed to react with 10 g. of copper sulfate. Slake about **5 times** this weight in 200 cc. of warm water and dilute to 400 cc. Strain this slaked lime through cheese cloth. Dissolve 10 g. of copper sulfate crystals in 25 cc. of hot water and dilute it to 400 cc. with cold water. Pour the copper sulfate solution into the lime while stirring. Bordeaux mixture is composed of three insoluble compounds in a fine state of division. When used, it is diluted with water to permit of its being sprayed on vegetation. The copper hydroxid kills the fungi, and the excess lime prevents injury to the foliage.

IV. **Arsenate of Lead**. This substance has largely displaced Paris green as an insecticide because it does not harm foliage and can be used in much greater quantities, making it more effective in destroying insects.

Dissolve 20 g. of lead acetate in 50 cc. of warm water and add several drops of acetic acid. Crude lead acetate contains some lead carbonate which may not dissolve. Dilute the solution with 250 cc. of cold water. Calculate the weight of the sodium arsenate crystals needed to react with the 20 g. of lead acetate, the equation for the reaction being

$$2\,Na_2HAsO_4 \cdot 7\,H_2O + 3\,Pb(CO_2CH_3)_2 \cdot 3\,H_2O \rightleftharpoons Pb_3(AsO_4)_2\downarrow + 4\,Na(CO_2CH_3) + 2\,H(CO_2CH_3) + 16\,H_2O.$$

Dissolve the calculated quantity of the sodium arsenate in 350 cc. of water, and let it run drop by drop into the lead acetate solution while it is being stirred. Let the precipitate settle and wash it several times by decantation. Finally filter it. It is used mixed with considerable water or with Bordeaux mixture. It is said never to harm foliage.

QUALITATIVE ANALYSIS

Qualitative analysis is the process of determining what elements, radicals, or compounds are present in the material to be analyzed. Use is made of differences in the properties of the various compounds of the elements and radicals to be detected. The properties most frequently used are solubility and color. When determining the metals present, a solution of the material is treated with hydrochloric acid to form the insoluble chlorids of silver, lead [1] and mercury (ous), if they happen to be present. After filtering, the precipitate is subjected to further tests to find which of the three metals are present. From the filtrate, insoluble compounds of four other groups of metals are successively removed by precipitation and filtration.

No.	Name of Group and Group Reagent	Metals	Characteristic Properties of the Group
I	Hydrochloric Acid..........	Pb, Ag, Hg^+	Chlorids insoluble in water and dilute acids.
II	Hydrogen Sulfid............	Sn, As, Sb, Hg^{++} (Pb), Bi, Cu, Cd	Sulfids insoluble in dilute acids. First three are soluble in ammonium polysulfid.
III	Ammonium Sulfid..........	Al, Cr, Fe, Co, Ni, Mn, Zn	The first three give insoluble hydroxids on adding ammonium hydroxid. The rest and also iron give sulfids, with ammonium sulfid, insoluble in neutral or alkaline liquids.
IV	Ammonium Carbonate	Ba, Sr, Ca, Mg	Precipitated as carbonates by ammonium carbonate, except the last, which is soluble in ammonium salt solutions.
V	Soluble Group, or Alkali Metals	Li, Na, K, NH_4	No group reagent; detected by special tests.

General Suggestions. 1. Neatness and cleanliness are prime essentials for successful work in analysis. Return bottles to the shelves immediately. Be careful not to mix stoppers, and not to get foreign matter on the stoppers by laying them on the table. Keep all apparatus clean, and rinse with distilled water just before using.

2. Use only distilled water in making solutions, and do not return solutions to the reagent bottles.

3. Use the concentrated acids and bases only when specifically called for.

4. Precipitations must be complete, no further precipitate being formed by adding a few drops of the reagent to a little of the filtrate. Wash precipitates carefully to free them from adhering liquid. Substances dissolved in this liquid may seriously interfere with the tests.

5. In the preliminary tests, the product of every chemical action should be described, named, and the equation written. To determine which of several products is the precipitate, look up their solubilities in the table on page 241.

Fig. 105. Wash bottle.

---

[1] See notes on page 222.

HYDROCHLORIC ACID GROUP

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

I. **Preliminary Tests**. (*a*) Lead. Add hydrochloric acid to 5 cc. of a solution of a lead salt (?). Wash the precipitate twice by decantation, and boil it with a half test-tubeful of water (?). To one-third of the liquid add sulfuric acid (?). Allow the precipitate to settle, decant the liquid, and test the solubility of the precipitate in concentrated hydrochloric acid (?) and in 10 % ammonium acetate solution (?). To another third add potassium dichromate solution (?). (The precipitate is $PbCrO_4$; water enters into this action.) Test the solubility of the precipitate in sodium hydroxid solution (?). Test the last portion with hydrogen sulfid water or gas (?). Wash the precipitate by decantation, divide the precipitate into two parts, and try their solubility in yellow ammonium sulfid and in warm dilute nitric acid (1 vol. of 1.20 acid and 2 of water) (?).

(*b*) Silver. Add hydrochloric acid to 5 cc. of silver nitrate solution (?). Wash the precipitate twice by decantation, and then try the solubility of a very small portion of it in hot water (?). Add ammonium hydroxid to the remainder of the precipitate (?). (New substance is ammonio-silver chlorid, $Ag(NH_3)_2Cl$.) Write the equation for its formation. Acidify with nitric acid (?).

(*c*) Mercury (Mercurous). Add hydrochloric acid to 5 cc. of mercurous nitrate solution (?). Wash the precipitate by decantation, and test the solubility of a very small portion of it in 20 cc. of hot water (?). To the other portion of the precipitate add ammonium hydroxid (?). The product is a mixture of mercury and mercuric ammonium chlorid, $Hg + NH_2HgCl$.

II. **Preparation of the Solution**. Test the unknown for organic matter by heating a pinch (evaporate solutions and use the residue) of the material in a hard-glass test-tube. Smell very cautiously, transferring operations to the hood if bad odors are evolved, as they may be poisonous. If the material becomes black, deposits a tarry matter near the mouth of the tube, or gives off a burnt odor, organic matter is present and must be removed. See W. A. Noyes' *Qualitative Analysis*, p. 99.

Liquids. If the unknown is a liquid free from organic matter, test it with litmus paper. If alkaline, make slightly acid with hydrochloric acid. If neutral or slightly acid, it is ready for analysis. If strongly acid, add 3 or 5 cc. concentrated nitric acid and evaporate to dryness enough to give 1 g. of residue. Dissolve the residue with the smallest possible amount of water, 35 cc. if possible, adding a few drops of nitric acid if necessary, to get a clear solution. Proceed to the analysis.

Solids. If the unknown is a solid, try to dissolve a little of it, using the following solvents in order, first cold and then hot: water, dilute hydrochloric acid, concentrated hydrochloric acid, dilute nitric acid, concentrated nitric acid, aqua regia (3 vols. concentrated hydrochloric acid to 1 vol. concentrated nitric acid). A residue left after boiling with concentrated acid may dissolve in water. If the solid is not completely soluble in any one solvent, dissolve as much as possible in water, and try the residue in the various solvents. If these methods are not sufficient to get the unknown entirely into solution, see Stieglitz' *Qualitative Chemical Analysis*, pp. 94–99, or W. A. Noyes' *Qualitative Analysis*, pp. 100–105.

Having determined upon the solvent, get about 1 g. of the unknown into solution, using as little solvent as possible, and measuring the volume of acid needed. If two solutions have been prepared, it is usually best to analyze them separately. If the solution contains more than 4 cc. of the dilute acid in 40 cc. of the total liquid, or if aqua regia was used, it should be treated as described above for strongly acid liquids. If hydrochloric acid was used in making the solution, the first group must be absent. Proceed at once to the second group.

III. **Method of Analysis**. If the solution of the unknown is neutral or slightly acid, add 4 cc. of dilute hydrochloric acid and dilute to 40 cc. with distilled water (or add enough dilute hydrochloric acid to give a liquid containing $\frac{1}{10}$ its volume of dilute acid). Shake or stir vigorously, and filter, setting the filtrate aside to be analyzed for the metals of the second group. Filter and wash the precipitate (ppt.) twice with cold water, discarding the wash water.

<table>
<tr><td colspan="3">Boil the precipitate two minutes with 20 cc. of water and filter.</td></tr>
<tr><td colspan="2">Residue: $AgCl$, Hg_2Cl_2
Remove all soluble matter by washing several times with boiling water.[2] Then cover the residue with ammonium hydroxid.</td><td rowspan="2">Filtrate: $PbCl_2$
To one-half of the filtrate add potassium dichromate solution. A yellow ppt. soluble in sodium hydroxid solution indicates the presence of lead.
To the other half of the filtrate add sulfuric acid, filter, and pour a few drops of potassium iodid solution on the ppt. A white ppt. turning yellow proves the presence of lead.</td></tr>
<tr><td>Residue: $NH_2HgCl + Hg$
A black residue proves the presence of mercury.</td><td>Filtrate: $Ag(NH_3)_2Cl$
Acidify with nitric acid.[3] A white ppt. of silver chlorid proves the presence of silver.</td></tr>
</table>

IV. **Compounds Formed during the Analysis.**

$PbCl_2$	$PbCrO_4$	$PbSO_4$	PbI_2
Hg_2Cl_2	$NH_2HgCl + Hg$		
$AgCl$	$Ag(NH_3)_2Cl$	$AgCl$	

HYDROGEN SULFID GROUP, TIN DIVISION

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

I. **Preliminary Tests.** (*a*) Arsenic. Add one-half cc. of hydrochloric acid to 5 cc. of a solution of arsenic trichlorid and pass hydrogen sulfid through it (?). Filter, wash, and divide the precipitate into two parts. Test the solubility of one portion of the precipitate in ammonium carbonate solution (?), and then acidify with hydrochloric acid (?). The treatment with the ammonium carbonate gives ammonium thioarsenite $(NH_4)_3AsS_3$, and ammonium arsenite $(NH_4)_3AsO_3$; on adding hydrochloric acid, arsenic trisulfid is reprecipitated, As_2S_3. To the other portion of the precipitate saved above, add ammonium polysulfid $(NH_4)_2S_x$ (?). The product formed is ammonium thioarsenate $(NH_4)_3AsS_4$. Acidify with hydrochloric acid (?), forming arsenic pentasulfid, As_2S_5. Decant, and boil this precipitate with concentrated hydrochloric acid (?).

Repeat the above operations with a solution of sodium arsenate, Na_2HAsO_4 acidified with hydrochloric acid (?). After passing the hydrogen sulfid through for 2 minutes (?), heat to boiling and continue the treatment with hydrogen sulfid (?). The precipitate is arsenic pentasulfid, As_2S_5.

(*b*) Antimony. Add a few drops of hydrochloric acid to 5 cc. of a solution of an antimony salt, and pass hydrogen sulfid through it (?). After filtering and washing the precipitate, test its solubility separately in ammonium carbonate solution (?) and in ammonium polysulfid (?). Acidify the latter with hydrochloric acid (?). Excepting for the ammonium carbonate, the products of these actions are very similar to those of arsenic. Filter, wash, and boil the antimony pentasulfid with concentrated hydrochloric acid until no more hydrogen sulfid is expelled. Antimony trichlorid is formed. Filter if not clear, lay a piece of platinum foil in the solution, and lay a piece of pure tin upon it (?). The deposit is metallic antimony, and is a characteristic test. Clean the platinum with tartaric acid solution containing a few drops of concentrated nitric acid.

(*c*) Tin. Acidify 5 cc. of stannous chlorid solution, $SnCl_2$, with hydrochloric acid, precipitate with hydrogen sulfid (?), filter, wash, and treat the stannous sulfid with ammonium carbonate solution (?) and ammonium polysulfid (?), as was directed for arsenic and antimony. The second forms ammonium thiostannate, $(NH_4)_2SnS_3$. Acidify this with hydrochloric acid (?), filter, wash, and boil with concentrated hydrochloric acid (?). This gives stannic chlorid, $SnCl_4$. Place a couple of bright iron nails in this solution and boil several minutes (?). The result is stannous chlorid. Dilute with an equal volume of water and pour immediately into 5 cc. of mercuric chlorid solution (?). The precipitate is mercurous chlorid, Hg_2Cl_2, and finally mercury. Its formation, being due to the reduction by stannous chlorid, is used as a test for tin.

Repeat the above operations, using stannic chlorid, $SnCl_4$, in place of the stannous chlorid. The formula of the sulfid is SnS_2, and in the ammonium polysulfid it gives the same compound as that above.

II. **Precipitation of the Hydrogen Sulfid Group.** Heat the solution from which the metals of the first group have been precipitated, nearly to boiling, and saturate while hot with hydrogen sulfid. Continue the treatment with hydrogen sulfid for 15 minutes, not allowing the temperature to fall below 70° C. This operation is best carried out in a 150 cc. Erlenmeyer flask fitted with a two-hole stopper, the gas entering through an L-tube extending nearly to the bottom of the flask and escaping through a short L-tube which carries a rubber tube and pinch clamp. At first pass in a rapid stream of hydrogen sulfid, leaving the pinch clamp open so that the air may be swept out. Then the pinch clamp is closed and the solution saturated with the gas under pressure. Open the clamp occasionally for an instant.

Cool and add about one and one-half times as much water as solution. In order to be certain that all of the second group metals have been precipitated, the concentration of the hydrogen ions must not be greater than that of a 0.25 molar solution of hydrochloric acid. Methyl violet, either in solution or in pencil form — copying pencil — is used as the indicator. A drop of the diluted solution should turn a drop of the methyl violet solution or a mark on sized paper to a blue-green color. A blue color indicates that there is not enough acid, while a yellow-green or yellow color indicates too much acid. If in doubt as to the proper color, a quarter molar hydrochloric acid solution may be made up and the color it gives may be used as a standard of comparison. If the solution is too strongly acid, add more water, but if very much is required it will be better to evaporate to dryness and redissolve the residue in 4 cc. dilute hydrochloric acid and 100 cc. of water. When the concentration of the acid has been properly adjusted, saturate the cold liquid with hydrogen sulfid and let it stand 15 minutes.

If no precipitate, or only a fine white powder, forms, Group II is absent. The fine white powder is sulfur, due to the reduction of some compound of one of the metals of Group III. A change in the color of the solution will usually accompany this white precipitate. In either case, proceed directly with Group III. If a colored precipitate is produced, the presence of metals of either or both the tin and copper divisions of Group II is indicated. Filter, and save the filtrate for Group III.

III. **Separation of the Tin and Copper Divisions.** Wash thoroughly with hydrogen sulfid water (prepare this while the generator is not in use otherwise), and transfer a small portion of the precipitate to an evaporating dish. Add a little ammonium polysulfid and warm gently. If the precipitate dissolves completely, the copper division is absent; but if a colored residue is left, the copper division is present. In the latter case, filter, and add hydrochloric acid to excess: a pure white precipitate is sulfur from the ammonium polysulfid, and indicates that the tin division is absent; a colored precipitate indicates that the tin division is present. If only one division is present, the rest of the precipitate is treated according to the method of analysis for that division. If both divisions are present, transfer the rest of the precipitate to the evaporating dish, add 5 to 10 cc. ammonium polysulfid, cover, and warm (**not boil**) for 10 minutes with frequent stirring. Filter, wash thoroughly, and save the precipitate for analysis for the copper division. Acidify the filtrate with hydrochloric acid to reprecipitate the sulfids of the tin division, filter, wash thoroughly, and analyze the precipitate as follows:

IV. **Analysis of the Tin Division.** Transfer the precipitate to an evaporating dish and warm it with about four times its volume of ammonium carbonate solution. Decant the liquid through a filter, and repeat the treatment with the ammonium carbonate solution, pour the liquid through the same filter, combining the filtrates.

<table>
<tr>
<td colspan="2">Residue: Sb_2S_5, SnS_2
Press the residue dry, transfer it to an evaporating dish, and boil with conc. hydrochloric acid until all but sulfur is dissolved. Filter and divide the solution into two portions after diluting with 4 vols. water.</td>
<td rowspan="2">Filtrate: $(NH_4)_3AsS_4$ and $(NH_4)_3AsO_4$
Acidify with hydrochloric acid. A yellow precipitate of arsenic pentasulfid indicates the presence of arsenic.[4]</td>
</tr>
<tr>
<td>Portion I
$SbCl_3$
Place a piece of pure tin on platinum foil in the solution and let it stand 15 minutes. A jet-black stain on the platinum proves the presence of antimony.
Tartaric acid solution with nitric acid will remove the stain.</td>
<td>Portion II
$SnCl_4$
Boil with two or three bright iron nails for several minutes. Filter, letting the filtrate run into a test-tube of mercuric chlorid. A white to gray precipitate of mercurous chlorid, Hg_2Cl_2, and mercury, Hg, proves the presence of tin.</td>
</tr>
</table>

V. Compounds Formed during the Analysis.

As_2S_3	$(NH_4)_3AsS_4$	As_2S_5	$(NH_4)_3AsS_4$ and $(NH_4)_3AsO_4$	As_2S_5	
Sb_2S_3	$(NH_4)_3SbS_4$	Sb_2S_5	$SbCl_3$	Sb	
SnS or SnS_2	$(NH_4)_2SnS_3$	SnS_2	$SnCl_4$	$SnCl_2$	$SnCl_4(Hg_2Cl_2)$

HYDROGEN SULFID GROUP, COPPER DIVISION

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

I. **Preliminary Tests.** (*a*) Mercury (mercuric). Pass hydrogen sulfid through a solution of mercuric nitrate acidified with hydrochloric acid (?). Divide the mercuric sulfid, HgS, into three portions, and test its solubility, separately, in ammonium polysulfid (?), in a warm mixture of one volume of nitric acid (sp. gr. 1.20) and two volumes of water (?), and in warm aqua regia (?). To a solution of mercuric nitrate add stannous chlorid solution, at first slowly (?), and then in excess (?).

(*b*) Copper. To a solution of a copper salt add a few drops of hydrochloric acid, pass in hydrogen sulfid (?), filter, wash, and test the solubility of the copper sulfid, CuS, separately in ammonium polysulfid (?), and in the warm nitric acid mixture described above (?). Add excess of ammonium hydroxid to the latter (?); the new compound is ammoniocopper nitrate, $Cu(NH_3)_4(NO_3)_2$. Divide into two portions, acidify one portion with acetic acid, and add a solution of potassium ferrocyanid, $K_4Fe(CN)_6$ (?). Slightly acidify the other portion with sulfuric acid, boil several minutes with some iron filings, filter, and examine both solution and precipitate (?). Pass hydrogen sulfid through the filtrate (?).

(*c*) Cadmium. Treat a solution of a cadmium salt exactly as directed above for the copper salt, recording all observations, and noting particularly the differences between the properties of the copper and cadmium compounds.

(*d*) Bismuth. Acidify a solution of a bismuth salt with hydrochloric acid, saturate it with hydrogen sulfid (?), filter, wash, and try the solubility of the bismuth trisulfid, Bi_2S_3, in ammonium polysulfid (?), and in the warm nitric acid mixture (?). Add ammonium hydroxid until strongly alkaline (?). Bismuthyl hydroxid, BiOOH, is formed. Filter, wash, and add a few drops of hydrochloric acid to the precipitate on the filter (?), letting the bismuth trichlorid, $BiCl_3$, which is formed drop into a beaker of water (?). Filter, wash, and treat the residue, bismuth oxychlorid, $BiOCl$, on the paper with a fresh sodium stannite solution (?). The residue is metallic bismuth, and its formation in this way is a very delicate test for bismuth.

II. **Analysis of the Copper Division.** As lead chlorid is slightly soluble, even in cold water, it will not be completely precipitated in Group I, and the rest is found in Group II. If the unknown contains only a trace of lead compounds, lead may not be found in Group I at all, and must then be detected in Group II.

Warm a small portion of the residue left after the ammonium polysulfid treatment with 1 to 2 cc. of fresh ammonium polysulfid, filter, and acidify the filtrate with hydrochloric acid. If the precipitate is white or very pale yellow, it is sulfur, and the separation of the tin division was complete. If the precipitate is colored, the main part of the residue should again be treated with ammonium polysulfid until complete separation is accomplished. The filtrate may be rejected. Wash the residue thoroughly and proceed at once to analyze it, as exposure to the air oxidizes the sulfids into sulfates.

Boil the sulfids insoluble in ammonium polysulfid, with 10 to 20 cc. of a mixture of one volume of nitric acid (sp. gr. 1.20) and two volumes water, in a covered beaker, for 2 or 3 minutes. Nitric acid of this concentration dissolves the sulfids of lead, bismuth, copper, and cadmium quite easily, but has practically no effect upon the mercuric sulfid. Filter and wash.

<table>
<tr>
<td rowspan="3">Residue: HgS
Transfer the residue, with filter if necessary, to a test-tube, and boil with 2 to 3 cc. of aqua regia.[5] Dilute to 15 cc., filter, and add to the solution clear stannous chlorid solution slowly and then to excess. A white to grayish black precipitate proves mercury present.</td>
<td colspan="3">Filtrate: $Pb(NO_3)_2$, $Bi(NO_3)_3$, $Cu(NO_3)_2$, $Cd(NO_3)_2$
Add 3 to 4 cc. of conc. sulfuric acid, and evaporate in a porcelain dish until dense white fumes of sulfuric acid just begin to come off. Cool, pour slowly into 15 cc. water, rinse the dish with this solution, shake, cool, and let it stand 5 minutes. If there is a white precipitate, filter and wash.</td>
</tr>
<tr>
<td rowspan="2">Precipitate: $PbSO_4$
Pour a few drops of potassium iodid solution over the precipitate. If the white precipitate becomes yellow, lead is proved present.</td>
<td colspan="2">Filtrate: $Bi_2(SO_4)_3$, $CuSO_4$, $CdSO_4$
Add ammonium hydroxid until strongly alkaline. A deep-blue color proves the presence of copper. A white precipitate indicates bismuth. Filter and wash.</td>
</tr>
<tr>
<td>Precipitate: $Bi(OH)_3$
Add 2 or 3 drops of hydrochloric acid to the precipitate on the filter and let the solution fall into a beaker of water. A white precipitate proves the presence of bismuth.</td>
<td>Filtrate: $Cu(NH_3)_4SO_4$, $Cd(NH_3)_4SO_4$
If copper is absent, acidify the filtrate slightly with sulfuric acid and treat with hydrogen sulfid. A yellow precipitate proves cadmium. If copper is present, acidify with sulfuric acid, add iron filings, and boil. Filter,[6] and treat with hydrogen sulfid. A yellow precipitate proves cadmium present.</td>
</tr>
</table>

III. **Compounds Formed during the Analysis.**

HgS	HgS	HgS	$HgCl_2$	Hg_2Cl_2	Hg	
PbS	PbS	$Pb(NO_3)_2$	$PbSO_4$	PbI_2		
Bi_2S_3	Bi_2S_3	$Bi(NO_3)_2$	$Bi_2(SO_4)_3$	$BiOOH$	$BiCl_3$	$BiOCl$
CuS	CuS	$Cu(NO_3)_2$	$CuSO_4$	$Cu(NH_3)_4SO_4$	$CuSO_4$	Cu
CdS	CdS	$Cd(NO_3)_2$	$CdSO_4$	$Cd(NH_3)_4SO_4$	$CdSO_4$	CdS

AMMONIUM SULFID GROUP, ALUMINUM DIVISION

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

I. **Preliminary Tests.** (*a*) Aluminum. To an aluminum salt solution add ammonium chlorid solution and ammonium hydroxid (?). Filter, wash, dissolve in a little dilute hydrochloric acid (?), dilute to 10 cc. and add sodium hydroxid solution to excess (?). (See experiment on amphoters.) Acidify with dilute nitric acid (?), and then make alkaline with ammonium hydroxid (?).

(*b*) Chromium. Acidify a solution of potassium bichromate with hydrochloric acid and saturate it with hydrogen sulfid (?). The color of the solution after this treatment is characteristic of chromium salts, as chromium chlorid, $CrCl_3$. Identify and account for the precipitate (?). Boil, filter, add ammonium chlorid solution and ammonium hydroxid (?). Filter, wash, dissolve the precipitate with a little hydrochloric acid (?), dilute to 10 cc., make alkalin with sodium hydroxid solution (?), cool, and add 1 to 2 g. of solid sodium peroxid, Na_2O_2, in small portions, while stirring (?). The yellow color is characteristic of chromates, here sodium chromate, Na_2CrO_4.

(*c*) Iron. Acidify a solution of a ferric salt with hydrochloric acid, saturate with hydrogen sulfid (?), boil to expel the excess of the gas, filter, and add bromin water until the odor of bromin persists after stirring (?). Boil to remove the excess of bromin from the solution. Add ammonium chlorid solution and ammonium hydroxid to alkaline reaction (?). Filter, wash, dissolve the precipitate with a little hydrochloric acid (?), dilute to 10 cc., and make alkaline with sodium hydroxid solution (?). Filter, redissolve in a little hydrochloric acid, dilute, divide into three portions, add potassium sulfocyanate, KCNS, solution to one (?), and potassium ferrocyanid, $K_4Fe(CN)_6$, solution to the second (?), and ammonium hydroxid to the third (?).

II. **Preparation of the Solution.** If organic matter has not been removed, it must be tested for at this time and removed, as it interferes with the separation of the members of this group. Phosphates must also be tested for and removed to prevent the precipitation of the metals of Group IV along with those of Group III. Boil[7] the filtrate from the precipitation of the metals of Group II; warm a little of it with three or four times its volume of ammonium molybdate solution to a temperature of 60 to 70° C., and let it stand a short time. A crystalline yellow precipitate, ammonium phosphomolybdate, indicates the presence of phosphates. Discard.

If phosphates are present, test one-tenth of the solution for iron by means of potassium sulfocyanate and potassium ferrocyanid solutions. To the remainder of the solution add ammonium hydroxid until the precipitate just fails to redissolve on shaking. The addition of too much ammonium hydroxid may be rectified by adding acetic acid until distinctly acid. Add 5 cc. of 50% ammonium acetate solution, and then add ferric chlorid solution, drop by drop, until the mixture is brownish red. This color is not produced until all of the phosphate has been precipitated as ferric phosphate. Dilute the mixture to 100 cc., boil for 5 minutes, and after it has stood a short time, filter while hot. Reserve the precipitate and test it later for aluminum and chromium if these are not found in the regular course of analysis, as this precipitate sometimes carries down most or all of these metals.

To get rid of the excess of iron added to remove the phosphate, add 4 to 5 cc. of the ammonium acetate solution to the filtrate, boil, and if a precipitate forms, filter it off and reject it. The solution is now ready to be tested by the following method, omitting, however, the tests for iron.

If phosphates were not found in the solution, treat it according to the following outline.

III. **Analysis of the Ammonium Sulfid Group, Aluminum Division.** Boil the solution to remove hydrogen sulfid, add a little potassium chlorate or bromin water to the hot liquid, until the odor of chlorin or bromin persists, and then boil to remove the excess chlorin or bromin. Cool, add 10 cc. ammonium chlorid solution, and then ammonium hydroxid until decidedly alkaline. Warm, filter at once, and wash thoroughly. The filtrate is to be analyzed for members of the zinc division of Group III and those of Group IV.

The precipitate may contain the hydroxids of aluminum, iron, and chromium and possibly zinc chromite. Transfer it to a procelain dish, cover with water, add 3 cc. sodium hydroxid solution, and then add 1 to 3 g. of solid sodium peroxid[8] in small portions with constant stirring.

Residue: $Fe(OH)_3$	Filtrate: $NaAlO_2$, Na_2CrO_4	
Dissolve the residue in hydrochloric acid, and divide the solution into three portions. A blood-red color with potassium sulfocyanate, a brown precipitate with ammonium hydroxid, and a blue precipitate with potassium ferrocyanid, prove the presence of iron.	Divide the filtrate into two portions, testing one for aluminum, the other for chromium.	
	PORTION I $NaAlO_2$	PORTION II Na_2CrO_4
	Acidify with nitric acid, boil, add ammonium hydroxid, boil; a flocculent white precipitate proves the presence of aluminum.	Acidify slightly with acetic acid. A yellow solution proves chromium. If the solution is at all yellow, add 10 cc. of 10 % barium chlorid solution and allow the mixture to stand 5 minutes. Filter, warm the filtrate to about 50° C., saturate with hydrogen sulfid, and let it stand 10 minutes. A white flocculent precipitate proves zinc.[9]

IV. Compounds Formed during the Analysis.

$Al(OH)_3$ $NaAlO_2$ $Al(NO_3)_3$
$Cr(OH)_3$ Na_2CrO_4 $BaCrO_4$
$Fe(OH)_3$ $Fe(OH)_3$ $FeCl_3$ $Fe(CNS)_3$ $Fe(OH)_3$ $Fe_4(Fe(CN)_6)_3$

AMMONIUM SULFID GROUP, ZINC DIVISION

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

I. **Preliminary Tests.** (*a*) Nickel. To a cobalt solution add ammonium chlorid solution and ammonium hydroxid until decidedly alkaline (?). Pass hydrogen sulfid gas into the liquid (?), and filter; if the filtrate is dark brown in color, boil it to remove ammonia and pour it again through the same paper. Test the residue in the borax bead (?). Transfer the paper and precipitate to a porcelain dish, and boil with 5 to 15 cc. dilute hydrochloric acid (?). Add a few drops of concentrated nitric acid and continue the boiling (?). When the black precipitate is completely dissolved, filter, evaporate to small bulk. While hot, add an equal volume of bromin water, and then of sodium hydroxid solution. Boil vigorously, and add a little more bromin water at intervals until a drop of the liquid turns a piece of iodo-starch paper blue. Describe the precipitate (?); it is nickelic hydroxid, $Ni(OH)_3$. Boil it with ammonium chlorid and ammonium hydroxid solutions (?). The solution has a variable composition. Treat the solution with hydrogen sulfid (?). Filter, wash, and test the precipitate with the borax bead.

(*b*) Cobalt. Repeat the above work, using a solution of a cobalt salt instead of the nickel salt. Record all observations and note especially the differences between the compounds of the two metals.

(*c*) Manganese. To a solution of potassium permanganate add hydrochloric acid and saturate with hydrogen sulfid (?). The color of the solution after this treatment is characteristic of manganous salts, as manganous chlorid, $MnCl_2$. Identify and account for the precipitate (?). Boil, filter, and to a small portion, add ammonium hydroxid solution (?); discard. To the rest of the solution add ammonium chlorid solution and ammonium hydroxid (?), and then saturate with hydrogen sulfid (?). Filter, wash, and treat the precipitate on the paper with very dilute hydrochloric acid (1 vol. dilute acid to 5 vols. water), pouring the same acid over the precipitate several times if necessary (?). Add sodium hydroxid solution in excess (?), filter, and wash. Transfer a little of the precipitate to a dish, add 1 to 2 g. of lead dioxid, PbO_2, and about 10 cc. nitric acid. Heat to boiling for 2 minutes, pour into a test-tube and allow to settle (?). The violet-red color of the permanganate ion of the permanganic acid, $HMnO_4$, produced is a very delicate test for manganese.

(*d*) Zinc. To a solution of a zinc salt add ammonium chlorid solution and ammonium hydroxid, slowly first, then to excess (?). The solution contains an ammoniozinc salt, as $Zn(NH_3)_4(NO_3)_2$. Treat the liquid with hydrogen sulfid (?), filter, wash, and treat the precipitate on the filter with very dilute hydrochloric acid as was done with the manganese compound above (?). Boil to remove hydrogen sulfid, add sodium hydroxid slowly at first and then to excess (?), acidify with acetic acid (?), and then saturate with hydrogen sulfid (?).

II. **Analysis of the Ammonium Sulfid Group, Zinc Division.** Saturate the filtrate obtained from the precipitation of the aluminum division metals with hydrogen sulfid, boil for a few minutes, filter, wash thoroughly; save the filtrate for analysis for the metals of Group IV. Transfer the precipitate to a beaker, treat it with 75 to 150 cc. of cold, very dilute hydrochloric acid (1 vol. dilute acid to 4 vols. water), and filter at once.

<table>
<tr>
<td colspan="2">Residue: CoS, NiS
After washing, test the residue in the borax bead. A blue bead proves cobalt, a brown one proves nickel. Even a small amount of cobalt will conceal the color of considerable nickel. To detect both, dissolve the residue in a little dilute aqua regia.[10] Filter, evaporate to small bulk, add an equal volume of bromin water, and then of sodium hydroxid; boil, and add more bromin water at intervals until a drop of the liquid turns a piece of iodo-starch paper blue. Wash thoroughly with boiling water by decantation, filter, and boil the residue with ammonium chlorid and ammonium hydroxid solutions, and filter.</td>
<td colspan="2">Filtrate: $MnCl_2$, $ZnCl_2$
Add sodium hydroxid solution in excess, boil, filter, and wash.</td>
</tr>
<tr>
<td>Residue: $Co(OH)_3$
Test the residue with the borax bead. A blue bead proves the presence of cobalt.</td>
<td>Filtrate: Nickel solution
Saturate with hydrogen sulfid. A black precipitate indicates the presence of nickel. Test with borax bead. A brown bead proves the presence of nickel.</td>
<td>Precipitate: $Mn(OH)_2$
If the precipitate is brown, the presence of manganese is indicated. Confirm by boiling part of the precipitate with lead dioxid and 10 cc. nitric acid. Let settle. A violet-red liquid proves the presence of manganese.</td>
<td>Filtrate: Na_2ZnO_2
Acidify with acetic acid and saturate with hydrogen sulfid. A white precipitate proves the presence of zinc.[9]</td>
</tr>
</table>

III. Compounds Formed during the Analysis.

MnS	$MnCl_2$	$Mn(OH)_2$	$HMnO_4$	
ZnS	$ZnCl_2$	Na_2ZnO_2	$Zn(C_2H_3O_2)_2$	ZnS
CoS	CoS	$CoCl_2$	$Co(OH)_3$	
NiS	NiS	$NiCl_2$	$Ni(OH)_3$	NiS

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

I. **Preliminary Tests.** (*a*) Barium. Add ammonium chlorid solution, and then ammonium carbonate solution to a barium salt solution (?). Filter, wash, and dissolve the precipitate by pouring a little dilute acetic acid repeatedly over it while on the filter (?). Add potassium chromate solution to the filtrate (?), filter, wash, dissolve in a little hydrochloric acid (?), make a flame test with platinum wire (?), and add a few drops of sulfuric acid (?).

(*b*) Strontium. To a strontium salt solution, add ammonium chlorid solution and then ammonium carbonate solution (?). Filter, wash, dissolve from the filter in the smallest possible volume of dilute acetic acid (?), dip a clean platinum wire into the solution, and hold it in the outer edge of the blue flame (?). Divide the solution into two portions, add potassium chromate solution to one (?), and ammonium sulfate solution to the other (?).

(*c*) Calcium. To a calcium salt solution add ammonium chlorid solution and then ammonium carbonate solution (?). Filter, wash, dissolve from the filter in the smallest possible volume of dilute acetic acid (?), and make a flame test (?). Divide the solution into three portions. To one add potassium chromate solution (?). To another add ammonium sulfate solution (?). To the third portion add ammonium oxalate, $(NH_4)_2C_2O_4$, solution (?). Test the solubility of the precipitate in hydrochloric acid (?).

(*d*) Magnesium. Add ammonium carbonate solution to a solution of a magnesium salt (?). Reject. Repeat the test, but add ammonium chlorid solution before adding the carbonate solution (?). If a precipitate forms, add more ammonium chlorid. Add ammonium hydroxid and then a little disodium hydrogen phosphate, Na_2HPO_4, solution (?). The crystalline precipitate is magnesium ammonium phosphate, $MgNH_4PO_4$. If it does not appear, rub the inside of the test-tube with a glass rod.

II. **Analysis of the Ammonium Carbonate Group.** Acidify the filtrate from the ammonium sulfid group and boil until the odor of hydrogen sulfid disappears and the solution is colorless. Evaporate to a volume of 25 to 30 cc., and filter off any sulfur that is precipitated. To the filtrate add ammonium hydroxid solution to alkaline reaction, boil, add ammonium carbonate solution as long as a precipitate forms, and then set the mixture aside for 10 minutes. Filter, and wash the precipitate thoroughly.

<table>
<tr>
<td colspan="3">Precipitate: $BaCO_3$, $SrCO_3$, $CaCO_3$
Dissolve the precipitate in a little warm dilute acetic acid. To a little of this solution add potassium chromate solution; a yellow precipitate indicates the presence of barium. If barium is present, add an excess of potassium chromate solution to the rest of the solution, filter, and wash. If barium is absent, use the rest of the solution as the filtrate below.</td>
<td rowspan="3">Filtrate: $MgCl_2$
To half the filtrate add ammonium hydroxid, cool, and add a little disodium hydrogen phosphate solution. A crystalline precipitate which may not appear until after rubbing the inside of the tube with a glass rod proves the presence of magnesium.
If the precipitate is flocculent, it is due to aluminum. Acidify the rest of the filtrate with sulfuric acid; filter. Add ammonium hydroxid to the filtrate until alkaline, and then add ammonium oxalate solution. Warm and filter. Test the filtrate for magnesium with disodium phosphate solution as directed above.</td>
</tr>
<tr>
<td rowspan="2">Precipitate: $BaCrO_4$
Dissolve the precipitate in hydrochloric acid and add a few drops of sulfuric acid. A white precipitate proves the presence of barium.</td>
<td colspan="2">Filtrate: $Sr(C_2H_3O_2)_2$, $Ca(C_2H_3O_2)_2$
Make alkaline with ammonium hydroxid, add excess of ammonium carbonate solution, warm, filter, wash, rejecting the filtrate. Dissolve the precipitate in warm dilute acetic acid and boil with ammonium sulfate solution. Filter and wash.</td>
</tr>
<tr>
<td>Precipitate: $SrSO_4$
Moisten the precipitate with hydrochloric acid and test in the flame. A lasting crimson color proves the presence of strontium.</td>
<td>Filtrate: $Ca(C_2H_3O_4)_2$
Add a few drops of acetic acid and 1 cc. of ammonium oxalate solution. A fine white precipitate proves the presence of calcium.</td>
</tr>
</table>

III. Compounds Formed during the Analysis.

$BaCO_3$	$Ba(C_2H_3O_2)_2$	$BaCrO_4$	$BaCl_2$	$BaSO_4$	
$SrCO_3$	$Sr(C_2H_3O_2)_2$	$Sr(C_2H_3O_2)_2$	$SrCO_3$	$Sr(C_2H_3O_2)_2$	$SrSO_4$
$CaCO_3$	$Ca(C_2H_3O_2)_2$	$Ca(C_2H_3O_2)_2$	$CaCO_3$	$Ca(C_2H_3O_2)_2$	CaC_2O_4
$MgCl_2$	$MgNH_4PO_4$				

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

I. **Preliminary Tests.** (*a*) Ammonium. Place a small portion of an ammonium salt in a small beaker, add sodium hydroxid solution, cover the beaker with a watch crystal or glass plate on the under side of which a piece of moistened red litmus paper has been placed, and warm but do not boil (?).

(*b*) Sodium. Test a little of some sodium compound on a clean platinum wire in the blue flame (?). Note the length of time the color lasts (?). Repeat the test, viewing the flame through a double thickness of cobalt blue glass (?).

(*c*) Potassium. Test a potassium salt on a clean platinum wire in the blue flame (?). Note the duration of the color (?). Repeat the test, viewing the flame through a double thickness of cobalt blue glass (?).

II. **Analysis of the Alkali Group.** The tests for the members of the alkali group are best made with the original solution of the unknown, or that for ammonium can be made with the solid unknown. Many analysts, however, use only a small portion of the filtrate from which barium, strontium, and calcium have been removed for the test for magnesium, and use the rest of the filtrate for tests for sodium and potassium. If magnesium was found present, add sufficient ammonium phosphate solution, $(NH_4)_2HPO_4$, to insure its complete removal,[11] filter, evaporate the filtrate to dryness, and then ignite the residue at a dull-red heat to expel the ammonium salts. Heat the sides as well as the bottom of the dish. Use a little of the residue to make the flame tests for sodium and potassium described above.

If it is desired to form an estimate of the amount of sodium and potassium present, the rest of the residue should be dissolved in as little hot water[12] as possible (1 to 2 cc.), half of the solution acidified with a drop or two of acetic acid, and an equal volume of sodium cobaltinitrite solution, $Na_3Co(NO_2)_6$, added to it. Allow the mixture to stand 30 minutes unless a precipitate forms sooner. A yellow precipitate, $K_2NaCo(NO_2)_6$, proves the presence of potassium.

Make the rest of the solution slightly alkaline with potassium hydroxid solution and add twice its bulk of potassium pyroantimonate solution, $K_2H_2Sb_2O_7$. Shake the mixture, scratch the inside of the tube with a clean glass rod, and let it stand for 30 minutes. A crystalline, white precipitate, $Na_2H_2Sb_2O_7$, proves the presence of sodium.

There is considerable danger that sodium compounds will be introduced into the liquid during the course of analysis, either through carelessness or through impure reagents, so there is some advantage in making the tests with the original solution. If members of the ammonium carbonate group were found present, they must be removed, because their flames will interfere with the tests for sodium and potassium. To remove them, make the solution alkaline with ammonium hydroxid, and then add sufficient ammonium carbonate solution to assure complete precipitation. Filter, and use the filtrate to test for sodium and potassium according to the methods given above.

Ammonium is tested for by placing some of the unknown substance, or some of the original solution before any reagents have been added to it, in a small beaker, adding sodium hydroxid solution, covering with a watch glass having moist red litmus paper stuck to its under side, and warming carefully. If the paper turns blue, ammonium salts are present. Usually the odor of ammonia gas will be perceptible.

EXPERIMENT

Purpose: State it on a separate sheet used for the record.

I. **Special Tests**. (*a*) Carbonates. Review the test for carbonates on page 82.

(*b*) Chlorids. Review the test for chlorids on page 58.

(*c*) Nitrates. Review the test for nitrates on page 118.

(*d*) Sulfates. Review the test for sulfates on page 58.

(*e*) Nitrites. Add dilute sulfuric acid to a small amount of a nitrite, noting the color of the gases by looking down through the tube (?). Treat potassium permanganate solution with an acidified solution of a nitrite (?). Treat iodo-starch paper with an acidified nitrite solution (?). (See page 96.)

(*f*) Sulfids. Moisten a sulfid and place it on a clean silver coin (?). If there is no effect, fuse a bit of the sulfid with sodium carbonate in a crucible, moisten, and place on the coin (?).

(*g*) Sulfites. To a small amount of a sulfite add hydrochloric acid and note the odor of the gas given off (?), warming if necessary. The tests described on page 104 are more delicate.

(*h*) Bromids. Heat a small amount of a bromid with concentrated sulfuric acid (?), and hold moist starch in the fumes. The sulfur dioxid comes from the sulfuric acid.

(*i*) Iodids. Heat a small amount of an iodid with concentrated sulfuric acid (?). Hold moist starch in the fumes.

(*j*) Phosphates. Review the test for phosphates on page 124.

(*k*) Acetates. Add concentrated sulfuric acid to a solution of an acetate, and warm, noting the odor (?). Repeat, adding ethyl alcohol before adding the sulfuric acid (?).

(*l*) Tartrates. Heat a small amount of a tartrate on a crucible cover, or in a spoon, noting the peculiar odor (?) and the change in color (?).

(*m*) Permanganates. Dissolve a permanganate in water; the color (?) is characteristic of the permanganate ion. Acidify the solution with sulfuric acid and pass in hydrogen sulfid (?).

(*n*) Fluorids. See the test for fluorids on page 127.

(*o*) Borates. See the test for borates on page 128.

(*p*) Thiosulfates. Add dilute hydrochloric acid to a solution of a thiosulfate, and warm; note the odor (?) and the precipitate of sulfur (?). Add a little silver nitrate solution to a thiosulfate solution and describe the precipitate of silver thiosulfate, $Ag_2S_2O_3$, and note the changes in color due to the separation of silver sulfid, Ag_2S (?).

(*q*) Oxalates. Acidify a solution of an oxalate with hydrochloric acid, add an equal volume of saturated sodium acetate solution, and then some calcium sulfate solution. Describe the precipitate of calcium oxalate, CaC_2O_4. A fluorid would also give a precipitate under these conditions. Mix a little solid oxalate with a little powdered manganese dioxid, add 2 or 3 cc. dilute sulfuric acid in a tube, and heat gently. Test the gas evolved with a drop of lime-water held in a wire loop (?).

II. **Analysis of the Acid Radicals**. Whether or not certain acid radicals can be present in an unknown may be determined by comparing the list of metals present in it with the table of solubilities on page 241. In this connection, one must also take into account the solubility of the unknown itself. A material which is soluble in water and gives a neutral or slightly acid solution containing Zn^{++} and Ba^{++} ions need not be tested for sulfids, sulfites, carbonates, oxalates, tartrates, phosphates, arsenites, arsenates, chromates, thiosulfates, sulfates, and fluorids.

(*a*) **Preliminary Examination**. Moisten a small amount of the finely powdered solid (evaporate solutions) with water and then add about 1 cc. concentrated sulfuric acid, warming carefully if nothing happens. This should be done in a hood, and the vapors given off should be smelled cautiously, for poisonous gases may be given off. The results should be regarded as indications only, and should be corroborated by the special tests given above, especially when they are negative.

Chlorids give hydrogen chlorid gas — characteristic odor, and white precipitate in drop of silver nitrate.

Carbonates give carbon dioxid, turning a drop of lime-water milky.

Nitrates give nitric acid and some brown fumes — odor.

Nitrites give dense brown fumes.

Sulfids give hydrogen sulfid — odor.

Sulfites give sulfur dioxid — odor.

Bromids give bromin fumes and sulfur dioxid — color and odor.

Iodids give iodin vapor, hydrogen sulfid, and sulfur dioxid — color and odor.

Acetates give acetic acid vapor — odor of vinegar.

Tartrates char and give sulfur dioxid.

Fluorids give hydrogen fluorid which etches the test-tube, also makes drop of water held in loop of platinum wire turbid. Warm if necessary.

Permanganates give violet fumes, explosive.

(*b*) **Groups of the Acids.** The acids may be grouped as follows :

Group I. Acids whose silver salts are insoluble in water or dilute nitric acid : hydrochloric, hydrobromic, hydriodic, (hydrosulfuric, H_2S).

Group II. Acids whose silver salts are soluble in dilute nitric acid, but insoluble or but slightly soluble in water.

(*a*) Acids whose silver salts are colorless (white) : oxalic, boric, tartaric, (sulfurous), (carbonic), (nitrous), and acetic, whose silver salts are soluble in hot water.

(*b*) Acids whose silver salts are colored : phosphoric, arsenic, arsenious, (thiosulfuric), chromic.

Group III. Acids whose silver salts are soluble in water and dilute nitric acid : nitric, chloric, permanganic, sulfuric, hydrofluoric, acetic, (nitrous). The last two give silver salts which are difficultly soluble in water.

Group IV. Organic acids which carbonize on heating : tartaric, acetic, citric, benzoic, malic, and hundreds of others.

(*c*) **Preparation of the Solution.**[13] Boil 1 gram of the finely powdered substance in a porcelain dish with about 10 cc. concentrated sodium carbonate solution for 10 minutes, adding water as it evaporates, and filter.

(*d*) **General Tests for the Groups.** Neutralize half of the filtrate with nitric acid, filter if a precipitate forms, acidify strongly with nitric acid, boil to expel carbon dioxid, and add 1 to 2 cc. silver nitrate solution. A precipitate indicates the presence of acids of Group I. Add a little more silver nitrate solution to insure complete precipitation, filtering if necessary. Hold the tube in a slanting position and carefully pour a little dilute ammonium hydroxid down the sides so as to form a layer on top. A white precipitate at the junction of the two liquids indicates the presence of acids of Group II (*a*), while a colored precipitate indicates the presence of acids of Group II (*b*). Acidify another small portion of the sodium carbonate solution with acetic acid for the oxalate test, another portion with sulfuric acid for the nitrate test, and the rest of it with hydrochloric acid for the sulfate test.

The presence or absence of those acids of Groups I, II, and III which are inclosed in parentheses, cannot be determined by the silver nitrate test above, because the boiling with the dilute nitric acid will expel them. So the original substance must be tested for sulfids, carbonates, sulfites, and thiosulfates.

IMPROVED TEST FOR ARSENIC[a]

Arsenic is very poisonous. It frequently gets into foods by contact with other metals containing it, in quantities sufficient to cause dangerous poisoning.

Apparatus. The generator bottle (Fig. 106) is a 2-oz. wide-mouth bottle with a No. 4 rubber stopper, in which is fitted the lower absorption tube. The absorption tubes are glass tubes about 12 mm. inside diameter and about 5 cm. long, with a piece of 3 mm. tubing about 3 cm. long fused to the end of each. Each of these tubes should be provided with a No. 00 rubber stopper. Into the upper stopper insert the reaction tube, a glass tube 3 mm. inside diameter and 11 or 12 cm. in length. The upper end of this tube should be filed or ground down to take off the sharp edge, instead of being rounded in the flame. The lower absorption tube should be filled with absorbent cotton soaked in 20 % lead acetate solution, squeezed out moderately by hand, and packed comfortably (not too tightly) into the tube. The upper tube is filled in the same way except that 5 % lead acetate solution is used. In packing the lower tube, the lower 1.5 or 2 cm. of the wide part should be left empty, so that the cotton will not get soaked with the acid that gathers in the inlet of the tube.

FIG. 106. Apparatus for arsenic test.

Preparation of Standards. Into each generator bottle put about 40 cc. of sulfuric acid solution containing about 3½ cc. of concentrated, arsenic-free sulfuric acid. Add about ½ cc. of stannous chlorid solution[b] and about ½ cc. of ferrous ammonium sulfate solution.[c]

In the reaction tube of each generator place a strip of mercuric bromid paper ⅛ of an inch wide. The mercuric bromid paper is prepared by soaking 3 × 4-inch pieces of No. 00 Munktell's filter paper in 1½ % alcoholic mercuric bromid solution, removing, drying in air, and cutting into strips (use a photo trimmer). The ends of these strips should be cut off before using.

To the different generator bottles add such amounts of a standard arsenic solution[d] that they will contain respectively ½, 1, 1½, 2, 3, 4, 6, 8, and 10 micromilligrams (millionths of a gram) of arsenious oxid, As_2O_3. Keep the bottles in a water bath kept at 90° C. for about 10 minutes. Remove them from the bath and test the contents of each bottle for ferric iron by mixing one drop with one drop of $N/_{10}$ potassium sulfocyanid solution (1 g. to 100 cc. water). A red color indicates the presence of ferric iron. A porcelain spot plate is very convenient in making this test. If no ferric iron is present, cool in a water bath to about 25° C. Add two sticks of arsenic-free zinc about 1 inch long and ¼ inch diameter. Place the stopper with the absorption tubes into the bottles at once. See that the water in the bath does not get much above 25° C. and allow the reaction to proceed about 45 minutes. At the end of this time, remove the test papers and put in fresh papers to make sure that the evolution is complete (the ends originally cut off will do for this).

Arrange the test strips in proper order on cross-section paper, using one of the horizontal rulings as a common base line. Leave the proper intervals between the test strips, so that each centimeter division will represent ½ micromilligram of arsenious oxid. Mark on the cross-section paper

---

[a] David Klein, State Analyst, Division of Foods and Dairies, Department of Agriculture, State of Illinois.

[b] Dissolve 40 g. of stannous chlorid crystals in concentrated hydrochloric acid and water, and dilute to 100 cc., adding additional acid if a precipitate forms due to the hydrolysis of the stannous chlorid.

[c] Dissolve 30 g. of ferrous ammonium sulfate crystals in 100 cc. water, and add a few drops of concentrated sulfuric acid to prevent hydrolysis.

[d] Any convenient concentration will serve this purpose. One suggested is made as follows: weigh 1.000 g. of chemically pure arsenious oxid, As_2O_3, into a beaker, dissolve it in 27 cc. of sodium hydroxid solution, neutralize with sulfuric acid, add 10 cc. of concentrated sulfuric acid, dilute to 1 L. and keep for preparing any more dilute solution. One cubic centimeter of this solution contains 1 mg. of arsenious oxid. Dilute 10 cc. of this solution to 2 L.; 1 cc. of the diluted solution contains 5 micromilligrams of the arsenious oxid.

the upper limits of the prominent stain zone. Through these points draw a straight line. This line can be used for comparisons in place of the standard papers.

Procedure. In the tests for arsenic in inorganic substances, including baking powders and wall papers, a weighed amount of the sample is digested in moderately concentrated hydrochloric acid for several hours. The acid solution is put into the generator with arsenic-free sulfuric acid, and stannous chlorid and ferrous ammonium sulfate solutions, and warmed in a water bath to 90° C. as above stated. Then the same procedure is followed as in the standardization of the apparatus.

Food products in general should be first treated with C. P. concentrated nitric and sulfuric acids (in the hood) to oxidize the arsenic compounds to arsenic acid and char the sample. In this form the arsenic can be extracted by repeated treatments of boiling water. This solution is cooled, filtered, and put into the generator as before for the test.

In the case of meats and vegetables, heat about 100 g. of the sample in a porcelain dish with 23 cc. of concentrated nitric acid to 150° C. and stir with a long glass rod. Work in a hood, and protect the hands from the fumes. When the sample becomes yellow or orange, cease heating and add 3 cc. of concentrated sulfuric acid and stir the mixture in a hood while nitrous fumes are given off. Heat again to 180° C. and add, drop by drop, 8 cc. of concentrated nitric acid while stirring. Finally heat until dense white fumes of sulfuric acid begin to appear, and the residue is dry and black. The arsenic content of this residue may be extracted by boiling water and tested as given above.

In case there is reason to believe that the article being tested contains considerable arsenic, the extract should be diluted to some definite volume in a volumetric flask, and then such aliquot portion used as will contain from 6 to 10 micromilligrams of arsenious oxid. A trial test may be necessary to determine this. Due allowance must be made for this, of course, in the calculations.

Since hydrogen sulfid interferes with this test, it must be prevented from reaching the test paper by the lead acetate reagent in the absorption tubes. The quantitative determination of arsenic in organic substances is not as accurate as that for inorganic substances.

If only qualitative results are desired, it is not necessary to standardize the apparatus. If arsenic compounds are present, the mercuric bromid paper turns a grayish yellow. If the coloration approaches the top of the test paper, the apparatus should be taken to a hood with a good draft, taken apart, and the generator rinsed out. The zinc may be used again. Be sure that no arsin from preceding tests remains in the absorption tubes when beginning a new test.

NOTES ON QUALITATIVE ANALYSIS

1. Lead chlorid is slightly soluble in water, so small amounts of lead escape detection in the first group, but are detected in the second group.

2. Wash with the hot water until the wash water running through the filter will no longer give a precipitate with diluted sulfuric acid.

3. Whenever directed to make a liquid acid or alkaline, test it with litmus paper after mixing thoroughly.

4. The presence of arsenic may be confirmed by the Improved Test for Arsenic described on p. 220, or by the arsin test on p. 126. The latter method should be used only when hoods having good drafts are available. Either the precipitate supposed to be arsenic sulfid, or some of the original solution of the unknown may be used in the test.

5. Boil until chlorin is no longer evolved.

6. The solution should still be slightly acid. If not, add dilute sulfuric acid very carefully to avoid an excess which would prevent the precipitation of the cadmium sulfid. If the precipitate given by the hydrogen sulfid in the next operation is black (containing HgS, CuS, FeS) it should be filtered, washed, and boiled with a mixture of 3 cc. dil. sulfuric acid and 10 cc. water. Filter, dilute the filtrate with water, and pass in hydrogen sulfid. A yellow precipitate proves cadmium present.

7. Boil until the odor of hydrogen sulfid can no longer be detected, as this substance interferes with the test for phosphates.

8. Sodium peroxid almost invariably, and sodium hydroxid very often, contain aluminum and silica. A blank test should be made with these reagents by following the directions of this paragraph and the subsequent test for aluminum, and comparing the final precipitate with that obtained in the regular analysis.

9. If the precipitate is not white and flocculent, confirm the presence of zinc by pouring 5 cc. dil. nitric acid several times through the paper containing the precipitate supposed to be zinc sulfid. To the solution add 8 to 25 drops of cobalt nitrate solution, according to the size of the precipitate, evaporate in a dish nearly to dryness, and neutralize with sodium carbonate solution, using about 0.5 cc. in excess. Evaporate to dryness and ignite gently until the purple color has disappeared. A green residue when cool proves the presence of zinc.

10. Optional method for cobalt and nickel. Evaporate the solution of the sulfids in aqua regia almost to dryness on the water bath; dissolve the residue in water containing a few drops of hydrochloric acid, and then add sodium carbonate solution to alkalinity and a very little acetic acid until the solution is slightly acid to litmus paper. To one-half of this solution add ammonium hydroxid until clear, then a few cubic centimeters of hydrogen peroxid solution and boil the solution for 3 minutes. Add 10 cc. of dimethylglyoxime solution (0.1 g. dissolved in 10 cc. alcohol), and boil the mixture for 3 to 4 minutes. Allow to cool and filter. A red crystalline precipitate proves the presence of nickel.

To the other half of the cobalt nickel solution add 2 cc. of acetic acid and 10 cc. of 30% potassium nitrite solution. Allow to stand 30 minutes unless a precipitate forms sooner. A yellow precipitate (not dark), $K_3Co(NO_2)_6$, proves the presence of cobalt.

11. If the sodium cobaltinitrite and the potassium pyroantimonate tests are not to be used for potassium and sodium, the removal of magnesium is not necessary.

12. If ammonium phosphate was used to precipitate magnesium, the residue, metaphosphates, will be found to be glassy and difficultly soluble in water. Heat it with 1 cc. of concentrated hydrochloric acid for a few minutes, evaporate the acid on a steam bath, and then dissolve the residue in the smallest possible amount of water.

13. If the substance is soluble in water and contains no metals of the first three groups, 1 gram of the substance may be dissolved in 20 to 30 cc. of water, using this solution as directed for the sodium carbonate solution. If metals of the first three groups are present, sodium carbonate solution should be added to the boiling-hot solution of the substance in water as long as a precipitate continues to form.

APPENDIX

CHEMICAL LIBRARY

The books listed first in each division are more desirable for secondary students than those which follow. It is recommended that as many as possible of the books from the popular class be obtained. Books for libraries are subject to a discount from the given list prices.

POPULAR CHEMISTRY

Duncan, *The Chemistry of Commerce*. $1.50. Harpers.
Duncan, *Some Chemical Problems of To-day*. $1.50. Harpers.
Duncan, *New Knowledge*. $1.50. A. S. Barnes.
Martin, *Triumphs and Wonders of Modern Chemistry*. $2.00. Van Nostrand.
Martin, *Modern Chemistry and its Wonders*. $2.00. Van Nostrand.
Baskerville, *Municipal Chemistry*. $5.00. McGraw, Hill.
Martin, E. A., *A Story of a Piece of Coal*. $.50. D. Appleton.
Tower, *The Story of Oil*. $1.00. D. Appleton.
Tidy, *The Story of a Tinder Box*. $.50. S. P. C. K.
Faraday, *The Chemical History of a Candle*. $.75. Harpers.
Bird, *Modern Science Reader*. $1.00. Macmillan.
Sadtler, *The Chemistry of Familiar Things*, 2d ed. $1.60. Lippincott.
Findlay, *Chemistry in the Service of Man*. $2.00. Longmans.
Leighou, *Chemistry of Materials*. $2.00. McGraw, Hill.
Phillips, *The Romance of Modern Chemistry*. $1.25. Seeley.
Jenks, *Chemistry for Young People*. $1.25. Stokes.
Brownlee, *Chemistry of Common Things*. $1.25. Allyn & Bacon.
Stewart, *Chemistry and Its Borderland*. $1.50. Longmans.
Lassar-Cohn, *The Chemistry of Daily Life*. $1.75. Lippincott.
Johnston, *The Chemistry of Common Life*. $1.50. D. Appleton.
Jenks, *Photography for Young People*. $1.25. Stokes.
Abney, *Instruction in Photography*, 11th ed. $2.50. Iliffe.

HISTORICAL

Venable, *History of Chemistry*. $1.00. D. C. Heath.
Brown, *A History of Chemistry*. $3.50. Blakiston.
Muir, *The Story of Alchemy*. $.75. D. Appleton.
Muir, *Heroes of Science, Chemists*. $1.00. S. P. C. K.
Roberts, *Famous Chemists*. $1.00. George Allen & Co.
Roscoe, *John Dalton and the Rise of Modern Chemistry*. $.60. Macmillan.
Thorpe, *Humphrey Davy*. $.60. Macmillan.
Shenstone, *Liebig, His Life and Work*. $.60. Macmillan.
Thompson, *Michael Faraday*. $.60. Macmillan.
Jones, *A New Era in Chemistry*. $2.00. Van Nostrand.
Williams, *A History of Science*, 11 vols. $22.00. Harpers.
(All sciences. A cheap edition by Funk and Wagnalls is not recommended.)
Muir, *A History of Chemical Laws and Theories*. $4.00. Wiley.
Thorpe, *Essays in Historical Chemistry*. $4.00. Macmillan.
Ramsay. *Gases of the Atmosphere*. $2.00. Macmillan.

ADVANCED CHEMISTRY

SMITH, A., *General Chemistry for Colleges.* $2.25. Century Co.
MELLOR, *Modern Inorganic Chemistry.* $2.00. Longmans.
REMSEN, *Inorganic Chemistry.* Advanced. $2.80. Henry Holt.
NOYES, *A Text Book of Chemistry.* $2.25. Henry Holt.
ROSCOE AND SCHORLEMMER, *A Treatise on Chemistry,* 2 vols. $12.50. Macmillan.
BENEDICT, *Chemical Lecture Experiments.* $2.00. Macmillan.
NEWTH, *Chemical Lecture Experiments.* $2.00. Longmans.
JONES, *Introduction to Physical Chemistry.* $1.60. Macmillan.

ORGANIC CHEMISTRY

NOYES, *Organic Chemistry for the Laboratory.* $2.00. Henry Holt.
PERKIN AND KIPPING, *Organic Chemistry.* $2.50. Lippincott.

QUALITATIVE

NOYES, *Qualitative Analysis.* $1.25. Henry Holt.
STIEGLITZ, *Chemical Qualitative Analysis,* 2 vols. $2.50. Century Co.
SCOTT, *Standard Methods in Qualitative Analysis.* $6.00. Van Nostrand.

QUANTITATIVE

CLOWES AND COLEMAN, *Quantitative Analysis,* 10th ed. $3.50. Macmillan.
MAHIN, *Quantitative Analysis.* $3.00. McGraw-Hill.
LINCOLN AND WALTON, *Elementary Quantitative Agricultural Chemical Analysis.* $1.50. Macmillan.
TREADWELL AND HALL, *Quantitative Analysis,* 4th ed. $3.50. Wiley & Sons.

INDUSTRIAL

ROGERS, *Manual of Industrial Chemistry,* 2d ed. $5.00. Van Nostrand.
THORP, *Outlines of Industrial Chemistry,* 3d ed. $3.75. Macmillan.
THORPE, *Dictionary of Applied Chemistry,* 5 vols. $55.00. Macmillan.

PHYSIOLOGICAL

HAWK, *Practical Physiological Chemistry,* 5th ed. $2.50. Blakiston.
SNYDER, *Chemistry of Plant and Animal Life.* $1.25. Macmillan.
FISCHER, *The Physiology of Alimentation.* $2.00. Wiley & Sons.
HALLIBURTON, *Essentials of Chemical Physiology.* $1.75. Longmans.

FOODS

WOODMAN AND NORTON, *Air, Water, and Food from a Sanitary Standpoint,* 4th ed. $2.00. Wiley & Sons.

BAILEY, *Sanitary and Applied Chemistry*, 4th ed. $1.60. Macmillan.
WILEY, *Foods and Their Adulteration*, 2d ed. $4.00. Blakiston.
WINTON, *A Course in Food Analysis*. $1.50. Wiley & Sons.
SNELL, *Elementary Household Chemistry*. $1.50. Macmillan.
OLSEN, *Pure Foods*. $1.25. Ginn & Co.
BRUCE, *Pure Food Tests*. $1.25. Van Nostrand.
SHERMAN, *Chemistry of Food and Nutrition*. $1.50. Macmillan.
THRESH AND PORTER, *Preservatives in Food & Food Products*. $4.50. Blakiston.
LEFFMAN AND BEAM, *Food Analysis*, 2d ed. $2.50. Blakiston.
BAILEY, *Source, Chemistry, and Use of Food Products*. $1.60. Blakiston.
LEACH, *Food Inspection and Analysis*, 3d ed. $7.50. Wiley & Sons.
GRANT, *Chemistry of Bread Making*. $1.40. Longmans.
WINTON, *Microscopy of Vegetable Foods*. $6.50. Wiley & Sons.

WATER

HAZEN, *Clean Water and How to Get It*, 2d ed. $1.50. Wiley & Sons.
MASON, *Water Supply*, 5th ed. $3.75. Wiley & Sons.
WHIPPLE, *Microscopy of Drinking Water*, 3d ed. $4.00. Wiley & Sons.
PRESCOTT-WINSLOW, *Elements of Water Bacteriology*, 3d ed. $1.75. Wiley & Sons.
WHIPPLE, *The Value of Pure Water*. $1.00. Wiley & Sons.
THRESH, *The Examination of Water and Water Supplies*, 2d ed. $5.00. Blakiston.

SANITARY

RITCHIE, *Primer of Sanitation*. $.50. World Book Co.
ELLIOTT, *Household Bacteriology*. $1.00. Am. Sch. Home Ec.
COSGROVE, *History of Sanitation*. $1.50. Stand. San. Mnfg. Co.
RICHARDS, *Conservation by Sanitation*. $2.50. Wiley & Sons.
WOOD, *Sanitation Practically Applied*. $3.00. Wiley & Sons.
KERSHAW, *Sewage Purification and Disposal*. $3.75. Cambridge Press.
CONN, *Practical Dairy Bacteriology*. $1.25. Orange Judd Co.
BUCHANAN, *Household Bacteriology*. $2.25. Macmillan.
CONN, *Agricultural Bacteriology*, 2d ed. $2.00. Blakiston.
HUBBARD, *Ventilation Hand Book*. $2.00. Sheet Metal Pub.
PRICE, *Handbook on Sanitation*. $1.50. Wiley & Sons.

TEXTILES

WOOLMAN AND MCGOWAN, *Textiles*. $2.00. Macmillan.
MATTHEWS, *Textile Fibers*, 3d ed. $4.00. Wiley & Sons.
BARKER AND MIDGLEY, *Analysis of Woven Fibers*. $3.00. Van Nostrand.
FRAPS, *Principles of Dyeing*. $1.60. Macmillan.
DANERETH, *The Methods of Textile Chemistry*. $2.00. Wiley & Sons.

CYCLOPEDIAS

Handbook of Chemistry and Physics. $2.00. Chemical Rubber Co.
HOPKINS, *Scientific American Cyclopedia of Formulas*. $5.00. Munn & Co.

Henley, *Twentieth Century Book of Receipts, Formulas, and Processes.* $3.00. Henley Pub. Co.
Brannt-Wahl, *Techno-Chemical Receipt Book.* $2.00. Baird & Co.
Brannt-Wahl, *National Formulary,* 4th ed. $2.75. Am. Pharm. Assoc.
Merck, *Merck's Index.* Merck & Co., N. Y.
Merck, *U. S. Pharmacopœia,* 9th Revision. $3.00. Blakiston.
Olsen, *Chemical Annual,* 4th ed. $2.50. Van Nostrand.

TECHNICAL

Frary, *Laboratory Manual of Glass Blowing.* $.75. McGraw-Hill.
Boucart, *Insecticides, Fungicides, and Weed Killers.* $4.50. Van Nostrand.
Koller, *Utilization of Waste Products,* 2d ed. $3.00. Van Nostrand.
Chamot, *Elementary Chemical Microscopy.* $3.00. Wiley & Sons.
Lehner, *Ink Manufacture.* $2.50. Van Nostrand.
Deite, *Manual of Perfumery.* $3.00. Van Nostrand.
Deite, *Manual of Toilet Soap Making.* $5.00. Van Nostrand.
Koller, *Cosmetics.* $2.50. Van Nostrand.
Toch, *Chemistry and Technology of Mixed Paints.* $3.00. Van Nostrand.
Watt, *Art of Paper Making,* 3d ed. $3.00. Van Nostrand.
Watt, *Art of Soap Making,* 7th ed. $3.00. Van Nostrand.
Duckwall, *Canning and Preserving of Food Products with Bacteriological Technique.* $3.00. Van Nostrand.
Johnson and Cameron, *Elements of Agricultural Chemistry and Geology.* $2.60. Van Nostrand.

BULLETINS

Hygienic Laboratory Bulletin, No. 56. *Milk and Its Relation to Public Health,* 2d ed. Treasury Dept.
Write to the U. S. Department of Agriculture for a list of the Bureau of Chemistry Bulletins and the Farmers' Bulletins of the Experiment Station. Similarly, address the Experiment Station of the State University.

METRIC SYSTEM

The standard unit of the Metric System is the meter. It is a unit of length. Its multiples and submultiples are obtained decimally.

10 millimeters (mm.) = 1 centimeter (cm.)
10 cm. = 1 decimeter (dm.)
10 dm. = 1 meter (m.)

The unit of volume in the metric system is the liter. It is the **cubic decimeter**. Like the meter it is divided decimally.

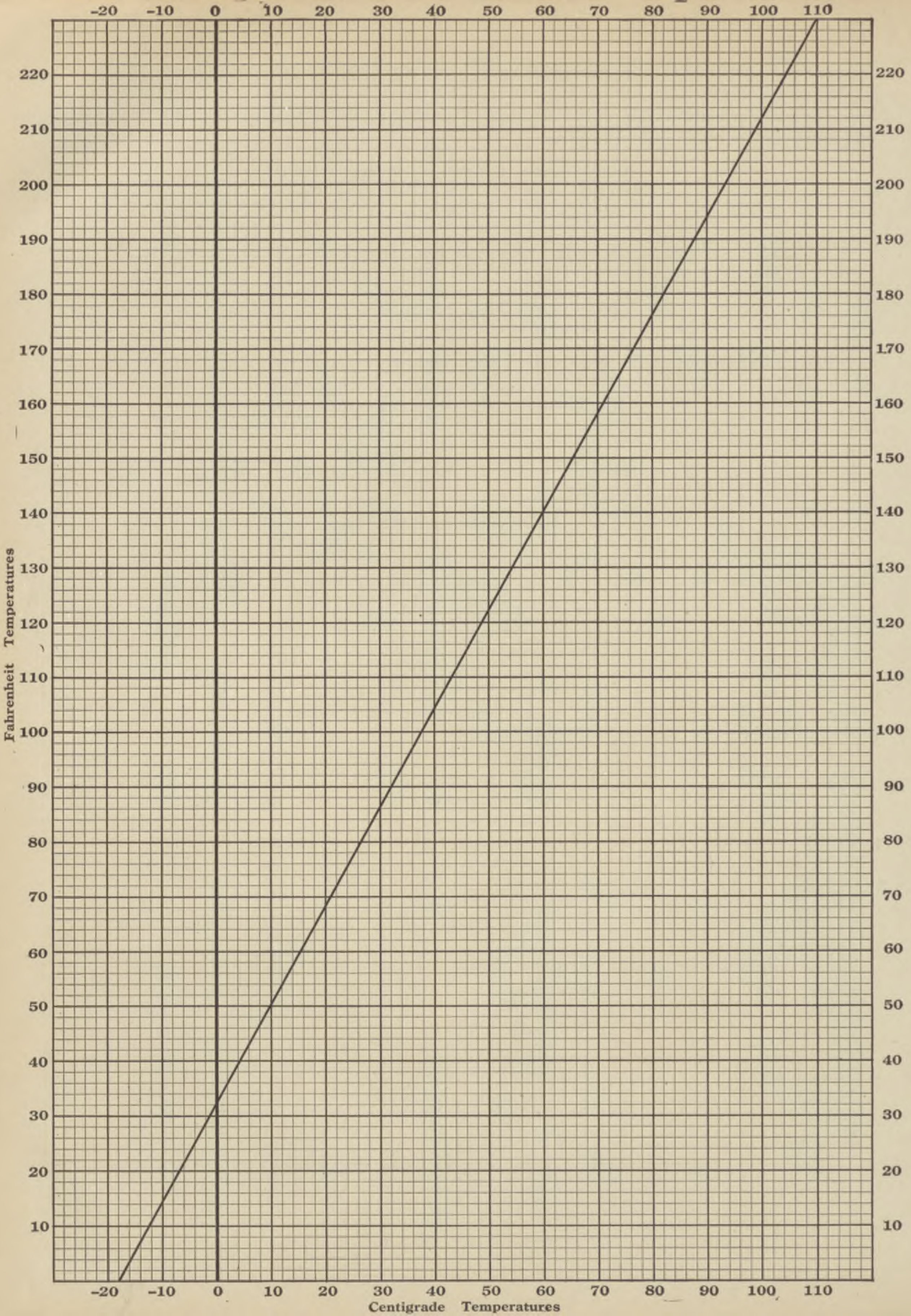
or $$\left.\begin{array}{r}\text{1000 cu. cm. (cc.)}\\ \text{1 cu. decimeter}\end{array}\right\} = 1\text{ liter}$$

The unit of weight is the gram. It is the weight of 1 cc. of water at 4° C. It is divided decimally.

10 milligrams (mg.) = 1 centigram (cg.)
10 cg. = 1 decigram (dg.)
10 dg. = 1 gram (g.)
1000 g. = 1 kilogram (kg.)

CONVERSION FACTORS

To Change	to	Multiply by
Inches	Centimeters	2.54
Feet	Meters	.305
Miles	Kilometers	1.609
Meters	Inches	39.37
Kilometers	Miles	.621
Square inches	Square centimeters	6.452
Square yards	Square meters	.836
Square centimeters	Square inches	.155
Square meters	Square yards	1.196
Cubic inches	Cubic centimeters	16.386
Cubic yards	Cubic meters	.765
Cubic centimeters	Cubic inches	.061
Cubic meters	Cubic yards	1.308
Minims	Grains	.95
Minims	Cubic centimeters	.0617
Fluid ounces	Cubic centimeters	29.57
Quarts	Liters	.946
Cubic centimeters	Fluid ounces	.034
Liters	Quarts	1.057
Grains	Milligrams	64.799
Ounces (av.)	Grams	28.35
Pounds (av.)	Kilograms	.454
Ounces (troy)	Grams	31.103
Pounds (troy)	Kilograms	.373
Grams	Grains	15.432
Kilograms	Pounds	2.205


-20
-10
0
10
20
30
40
50
60
70
80
90
100
110
220
210
200
190
180
170
160
150
140
130
120
110
100
90
80
70
60
50
40
30
20
10
Fahrenheit Temperatures
Centigrade Temperatures

VAPOR PRESSURE OF WATER IN MM. OF MERCURY

Temperature	Pressure	Temperature	Pressure	Temperature	Pressure
− 10	2.2	12	10.4	26	25.
− 5	3.2	13	11.1	27	26.5
0	4.6	14	11.9	28	28.1
+ 1	4.9	15	12.7	29	29.7
2	5.3	16	13.5	30	31.5
3	5.6	17	14.4	35	41.8
4	6.1	18	15.3	40	54.9
5	6.5	19	16.3	50	91.9
6	7.0	20	17.3	60	148.7
7	7.5	21	18.5	70	233.0
8	8.0	22	19.6	80	354.6
9	8.5	23	20.9	90	525.4
10	9.1	24	22.2	100	760.0
11	9.7	25	23.55		

REDUCTION OF BAROMETER READING TO 0° C.

Subtract the Correction

Temperature C°	Correction mm.	Temperature C°	Correction mm.	Temperature C°	Correction mm.
10	1.2	17	2.0	24	2.8
11	1.3	18	2.1	25	3.0
12	1.4	19	2.2	26	3.1
13	1.5	20	2.4	27	3.2
14	1.6	21	2.5	28	3.3
15	1.8	22	2.6	29	3.4
16	1.9	23	2.7	30	3.6

100 cc. OF A GAS AT THE TEMPERATURES AND PRESSURES GIVEN, REDUCED TO 0° C. AND 760 MM.

Temperature	Pressure in Millimeters of Mercury										
	720	**725**	**730**	**735**	**740**	**745**	**750**	**755**	**760**	**765**	**770**
15° C.	89.79	90.42	91.04	91.67	92.29	92.91	93.54	94.16	94.78	95.41	96.03
16°	89.48	90.11	90.73	91.35	91.97	92.59	93.21	93.83	94.45	95.08	95.70
17°	89.17	89.79	90.41	91.03	91.66	92.27	92.89	93.51	94.13	94.75	95.37
18°	88.87	89.48	90.10	90.72	91.34	91.97	92.57	93.19	93.80	94.42	95.04
19°	88.57	89.18	89.79	90.41	91.03	91.66	92.26	92.87	93.48	94.10	94.71
20°	88.26	88.87	89.48	90.10	90.72	91.34	91.94	92.55	93.16	93.78	94.39
21°	87.96	88.57	89.18	89.79	90.41	91.02	91.63	92.24	92.85	93.46	94.07
22°	87.66	88.27	88.87	89.48	90.10	90.72	91.31	91.93	92.53	93.14	93.75
23°	87.36	87.97	88.57	89.18	89.79	90.41	91.00	91.61	92.22	92.82	93.43
24°	87.07	87.67	88.27	88.87	89.48	90.09	90.69	91.31	91.91	92.51	93.11
25°	86.78	87.38	87.98	88.58	89.19	89.79	90.39	91.00	91.60	92.20	92.80

Nat. Number	0	1	2	3	4	5	6	7	8	9	Proportional Parts								
											1	2	3	4	5	6	7	8	9
10	0000	0043	0086	0128	0170	0212	0253	0294	0334	0374	4	8	12	17	21	25	29	33	37
11	0414	0453	0492	0531	0569	0607	0645	0682	0719	0755	4	8	11	15	19	23	26	30	34
12	0792	0828	0864	0899	0934	0969	1004	1038	1072	1106	3	7	10	14	17	21	24	28	31
13	1139	1173	1206	1239	1271	1303	1335	1367	1399	1430	3	6	10	13	16	19	23	26	29
14	1461	1492	1523	1553	1584	1614	1644	1673	1703	1732	3	6	9	12	15	18	21	24	27
15	1761	1790	1818	1847	1875	1903	1931	1959	1987	2014	3	6	8	11	14	17	20	22	25
16	2041	2068	2095	2122	2148	2175	2201	2227	2253	2279	3	5	8	11	13	16	18	21	24
17	2304	2330	2355	2380	2405	2430	2455	2480	2504	2529	2	5	7	10	12	15	17	20	22
18	2553	2577	2601	2625	2648	2672	2695	2718	2742	2765	2	5	7	9	12	14	16	19	21
19	2788	2810	2833	2856	2878	2900	2923	2945	2967	2989	2	4	7	9	11	13	16	18	20
20	3010	3032	3054	3075	3096	3118	3139	3160	3181	3201	2	4	6	8	11	13	15	17	19
21	3222	3243	3263	3284	3304	3324	3345	3365	3385	3404	2	4	6	8	10	12	14	16	18
22	3424	3444	3464	3483	3502	3522	3541	3560	3579	3598	2	4	6	8	10	12	14	15	17
23	3617	3636	3655	3674	3692	3711	3729	3747	3766	3784	2	4	6	7	9	11	13	15	17
24	3802	3820	3838	3856	3874	3892	3909	3927	3945	3962	2	4	5	7	9	11	12	14	16
25	3979	3997	4014	4031	4048	4065	4082	4099	4116	4133	2	3	5	7	9	10	12	14	15
26	4150	4166	4183	4200	4216	4232	4249	4265	4281	4298	2	3	5	7	8	10	11	13	15
27	4314	4330	4346	4362	4378	4393	4409	4425	4440	4456	2	3	5	6	8	9	11	13	14
28	4472	4487	4502	4518	4533	4548	4564	4579	4594	4609	2	3	5	6	8	9	11	12	14
29	4624	4639	4654	4669	4683	4698	4713	4728	4742	4757	1	3	4	6	7	9	10	12	13
30	4771	4786	4800	4814	4829	4843	4857	4871	4886	4900	1	3	4	6	7	9	10	11	13
31	4914	4928	4942	4955	4969	4983	4997	5011	5024	5038	1	3	4	6	7	8	10	11	12
32	5051	5065	5079	5092	5105	5119	5132	5145	5159	5172	1	3	4	5	7	8	9	11	12
33	5185	5198	5211	5224	5237	5250	5263	5276	5289	5302	1	3	4	5	6	8	9	10	12
34	5315	5328	5340	5353	5366	5378	5391	5403	5416	5428	1	3	4	5	6	8	9	10	11
35	5441	5453	5465	5478	5490	5502	5514	5527	5539	5551	1	2	4	5	6	7	9	10	11
36	5563	5575	5587	5599	5611	5623	5635	5647	5658	5670	1	2	4	5	6	7	8	10	11
37	5682	5694	5705	5717	5729	5740	5752	5763	5775	5786	1	2	3	5	6	7	8	9	10
38	5798	5809	5821	5832	5843	5855	5866	5877	5888	5899	1	2	3	5	6	7	8	9	10
39	5911	5922	5933	5944	5955	5966	5977	5988	5999	6010	1	2	3	4	5	7	8	9	10
40	6021	6031	6042	6053	6064	6075	6085	6096	6107	6117	1	2	3	4	5	6	8	9	10
41	6128	6138	6149	6160	6170	6180	6191	6201	6212	6222	1	2	3	4	5	6	7	8	9
42	6232	6243	6253	6263	6274	6284	6294	6304	6314	6325	1	2	3	4	5	6	7	8	9
43	6335	6345	6355	6365	6375	6385	6395	6405	6415	6425	1	2	3	4	5	6	7	8	9
44	6435	6444	6454	6464	6474	6484	6493	6503	6513	6522	1	2	3	4	5	6	7	8	9
45	6532	6542	6551	6561	6571	6580	6590	6599	6609	6618	1	2	3	4	5	6	7	8	9
46	6628	6637	6646	6656	6665	6675	6684	6693	6702	6712	1	2	3	4	5	6	7	7	8
47	6721	6730	6739	6749	6758	6767	6776	6785	6794	6803	1	2	3	4	5	5	6	7	8
48	6812	6821	6830	6839	6848	6857	6866	6875	6884	6893	1	2	3	4	4	5	6	7	8
49	6902	6911	6920	6928	6937	6946	6955	6964	6972	6981	1	2	3	4	4	5	6	7	8
50	6990	6998	7007	7016	7024	7033	7042	7050	7059	7067	1	2	3	3	4	5	6	7	8
51	7076	7084	7093	7101	7110	7118	7126	7135	7143	7152	1	2	3	3	4	5	6	7	8
52	7160	7168	7177	7185	7193	7202	7210	7218	7226	7235	1	2	2	3	4	5	6	7	7
53	7243	7251	7259	7267	7275	7284	7292	7300	7308	7316	1	2	2	3	4	5	6	6	7
54	7324	7332	7340	7348	7356	7364	7372	7380	7388	7396	1	2	2	3	4	5	6	6	7
	0	1	2	3	4	5	6	7	8	9	1	2	3	4	5	6	7	8	9
											Proportional Parts								

Nat. Number	0	1	2	3	4	5	6	7	8	9	Proportional Parts								
											1	2	3	4	5	6	7	8	9
55	7404	7412	7419	7427	7435	7443	7451	7459	7466	7474	1	2	2	3	4	5	5	6	7
56	7482	7490	7497	7505	7513	7520	7528	7536	7543	7551	1	2	2	3	4	5	5	6	7
57	7559	7566	7574	7582	7589	7597	7604	7612	7619	7627	1	2	2	3	4	5	5	6	7
58	7634	7642	7649	7657	7664	7672	7679	7686	7694	7701	1	1	2	3	4	4	5	6	7
59	7709	7716	7723	7731	7738	7745	7752	7760	7767	7774	1	1	2	3	4	4	5	6	7
60	7782	7789	7796	7803	7810	7818	7825	7832	7839	7846	1	1	2	3	4	4	5	6	6
61	7853	7860	7868	7875	7882	7889	7896	7903	7910	7917	1	1	2	3	4	4	5	6	6
62	7924	7931	7938	7945	7952	7959	7966	7973	7980	7987	1	1	2	3	3	4	5	6	6
63	7993	8000	8007	8014	8021	8028	8035	8041	8048	8055	1	1	2	3	3	4	5	5	6
64	8062	8069	8075	8082	8089	8096	8102	8109	8116	8122	1	1	2	3	3	4	5	5	6
65	8129	8136	8142	8149	8156	8162	8169	8176	8182	8189	1	1	2	3	3	4	5	5	6
66	8195	8202	8209	8215	8222	8228	8235	8241	8248	8254	1	1	2	3	3	4	5	5	6
67	8261	8267	8274	8280	8287	8293	8299	8306	8312	8319	1	1	2	3	3	4	5	5	6
68	8325	8331	8338	8344	8351	8357	8363	8370	8376	8382	1	1	2	3	3	4	4	5	6
69	8388	8395	8401	8407	8414	8420	8426	8432	8439	8445	1	1	2	2	3	4	4	5	6
70	8451	8457	8463	8470	8476	8482	8488	8494	8500	8506	1	1	2	2	3	4	4	5	6
71	8513	8519	8525	8531	8537	8543	8549	8555	8561	8567	1	1	2	2	3	4	4	5	5
72	8573	8579	8585	8591	8597	8603	8609	8615	8621	8627	1	1	2	2	3	4	4	5	5
73	8633	8639	8645	8651	8657	8663	8669	8675	8681	8686	1	1	2	2	3	4	4	5	5
74	8692	8698	8704	8710	8716	8722	8727	8733	8739	8745	1	1	2	2	3	4	4	5	5
75	8751	8756	8762	8768	8774	8779	8785	8791	8797	8802	1	1	2	2	3	3	4	5	5
76	8808	8814	8820	8825	8831	8837	8842	8848	8854	8859	1	1	2	2	3	3	4	5	5
77	8865	8871	8876	8882	8887	8893	8899	8904	8910	8915	1	1	2	2	3	3	4	4	5
78	8921	8927	8932	8938	8943	8949	8954	8960	8965	8971	1	1	2	2	3	3	4	4	5
79	8976	8982	8987	8993	8998	9004	9009	9015	9020	9025	1	1	2	2	3	3	4	4	5
80	9031	9036	9042	9047	9053	9058	9063	9069	9074	9079	1	1	2	2	3	3	4	4	5
81	9085	9090	9096	9101	9106	9112	9117	9122	9128	9133	1	1	2	2	3	3	4	4	5
82	9138	9143	9149	9154	9159	9165	9170	9175	9180	9186	1	1	2	2	3	3	4	4	5
83	9191	9196	9201	9206	9212	9217	9222	9227	9232	9238	1	1	2	2	3	3	4	4	5
84	9243	9248	9253	9258	9263	9269	9274	9279	9284	9289	1	1	2	2	3	3	4	4	5
85	9294	9299	9304	9309	9315	9320	9325	9330	9335	9340	1	1	2	2	3	3	4	4	5
86	9345	9350	9355	9360	9365	9370	9375	9380	9385	9390	1	1	2	2	3	3	4	4	5
87	9395	9400	9405	9410	9415	9420	9425	9430	9435	9440	0	1	1	2	2	3	3	4	4
88	9445	9450	9455	9460	9465	9469	9474	9479	9484	9489	0	1	1	2	2	3	3	4	4
89	9494	9499	9504	9509	9513	9518	9523	9528	9533	9538	0	1	1	2	2	3	3	4	4
90	9542	9547	9552	9557	9562	9566	9571	9576	9581	9586	0	1	1	2	2	3	3	4	4
91	9590	9595	9600	9605	9609	9614	9619	9624	9628	9633	0	1	1	2	2	3	3	4	4
92	9638	9643	9647	9652	9657	9661	9666	9671	9675	9680	0	1	1	2	2	3	3	4	4
93	9685	9689	9694	9699	9703	9708	9713	9717	9722	9727	0	1	1	2	2	3	3	4	4
94	9731	9736	9741	9745	9750	9754	9759	9763	9768	9773	0	1	1	2	2	3	3	4	4
95	9777	9782	9786	9791	9795	9800	9805	9809	9814	9818	0	1	1	2	2	3	3	4	4
96	9823	9827	9832	9836	9841	9845	9850	9854	9859	9863	0	1	1	2	2	3	3	4	4
97	9868	9872	9877	9881	9886	9890	9894	9899	9903	9908	0	1	1	2	2	3	3	4	4
98	9912	9917	9921	9926	9930	9934	9939	9943	9948	9952	0	1	1	2	2	3	3	4	4
99	9956	9961	9965	9969	9974	9978	9983	9987	9991	9996	0	1	1	2	2	3	3	3	4
	0	1	2	3	4	5	6	7	8	9	1	2	3	4	5	6	7	8	9
											Proportional Parts								

LIST OF ELEMENTS,[1] 1911

The more common elements are printed in heavy type.

Name	Symbol	Valences	At. Wt.	Name	Symbol	Valences	At. Wt.
Aluminum	Al	3	27.1	Neodymium	Nd	3	144.3
Antimony	Sb	3–5	120.2	Neon	Ne	0	20
Argon	A	0	39.9	**Nickel**	Ni	2–3	58.7
Arsenic	As	3–5	75	Niobium[3]	Nb	5–3	93.5
Barium	Ba	2	137.4	Niton	Ni	0	222
Beryllium[2]	Be	2	9.1	**Nitrogen**	N	3–5	14
Bismuth	Bi	3	208	Osmium	Os	8–6–4	190.9
Boron	B	3	11	**Oxygen**	O	2	16
Bromin	Br	1–5	79.9	Palladium	Pd	4–2	106.7
Cadmium	Cd	2	112.4	**Phosphorus**	P	5–3	31
Cæsium	Cs	1	132.8	**Platinum**	Pt	4–2	195
Calcium	Ca	2	40.1	**Potassium**	K	1	39.1
Carbon	C	4–2	12	Praseodymium	Pr	3	140.6
Cerium	Ce	4–3	140.3	Radium	Ra	2	226.4
Chlorin	Cl	1–5–7–3	35.5	Rhodium	Rh	3	102.9
Chromium	Cr	3–6–7–2	52	Rubidium	Rb	1	85.5
Cobalt	Co	2–3	59	Ruthenium	Ru	4–3–6–8	101.7
Copper	Cu	2–1	63.6	Samarium	Sa	3	150.4
Dysprosium	Dy	3	162.5	Scandium	Sc	3	44.1
Erbium	Er	3	167.4	Selenium	Se	2–6–4	79.2
Europium	Eu	3	152	**Silicon**	Si	4	28.3
Fluorin	F	1	19	**Silver**	Ag	1	107.9
Gadolinium	Gd	3	157.3	**Sodium**	Na	1	23
Gallium	Ga	3	69.9	**Strontium**	Sr	2	87.7
Germanium	Ge	4–2	72.5	**Sulfur**	S	2–4–6	32.1
Gold	Au	1–3	197.2	Tantalum	Ta	5	181.5
Helium	He	0	4	Tellurium	Te	2–4–6	124.3
Hydrogen	H	1	1	Terbium	Tb	3	159.2
Indium	In	3	114.8	Thallium	Tl	1–3	204
Iodin	I	1–5–7	127	Thorium	Th	4–3	232.4
Iridium	Ir	3–4–2	193.1	Thullium	Tu	3	168.5
Iron	Fe	3–2	55.9	**Tin**	Sn	2–4	119
Krypton	Kr	0	83	Titanium	Ti	4	48.1
Lanthanum	La	3	139	Tungsten	W	6–4–2	184
Lead	Pb	2–4	207.1	Uranium	U	6–4	238.5
Lutecium	Lu	2	174	Vanadium	V	5–3	51
Lithium	Li	1	7	Xenon	Xe	0	130.7
Magnesium	Mg	2	24.3	Ytterbium	Yb	3	172
Manganese	Mn	2–7–4–6–3	54.9	Yttrium	Y	3	89
Mercury	Hg	2–1	200.6	**Zinc**	Zn	2	65.4
Molybdenum	Mo	6–4–2	96	Zirconium	Zr	4	90.6

---

[1] Calculated from the basis, O = 16.

[2] Also called Glucinum.

[3] Also called Columbium.

PHYSICAL PROPERTIES OF THE ELEMENTS

Element	Specific Gravity	Melting Point	Boiling Point
Aluminum	2.67	657	1800
Antimony	6.72	625	1440
Argon		– 187.9	– 186.1
Arsenic	5.70	500	449.5
Barium	3.78	850	
Beryllium	1.85	960	
Bismuth	9.65	269	1420
Boron	2.58	3000	3500
Bromin	3.10	– 7.3	59
Cadmium	8.72	320	778
Cæsium	1.88	26.37	270
Calcium	1.85	760	
Carbon	3.55d 2.25g 1.70a	3500	
Cerium	6.65	623	
Chlorin		– 102	– 33.6
Chromium	6.47	1515	2200
Cobalt	8.55	1530	
Copper	8.93	1045	2310
Erbium			
Fluorin		– 223	– 187
Gadolinium			
Gallium	5.95	30.15	
Germanium	5.47	900	1350
Gold	19.32	1065	
Helium		– 271	– 269.5
Hydrogen		– 252.5	– 259
Indium	7.40	155	800
Iodin	4.95	114	184.4
Iridium	21.83	1950	
Iron	7.84	1800	2450
Krypton		– 169	– 151.7
Lanthanum	6.10	810	
Lead	11.37	327	1525
Lithium	0.59	186	1400
Magnesium	1.75	632.6	1120
Manganese	7.25	1300	1900
Mercury	13.57	– 39.4	357
Molybdenum	8.65	2500	
Neodymium		840	
Neon		– 253	
Nickel	8.85	1484	
Niobium	7.06	1950	
Nitrogen		– 214	– 194
Osmium	22.48	2500	
Oxygen		– 227	– 183
Palladium	11.40	1586	
Phosphorus	1.82Y 2.11R	44.2	269
Platinum	21.48	1770	
Potassium	0.87	62.5	667
Praseodymium		940	
Radium			
Rhodium	11.50	2000	
Rubidium	1.52	38.5	696
Ruthenium	11.20	1900	
Samarium		1400	
Scandium		?	
Selenium	4.48	217	680
Silicon	2.49	1200	
Silver	10.50	955	1955
Sodium	0.97	95.6	742
Strontium	2.55	800	
Sulfur	2.06O 1.96M	115–119	448.5
Tantalum	10.70	2250	
Tellurium	6.24	452	1390
Terbium			
Thallium	11.85	302	1600–1800
Thorium	11.00	1700	
Thulium			
Tin	7.29	233	2270
Titanium	4.5	1800	
Tungsten	17.20	1700	
Uranium	18.68	800	
Vanadium	5.50	1630	
Xenon		— 140	— 109.1
Ytterbium			
Yttrium			
Zinc	7.15	420	918
Zirconium	6.4	2350	

SPECIFIC GRAVITY OF HYDROCHLORIC, NITRIC, SULFURIC ACIDS, AND SODIUM HYDROXID

Specific Gravity at $\frac{15^\circ}{4^\circ}$ (Vacuo)	Per Cent by Weight				Specific Gravity at $\frac{15^\circ}{4^\circ}$ (Vacuo)	Per Cent by Weight		
	HCl	HNO_3	H_2SO_4	NaOH		HNO_3	H_2SO_4	NaOH
1.000	0.16	0.10	0.09	.06	1.205	33.09	27.95	18.26
1.005	1.15	1.00	0.83	.43	1.210	33.82	28.58	18.72
1.010	2.14	1.90	1.57	.85	1.215	34.55	29.21	19.17
1.015	3.12	2.80	2.30	1.27	1.220	35.28	29.84	19.63
1.020	4.13	3.70	3.03	1.69	1.225	36.03	30.48	20.08
1.025	5.15	4.60	3.76	2.11	1.230	36.78	31.11	20.54
1.030	6.15	5.50	4.49	2.53	1.235	37.53	31.70	20.99
1.035	7.15	6.38	5.23	2.95	1.240	38.29	32.28	21.44
1.040	8.16	7.26	5.96	3.37	1.245	39.05	32.86	21.89
1.045	9.16	8.13	6.67	3.80	1.250	39.82	33.43	22.34
1.050	10.17	8.99	7.37	4.23	1.255	40.58	34.00	22.79
1.055	11.18	9.84	8.07	4.67	1.260	41.34	34.57	23.24
1.060	12.19	10.68	8.77	5.11	1.265	42.10	35.14	23.69
1.065	13.19	11.51	9.47	5.56	1.270	42.87	35.71	24.14
1.070	14.17	12.33	10.19	6.01	1.275	43.64	36.29	24.59
1.075	15.16	13.15	10.90	6.47	1.280	44.41	36.87	25.04
1.080	16.15	13.95	11.60	6.93	1.285	45.18	37.45	25.50
1.085	17.13	14.74	12.30	7.39	1.290	45.95	38.03	25.96
1.090	18.11	15.53	12.99	7.85	1.295	46.72	38.61	26.42
1.095	19.06	16.32	13.67	8.31	1.300	47.49	39.19	26.88
1.100	20.01	17.11	14.35	8.77	1.305	48.26	39.77	27.35
1.105	20.97	17.89	15.03	9.22	1.310	49.07	40.35	27.82
1.110	21.92	18.67	15.71	9.68	1.315	49.89	40.93	28.29
1.115	22.86	19.45	16.36	10.13	1.320	50.71	41.50	28.76
1.120	23.82	20.23	17.01	10.59	1.325	51.53	42.08	29.23
1.125	24.78	21.00	17.66	11.04	1.330	52.37	42.66	29.70
1.130	25.75	21.77	18.31	11.49	1.335	53.22	43.20	30.18
1.135	26.70	22.54	18.96	11.95	1.340	54.07	43.74	33.14
1.140	27.66	23.31	19.61	12.40	1.345	54.93	44.28	31.15
1.145	28.61	24.08	20.26	12.86	1.350	55.79	44.82	31.64
1.150	29.57	24.84	20.91	13.31	1.355	56.66	45.35	32.14
1.155	30.55	25.60	21.55	13.76	1.360	57.57	45.88	32.64
1.160	31.52	26.36	22.19	14.21	1.365	58.48	46.41	33.14
1.165	32.49	27.12	22.83	14.66	1.370	59.39	46.94	33.64
1.170	33.46	27.88	23.47	15.11	1.375	60.30	47.47	34.14
1.175	34.42	28.63	24.12	15.56	1.380	61.27	48.00	34.64
1.180	35.39	29.38	24.76	16.01	1.385	62.24	48.53	35.14
1.185	36.31	30.13	25.40	16.46	1.390	63.23	49.06	35.64
1.190	37.23	30.88	26.04	16.91	1.395	64.25	49.59	36.14
1.195	38.16	31.62	26.68	17.36	1.400	65.30	50.11	36.64
1.200	39.11	32.36	27.32	17.81				

SPECIFIC GRAVITY OF HYDROCHLORIC, NITRIC, SULFURIC ACIDS, AND SODIUM HYDROXID—*Continued*

Specific Gravity at $\frac{15°}{4°}$ (Vacuo)	Per Cent by Weight			Specific Gravity at $\frac{15°}{4°}$ (Vacuo)	Per Cent by Weight	Specific Gravity at $\frac{15°}{4°}$ (Vacuo)	Per Cent by Weight
	HNO_3	H_2SO_4	NaOH		H_2SO_4		H_2SO_4
1.405	66.40	50.63	37.14	1.580	66.71	1.755	82.00
1.410	67.50	51.15	37.65	1.585	67.13	1.760	82.44
1.415	68.63	51.66	38.16	1.590	67.59	1.765	82.88
1.420	69.80	52.15	38.67	1.595	68.05	1.770	83.32
1.425	70.98	52.63	39.18	1.600	68.51	1.775	83.90
1.430	72.17	53.11	39.70	1.605	68.97	1.780	84.50
1.435	73.39	53.59	40.22	1.610	69.43	1.785	85.10
1.440	74.68	54.07	40.74	1.615	69.89	1.790	85.70
1.445	75.98	54.55	41.26	1.620	70.32	1.795	86.30
1.450	77.28	55.03	41.79	1.625	70.74	1.800	86.90
1.455	78.60	55.50	42.30	1.630	71.16	1.805	87.60
1.460	79.98	55.97	42.82	1.635	71.57	1.810	88.30
1.565	81.42	56.43	43.34	1.640	71.99	1.815	89.05
1.470	82.90	56.90	43.86	1.645	72.40	1.820	90.05
1.475	84.45	57.37	44.38	1.650	72.82	1.825	91.00
1.480	86.05	57.83	44.90	1.655	73.23	1.830	92.10
1.485	87.70	58.28	45.42	1.660	73.64	1.835	93.43
1.490	89.60	58.74	45.94	1.665	74.07	1.840	95.60
1.495	91.60	59.22	46.46	1.670	74.51	1.8405	95.95
1.500	94.09	59.70	46.98	1.675	74.97	1.8410	97.00
1.505	96.39	60.18	47.50	1.680	75.42	1.8415	97.70
1.510	98.10	60.65	48.02	1.685	75.86	1.8410	98.20
1.515	99.07	61.12	48.54	1.690	76.30	1.8405	98.70
1.520	99.67	61.59	49.06	1.695	76.73	1.8400	99.20
1.525		62.06	49.58	1.700	77.17	1.8395	99.45
1.530		62.53	50.10	1.705	77.60	1.8390	99.70
1.535		63.00		1.710	78.04	1.8385	99.95
1.540		63.43		1.715	78.48		
1.545		63.85		1.720	78.92		
1.550		64.26		1.725	79.36		
1.555		64.67		1.730	79.80		
1.560		65.08		1.735	80.24		
1.565		65.49		1.740	80.68		
1.570		65.90		1.745	81.12		
1.575		66.30		1.750	81.56		

IMPORTANT TEMPERATURES

1. Zero absolute	– 273° C.
2. Boiling point of helium	– 268.5
3. Melting point of hydrogen	– 260
4. Boiling point of hydrogen	– 252.5
5. Melting point of nitrogen	– 214
6. Boiling point of nitrogen	– 194
7. Boiling point of oxygen	– 182.5
8. Freezing point of grain alcohol	– 130
9. Freezing point of mercury	– 39.5
10. Boiling point of ammonia	– 33
11. Freezing point of water	0
12. Average room temperature	20
13. Normal temperature of the body	37
14. Melting point of Wood's metal	60
15. Boiling point of grain alcohol	78.5
16. Boiling point of water	100
17. Melting point of rhombic sulfur	112.8
18. Melting point of tin	232
19. Melting point of lead	327
20. Melting point of zinc	419
21. Boiling point of sulfur	444.7
22. Incipient red heat	550
23. Dull red heat	650
24. Melting point of aluminum	660° C.
25. Melting point of common salt	792
26. Red heat	800
27. Melting point of silver	960
28. Bright red heat	1000
29. Melting point of gold	1063
30. Melting point of copper	1083
31. Yellow heat	1200
32. White heat	1350
33. Temperature of glass furnace	1375
34. Melting point of nickel	1452
35. Melting point of iron	1520
36. Melting point of sand	1600
37. Melting point of platinum	1755
38. Melting point of corundum	2000 ±
39. Melting point of iridium	2300
40. Temperature of oxyhydrogen flame	2500 ±
41. Melting point of osmium	2500
42. Temperature of oxy-acetylene flame	2700 +
43. Melting point of tungsten	3000
44. Temperature of alumino-thermy	3500 ±
45. Temperature of electric arc	4000 ±
46. Temperature of sun	6000

SPECIFIC GRAVITY AND WEIGHT OF GASES AT ATMOSPHERIC PRESSURE (760 mm.) AND 0°C.

	Air = 1	Vapor Density H = 1	Grams per Liter	Lb. per Cu. Ft.	Cubic Ft. per Lb.
Air	1.0000	14.444	1.2931	.080723	12.388
Oxygen, O_2	1.1052	15.963	1.4291	.08921	11.209
Hydrogen, H_2	0.0692	1.000	0.0895	.00559	178.931
Nitrogen, N_2	0.9701	14.012	1.2544	.07831	12.770
Carbon monoxid, CO	0.9671	13.968	1.2505	.07807	12.810
Carbon dioxid, CO_2	1.5197	21.950	1.9650	.12267	8.152
Methane, marsh-gas, CH_4	0.5530	7.987	0.7150	.04464	22.429
Ethylene, C_2H_4	0.9674	13.973	1.2510	.07809	12.805
Acetylene, C_2H_2	0.8982	12.973	1.1614	.07251	13.792
Ammonia, NH_3	0.5889	8.506	0.7615	.04754	21.036
Water vapor, H_2O	0.6218	8.981	0.8041	.05020	19.922
Chlorin, Cl_2	2.4910	35.79	3.22	.201	4.97
Sulfur dioxid, SO_2	2.264	32.77	2.927	.183	5.48

CENTIGRADE AND FAHRENHEIT SCALES

To reduce a Fahrenheit temperature to its equivalent in Centigrade degrees: Subtract 32°, and multiply the remainder by $\frac{5}{9}$.

To reduce a Centigrade temperature to its Fahrenheit equivalent: Multiply by $\frac{9}{5}$, and add 32° to the product.

QUALITATIVE SOLUBILITY OF THE COMMON SALTS, ETC.

	Acetates	Arsenates	Arsenites	Borates	Bromids	Carbonates	Chlorates	Chlorids	Chromates	Cyanids	Ferricyanids	Ferrocyanids	Fluorids	Hydroxids	Iodids	Nitrates	Oxalates	Oxids	Phosphates	Silicates	Sulphates	Sulphids	Tartrates
Al	W	A		A	W		W	W					W	A	W	W	A	A	A	sa	W	A	W
NH_4	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W		W		W	W	W
Sb	W	A	A		sw			sw	A				W		sw		A	A	sw		sw	A	A
Ba	W	A	A	A	W	A	W	W	A	sw	sw	sw	sa	W	W	W	A	W	A	A	I	W	A
Bi	W	A		A	sw	A	W	sw	A	sw		I	W	A	A	W	A	A	A		W	A	A
Cd	W	A		sw	W	A	W	W	A	A	A		sw	A	W	W	A	A	A	A	W	A	sw
Ca	W	A	A	A	W	A	W	W	sw	W	W	W	sa	sw	W	W	A	sw	A	A	sa	W	A
Cr	W	A		A	W		W	W	A	A			W	A	W	W	sw	A	A	A	W	A	W
Co	W	A	A	A	W	A	W	W	A	sa	I	I	sw	A	W	W	A	A	A	A	W	A	W
Cu	W	A	A	A	W	A	W	W	W	A	I	I	A	A	W	W	A	A	A	A	W	A	W
Au					W			W		A				A	A			A				A	
H	W	W		W	W		W	W		W	W	W	W	W	W	W	W	W	W	W	W	sw	W
Fe″	W	A	A	A	W	A	W	W		sa	I	I	sw	A	W	W	A	A	A	A	W	A	sw
Fe‴	W	A	A	A	W		W	W	W		W	I	W	A	W	W	A	A	A	A	W	A	W
Pb	W	A	A	A	sa	A	W	sa	sa	A	sw	A	A	A	sw	W	A	A	A	A	sa	A	A
Mg	W	A	A	sw	W	A	W	W	W	W	W	W	sa	A	W	W	A	A	A	A	W	A	W
Mn	W	A	A	A	W	A	W	W	W	A	I	A	A	A	W	W	sw	A	A	A	W	A	sw
Hg′	sw	A	A		sa	A	W	sa	A	W			W		A	W	A	A	A		sw	A	sw
Hg″	W	A	A		W	A	W	W	sw	W	A	A	sw		A	W	A	A	A		W	A	A
Ni	W	A	A	A	W	A	W	W	A	sa	I	I	sw	A	W	W	A	A	A	A	W	A	A
K	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	
Ag	sw	A	A	A	I	A	W	I	A	I	I	I	W		I	W	A	A	A		sw	A	A
Na	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W	W
Sn⁗	W	A			W			W				I	W	A	W	sw	W	A	A		W	A	
Sn″	W	A	A	A	W		W	W	A		I	I	W	A	W		A	A	A		W	A	A
Sr	W	A	A	A	W	A	W	W	sw	W	W	W	sa	A	W	W	A	sw	A	A	sa	W	A
Zn	W	A		A	W	A	W	W	W	A	A	sa	sw	A	W	W	A	A	A	A	W	A	A

W = soluble in water; A = insoluble in water, but soluble in acids; sw = sparingly soluble in water, but soluble in acids; sa = insoluble in water, and only sparingly soluble in acids; I = insoluble in both water and acids.

APPROXIMATE SOLUBILITY, IN GRAMS, IN 100 GRAMS OF WATER AT ORDINARY TEMPERATURE

	Fluorid	Chlorid	Bromid	Iodid	Chlorate	Bromate	Iodate	Nitrate	Sulphate	Carbonate	Chromate	Hydroxid
Ammonium	—	36.4	73.3	166.	—	—	—	175.	75.	27.	—	—
Barium	0.16	37.	104.	201.	35.4	0.8	0.05	8.7	.00023	.0023	.0004	3.7
Calcium	.0016	73.2	143.	200.	179.	85.	0.25	122.	0.2	.0013	0.4	0.17
Lead	0.07	1.5	0.6	0.08	150.	1.3	.002	51.7	0.004	.0001	.0002	0.01
Magnesium	.0076	56.	103.	148.	126.	43.	6.9	74.3	35.4	0.1	73.	0.001
Mercuric	—	7.2	1.2	0.004	—	0.15	—	—	—	—	—	—
Potassium	93.	33.	66.	137.	6.6	6.4	7.6	30.	11.	108.	63.	143.
Silver	195.	.00016	.00001	.0000003	12.3	0.59	.004	220.	0.55	0.003	.0025	0.01
Sodium	4.4	36.	89.	178.	97.	37.	8.3	84.	16.8	19.4	61.	116.
Strontium	0.012	51.	97.	169.	173.	30.	0.25	66.3	0.011	.0011	0.12	0.77
Zinc	0.005	204.	478.	419.	184.	58.	0.83	118.	53.	0.004	—	.0005

HEAT OF FORMATION AND HEAT OF SOLUTION IN KILOGRAM CENTIGRADE UNITS (1000 Calories)

Name	Formula	Heat of Formation	Heat of Solution in Water
Ozone	O_3	– 30.	—
Water (liquid)	H_2O	68.4	—
Hydrogen peroxid	H_2O_2	45.2	—
Hydrogen chlorid	HCl	22.	17.3
Hydrogen bromid	HBr	12.	19.9
Hydrogen iodid	HI	– 6.1	19.2
Hydrogen sulfid	H_2S	2.7	4.6
Sulfur dioxid	SO_2	71.	7.7
Sulfuric acid	H_2SO_4	193.1	17.8
Ammonia	NH_3	12.	8.4
Nitrogen tetroxid	N_2O_4	– 2.6	—
Nitrogen dioxid	NO_2	– 7.7	—
Nitric oxid	NO	– 21.6	—
Carbon disulfid	CS_2	– 19.6	—
Carbon dioxid	CO_2	97.	—
Carbon monoxid	CO	29.	—
Phosphorus, red, from yellow form	—	27.3	—
Phosphorus pentoxid	P_2O_5	370.	36.
Potassium hydroxid	KOH	103.2	13.3
Potassium carbonate	K_2CO_3	281.	6.5
Potassium nitrate	KNO_3	119.	– 8.5
Sodium chlorid	$NaCl$	97.6	– 1.2
Sodium hydroxid	$NaOH$	102.	9.9
Sodium carbonate	Na_2CO_3	272.6	5.6
Ammonium chlorid	NH_4Cl	75.8	– 4.
Calcium hydroxid	$Ca(OH)_2$	215.	3.
Magnesium hydroxid	$Mg(OH)_2$	217.	—
Aluminum hydroxid	$Al(OH)_3$	297.	—
Ferric hydroxid	$Fe(OH)_3$	198.	—
Ferrous-ferric oxid	Fe_3O_4	265.	—
Zinc oxid	ZnO	86.	—
Cupric oxid	CuO	37.	—
Mercuric oxid	HgO	20.7	—
Silver oxid	Ag_2O	6.	—
Lead monoxid	PbO	50.	—
Ethane	C_2H_6	23.3	—
Ethylene	C_2H_4	– 14.6	—
Acetylene	C_2H_2	– 58.1	—

DEGREE OF IONIZATION OF IONOGENS

Except where otherwise specified, the figures give the fraction ionized in a normal, aqueous solution (usually at 18° C.). Subtraction of the figures from unity gives the extent to which the ions will unite when brought together in normal concentration. At greater dilutions the ionization is greater and the union of ions less.

ACIDS

HNO_3	0.82	$H.H_2PO_4(N/2)$	0.17
HNO_3 (conc.)	0.09	$H.HC_2O_4$ (N/10)	0.50
HCl	0.78	$H.HC_4H_4O_6(N/10)$	0.08
HCl (conc.)	0.13	$H.C_2H_3O_2$	0.004
HCl (N/2)	0.85	$H.C_2H_3O_2(N/10)$	0.013
H_2SO_4	0.51	$H.HCO_3$ (N/10)	0.0017
H_2SO_4 (conc.)	0.007	$H.HCO_3$ (N/25)	0.0021
HBr (N/2)	0.90	$H.HS$ (N/10)	0.0007
HI (N/2)	0.90	$H.H_2BO_3$ (N/10)	0.0001
$HClO_3$ (N/2)	0.88	HNC (N/10)	0.0001
$HMnO_4$ (N/2)	0.93		

BASES

KOH	0.77	$Sr(OH)_2$ (N/64)	0.93
$NaOH$	0.73	$Ba(OH)_2$ (N/64)	0.92
$Ba(OH)_2$	0.69	$AgOH$ (N/1783)	0.39
NH_4OH	0.004	HOH	0.0000001
$Ca(OH)_2(N/64)$	0.90		

SALTS

KCl	0.75	Na_2CO_3	0.40
KBr (N/32)	0.92	$Na.HCO_3$	0.52
$KClO_3$ (N/2)	0.79	$Na_2.HPO_4$ (N/32)	0.83
KNO_3	0.64	$NaC_2H_3O_2$	0.53
K_2SO_4	0.53	$Na_2C_4H_4O_6$ (N/32)	0.78
K_2CO_3	0.49	$BaCl_2$	0.57
$KMnO_4$ (N/32)	0.92	$CaSO_4$ (N/100)	0.63
$K_2Cr_2O_7$ (N/32)	0.94	$CuSO_4$	0.22
NH_4Cl	0.74	$AgNO_3$	0.58
$NaCl$	0.67	$CdSO_4$	0.22
Na_2SO_4	0.44	$ZnSO_4$	0.24
Na_2SO_3 (N/32)	0.82	$HgCl_2$	(< 0.01)

TABLE SHOWING PER CENT OF TOTAL SOLIDS CORRESPONDING TO QUEVENNE LACTOMETER READINGS AND PER CENT OF FAT

Per Cent of Fat	Lactometer Readings at 60° F.														
	22	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**	**31**	**32**	**33**	**34**	**35**	**36**
0.0	5.50	5.75	6.00	6.25	6.50	6.75	7.00	7.25	7.50	7.75	8.00	8.25	8.50	8.75	9.00
0.1	5.62	5.87	6.12	6.37	6.62	6.87	7.12	7.37	7.62	7.87	8.12	8.37	8.62	8.87	9.12
0.2	5.74	5.99	6.24	6.49	6.74	6.99	7.24	7.49	7.74	7.99	8.24	8.49	8.74	8.99	9.24
0.3	5.86	6.11	6.36	6.61	6.86	7.11	7.36	7.61	7.86	8.11	8.36	8.61	8.86	9.11	9.36
0.4	5.98	6.23	6.48	6.73	6.98	7.23	7.48	7.73	7.98	8.23	8.48	8.73	8.98	9.23	9.48
0.5	6.10	6.35	6.60	6.85	7.10	7.35	7.60	7.85	8.10	8.35	8.60	8.85	9.10	9.35	9.60
0.6	6.22	6.47	6.72	6.97	7.22	7.47	7.72	7.97	8.22	8.47	8.72	8.97	9.22	9.47	9.72
0.7	6.34	6.59	6.84	7.09	7.34	7.59	7.84	8.09	8.34	8.59	8.84	9.09	9.34	9.59	9.84
0.8	6.46	6.71	6.96	7.21	7.46	7.71	7.96	8.21	8.46	8.71	8.96	9.21	9.46	9.71	9.96
0.9	6.58	6.83	7.08	7.33	7.58	7.83	8.08	8.33	8.58	8.83	9.08	9.33	9.58	9.83	10.08
1.0	6.70	6.95	7.20	7.45	7.70	7.95	8.20	8.45	8.70	8.95	9.20	9.45	9.70	9.95	10.20
1.1	6.82	7.07	7.32	7.57	7.82	8.07	8.32	8.57	8.82	9.07	9.32	9.57	9.82	10.07	10.32
1.2	6.94	7.19	7.44	7.69	7.94	8.19	8.44	8.69	8.94	9.19	9.44	9.69	9.94	10.19	10.44
1.3	7.06	7.31	7.56	7.81	8.06	8.31	8.56	8.81	9.06	9.31	9.56	9.81	10.06	10.31	10.56
1.4	7.18	7.43	7.68	7.93	8.18	8.43	8.68	8.93	9.18	9.43	9.68	9.93	10.18	10.43	10.68
1.5	7.30	7.55	7.80	8.05	8.30	8.55	8.80	9.05	9.30	9.55	9.80	10.05	10.30	10.55	10.80
1.6	7.42	7.67	7.92	8.17	8.42	8.67	8.92	9.17	9.42	9.67	9.92	10.17	10.42	10.67	10.92
1.7	7.54	7.79	8.04	8.29	8.54	8.79	9.04	9.29	9.54	9.79	10.04	10.29	10.54	10.79	11.04
1.8	7.66	7.91	8.16	8.41	8.66	8.91	9.16	9.41	9.66	9.91	10.16	10.41	10.66	10.91	11.17
1.9	7.78	8.03	8.28	8.53	8.78	9.03	9.28	9.53	9.78	10.03	10.28	10.54	10.78	11.04	11.29
2.0	7.90	8.15	8.40	8.65	8.90	9.15	9.40	9.65	9.90	10.15	10.40	10.66	10.91	11.16	11.41
2.1	8.02	8.27	8.52	8.77	9.02	9.27	9.52	9.77	10.02	10.27	10.52	10.78	11.03	11.28	11.53
2.2	8.14	8.39	8.64	8.89	9.14	9.39	9.64	9.89	10.14	10.39	10.64	10.90	11.15	11.40	11.65
2.3	8.26	8.51	8.76	9.01	9.26	9.51	9.76	10.01	10.26	10.51	10.76	11.02	11.27	11.52	11.77
2.4	8.38	8.63	8.88	9.13	9.38	9.63	9.88	10.13	10.38	10.63	10.88	11.14	11.39	11.64	11.89
2.5	8.50	8.75	9.00	9.25	9.50	9.75	10.00	10.25	10.50	10.75	11.00	11.26	11.51	11.76	12.01
2.6	8.62	8.87	9.12	9.37	9.62	9.87	10.12	10.37	10.62	10.87	11.12	11.38	11.63	11.88	12.13
2.7	8.74	8.99	9.24	9.49	9.74	9.99	10.24	10.49	10.74	10.99	11.24	11.50	11.75	12.00	12.25
2.8	8.86	9.11	9.36	9.61	9.86	10.11	10.36	10.61	10.86	11.11	11.37	11.62	11.87	12.12	12.37
2.9	8.98	9.23	9.48	9.73	9.98	10.23	10.48	10.73	10.98	11.23	11.49	11.74	11.99	12.24	12.49
3.0	9.10	9.35	9.60	9.85	10.10	10.35	10.60	10.85	11.10	11.36	11.61	11.86	12.11	12.36	12.61
3.1	9.22	9.47	9.72	9.97	10.22	10.47	10.72	10.97	11.23	11.48	11.73	11.98	12.23	12.48	12.74
3.2	9.34	9.59	9.84	10.09	10.34	10.59	10.84	11.09	11.35	11.60	11.85	12.10	12.35	12.61	12.86
3.3	9.46	9.71	9.96	10.21	10.46	10.71	10.96	11.22	11.47	11.72	11.97	12.22	12.48	12.73	12.98
3.4	9.58	9.83	10.08	10.33	10.58	10.83	11.09	11.34	11.59	11.84	12.09	12.34	12.60	12.85	13.10
3.5	9.70	9.95	10.20	10.45	10.70	10.95	11.21	11.46	11.71	11.96	12.21	12.46	12.72	12.97	13.22
3.6	9.82	10.07	10.32	10.57	10.82	11.08	11.33	11.58	11.83	12.08	12.33	12.58	12.84	13.09	13.34
3.7	9.94	10.19	10.44	10.69	10.94	11.20	11.45	11.70	11.95	12.20	12.45	12.70	12.96	13.21	13.46
3.8	10.06	10.31	10.56	10.81	11.06	11.32	11.57	11.82	12.07	12.32	12.57	12.82	13.08	13.33	13.58
3.9	10.18	10.43	10.68	10.93	11.18	11.44	11.69	11.94	12.19	12.44	12.69	12.94	13.20	13.45	13.70
4.0	10.30	10.55	10.80	11.05	11.30	11.56	11.81	12.06	12.31	12.56	12.81	13.06	13.32	13.57	13.83
4.1	10.42	10.67	10.92	11.17	11.42	11.68	11.93	12.18	12.43	12.68	12.93	13.18	13.44	13.69	13.95
4.2	10.54	10.79	11.04	11.29	11.54	11.80	12.05	12.30	12.55	12.80	13.05	13.31	13.56	13.82	14.07
4.3	10.66	10.91	11.16	11.41	11.66	11.92	12.17	12.42	12.67	12.92	13.18	13.43	13.68	13.94	14.19
4.4	10.78	11.03	11.28	11.53	11.78	12.04	12.29	12.54	12.79	13.04	13.30	13.55	13.80	14.06	14.31
4.5	10.90	11.15	11.40	11.65	11.90	12.16	12.41	12.66	12.91	13.16	13.42	13.67	13.92	14.18	14.43
4.6	11.02	11.27	11.52	11.78	12.03	12.28	12.53	12.78	13.03	13.28	13.54	13.79	14.04	14.30	14.55
4.7	11.14	11.40	11.65	11.90	12.15	12.40	12.65	12.90	13.15	13.40	13.66	13.91	14.16	14.42	14.67
4.8	11.27	11.52	11.77	12.02	12.27	12.52	12.77	13.02	13.27	13.52	13.78	14.03	14.28	14.54	14.79
4.9	11.39	11.64	11.89	12.14	12.39	12.64	12.89	13.14	13.39	13.64	13.90	14.15	14.40	14.66	14.91
5.0	11.51	11.76	12.01	12.26	12.51	12.76	13.01	13.26	13.51	13.76	14.02	14.27	14.52	14.78	15.03
5.1	11.63	11.88	12.13	12.38	12.63	12.88	13.13	13.38	13.63	13.89	14.14	14.39	14.64	14.90	15.15
5.2	11.75	12.00	12.25	12.50	12.75	13.00	13.25	13.50	13.75	14.01	14.26	14.51	14.76	15.02	15.27
5.3	11.87	12.12	12.37	12.62	12.87	13.12	13.37	13.62	13.87	14.13	14.38	14.63	14.88	15.14	15.39
5.4	11.99	12.24	12.49	12.74	12.99	13.24	13.49	13.74	14.00	14.25	14.50	14.76	15.01	15.26	15.51
5.5	12.11	12.36	12.61	12.86	13.11	13.36	13.61	13.86	14.12	14.37	14.62	14.88	15.13	15.38	15.63
5.6	12.23	12.48	12.73	12.98	13.23	13.48	13.73	13.99	14.24	14.49	14.75	15.00	15.25	15.50	15.75
5.7	12.35	12.60	12.85	13.10	13.35	13.60	13.85	14.11	14.36	14.61	14.87	15.12	15.37	15.62	15.87
5.8	12.47	12.72	12.97	13.22	13.47	13.72	13.97	14.23	14.48	14.74	14.99	15.24	15.49	15.74	15.99
5.9	12.59	12.84	13.09	13.34	13.59	13.84	14.10	14.35	14.60	14.86	15.11	15.36	15.61	15.86	16.12
6.0	12.71	12.96	13.21	13.46	13.71	13.96	14.22	14.47	14.72	14.98	15.23	15.48	15.73	15.98	16.24

TABLE FOR CORRECTING QUEVENNE LACTOMETER READINGS FOR TEMPERATURE (Vieth)

Degrees of Thermometer (Fahrenheit)	Degrees of Lactometer															
	20	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**	**31**	**32**	**33**	**34**	**35**
45	19.0	19.9	20.9	21.9	22.9	23.8	24.8	25.8	26.7	27.7	28.6	29.5	30.4	31.3	32.2	33.0
46	19.0	20.0	21.0	22.0	22.9	23.9	24.9	25.9	26.8	27.8	28.7	29.6	30.5	31.4	32.3	33.1
47	19.1	20.0	21.0	22.0	23.0	24.0	24.9	25.9	26.8	27.8	28.7	29.6	30.5	31.4	32.3	33.2
48	19.1	20.1	21.1	22.1	23.1	24.0	25.0	26.0	26.9	27.9	28.8	29.7	30.6	31.5	32.4	33.4
49	19.2	20.2	21.2	22.2	23.2	24.1	25.1	26.1	27.0	28.0	28.9	29.8	30.7	31.6	32.5	33.5
50	19.2	20.2	21.2	22.2	23.2	24.1	25.1	26.1	27.0	28.0	29.0	29.9	30.9	31.8	32.7	33.6
51	19.3	20.3	21.3	22.3	23.3	24.2	25.2	26.2	27.1	28.1	29.1	30.0	31.0	31.9	32.9	33.8
52	19.4	20.3	21.3	22.3	23.3	24.3	25.2	26.2	27.2	28.2	29.1	30.1	31.1	32.0	33.0	33.9
53	19.4	20.4	21.4	22.4	23.4	24.4	25.3	26.3	27.3	28.3	29.2	30.2	31.2	32.1	33.1	34.0
54	19.5	20.5	21.5	22.5	23.5	24.5	25.4	26.4	27.4	28.4	29.3	30.3	31.3	32.3	33.2	34.2
55	19.6	20.6	21.6	22.6	23.6	24.6	25.5	26.5	27.5	28.5	29.4	30.4	31.4	32.4	33.3	34.3
56	19.7	20.7	21.7	22.7	23.6	24.6	25.6	26.6	27.6	28.6	29.6	30.5	31.5	32.5	33.5	34.5
57	19.8	20.8	21.8	22.8	23.7	24.7	25.7	26.7	27.7	28.7	29.7	30.6	31.6	32.6	33.6	34.6
58	19.9	20.9	21.9	22.8	23.8	24.8	25.8	26.8	27.8	28.8	29.8	30.8	31.7	32.7	33.7	34.7
59	19.9	20.9	21.9	22.9	23.9	24.9	25.9	26.9	27.9	28.9	29.9	30.9	31.9	32.9	33.9	34.9
60	**20**	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**	**31**	**32**	**33**	**34**	**35**
61	20.1	21.1	22.1	23.1	24.1	25.1	26.1	27.1	28.1	29.1	30.1	31.2	32.2	33.2	34.2	35.2
62	20.2	21.2	22.2	23.2	24.2	25.2	26.2	27.3	28.3	29.3	30.3	31.3	32.3	33.3	34.3	35.3
63	20.2	21.3	22.3	23.3	24.3	25.3	26.3	27.4	28.4	29.4	30.4	31.4	32.5	33.5	34.5	35.5
64	20.3	21.4	22.4	23.4	24.4	25.4	26.5	27.5	28.5	29.5	30.5	31.5	32.6	33.6	34.6	35.6
65	20.4	21.5	22.5	23.5	24.5	25.5	26.6	27.6	28.6	29.6	30.7	31.6	32.7	33.8	34.8	35.8
66	20.5	21.6	22.6	23.6	24.6	25.6	26.7	27.7	28.7	29.8	30.8	31.7	32.9	33.9	34.9	35.9
67	20.6	21.7	22.7	23.7	24.7	25.7	26.8	27.8	28.8	29.9	30.9	31.8	33.0	34.0	35.0	36.1
68	20.7	21.8	22.8	23.8	24.9	25.9	27.0	28.0	29.0	30.1	31.1	32.0	33.2	34.2	35.2	36.2
69	20.9	22.0	23.0	24.0	25.0	26.0	27.1	28.1	29.1	30.2	31.2	32.2	33.3	34.3	35.3	36.4
70	21.0	22.1	23.1	24.1	25.1	26.1	27.2	28.2	29.2	30.3	31.3	32.4	33.4	34.5	35.5	36.5
71	21.1	22.2	23.2	24.2	25.2	26.2	27.3	28.3	29.4	30.4	31.5	32.5	33.6	34.6	35.6	36.7
72	21.2	22.3	23.3	24.3	25.3	26.4	27.4	28.4	29.5	30.5	31.6	32.6	33.7	34.7	35.8	36.8
73	21.3	22.4	23.4	24.4	25.5	26.5	27.5	28.6	29.7	30.7	31.8	32.8	33.9	34.9	36.0	37.0
74	21.5	22.5	23.5	24.6	25.6	26.6	27.7	28.7	29.8	30.9	31.9	33.0	34.0	35.1	36.1	37.2
75	21.6	22.6	23.7	24.7	25.7	26.8	27.8	28.9	29.9	31.0	32.1	33.1	34.2	35.2	36.3	37.3

When the temperature of the milk is above 60° F., the lactometer reading is too low, and if the temperature of the milk is below 60° F. the lactometer reading is too high. The lactometer reading for any given temperature may be corrected by finding the number which is in the same column as the observed lactometer reading in the first row, and also in the row indicated by the observed temperature (Fahrenheit). Thus a lactometer reading 33 at 50° F. should read 31.8 when the temperature of the milk is 60° F. — the temperature at which liquids are usually measured.

To convert readings on the New York Board of Health lactometer to Quevenne degrees, multiply them by .29.

PERCENTAGE OF ALCOHOL

Sp. Gr. 15.6° C.	% by Vol.	% by Wt.	Sp. Gr. 15.6° C.	% by Vol.	% by Wt.	Sp. Gr. 15.6° C.	% by Vol.	% by Wt.	Sp. Gr. 15.6° C.	% by Vol.	% by Wt.	Sp. Gr. 15.6° C.	% by Vol.	% by Wt.	Sp. Gr. 15.6° C.	% by Vol.	% by Wt.	Sp. Gr. 15.6° C.	% by Vol.	% by Wt.
1.0000	0.0	0.00	0.9811	15.0	12.13	0.9654	30.0	24.66	0.9436	45.0	37.84	0.9131	60.0	52.16	0.8772	75.0	67.87	0.8340	90.0	85.66
0.9992	0.5	0.40	06	15.5	12.54	48	30.5	25.08	28	45.5	38.30	20	60.5	52.66	59	75.5	68.43	23	90.5	86.32
85	1.0	0.79	01	16.0	12.95	42	31.0	25.51	19	46.0	38.75	09	61.0	53.15	45	76.0	68.99	06	91.0	86.97
78	1.5	1.19	0.9796	16.5	13.37	36	31.5	25.94	10	46.5	39.21	0.9098	61.5	53.66	32	76.5	69.54	0.8289	91.5	87.63
70	2.0	1.59	91	17.0	13.78	30	32.0	26.37	01	47.0	39.67	87	62.0	54.16	18	77.0	70.11	72	92.0	88.29
63	2.5	1.99	86	17.5	14.19	24	32.5	26.80	0.9392	47.5	40.13	76	62.5	54.67	04	77.5	70.68	55	92.5	88.95
56	3.0	2.39	81	18.0	14.60	17	33.0	27.23	82	48.0	40.60	65	63.0	55.16	0.8690	78.0	71.25	38	93.0	89.62
49	3.5	2.80	76	18.5	15.02	11	33.5	27.66	73	48.5	41.06	54	63.5	55.67	76	78.5	71.82	20	93.5	90.30
42	4.0	3.20	71	19.0	15.43	04	34.0	28.09	64	49.0	41.52	43	64.0	56.18	63	79.0	72.39	00	94.0	91.00
35	4.5	3.60	66	19.5	15.84	0.9598	34.5	28.52	54	49.5	41.99	32	64.5	56.69	50	79.5	72.96	0.8180	94.5	91.70
28	5.0	4.00	61	20.0	16.26	91	35.0	28.96	46	50.0	42.47	20	65.0	57.20	36	80.0	73.53	61	95.0	92.41
22	5.5	4.40	56	20.5	16.67	84	35.5	29.30	36	50.5	42.93	08	65.5	57.72	22	80.5	74.12	42	95.5	93.11
15	6.0	4.80	51	21.0	17.09	77	36.0	29.83	26	51.0	43.41	0.8996	66.0	58.24	10	81.0	74.68	23	96.0	93.81
09	6.5	5.21	46	21.5	17.51	70	36.5	30.26	17	51.5	43.88	84	66.5	58.76	0.8597	81.5	75.25	02	96.5	94.55
02	7.0	5.61	41	22.0	17.92	63	37.0	30.70	06	52.0	44.35	73	67.0	59.27	85	82.0	75.82	0.8081	97.0	95.29
0.9896	7.5	6.02	36	22.5	18.34	56	37.5	31.14	0.9296	52.5	44.84	61	67.5	59.79	71	82.5	76.41	60	97.5	96.02
90	8.0	6.42	30	23.0	18.76	49	38.0	31.58	85	53.0	45.31	48	68.0	60.32	57	83.0	77.00	37	98.0	96.79
84	8.5	6.83	25	23.5	19.17	41	38.5	32.03	75	53.5	45.80	35	68.5	60.86	43	83.5	77.58	14	98.5	97.57
78	9.0	7.23	20	24.0	19.59	34	39.0	32.46	64	54.0	46.27	23	69.0	61.39	28	84.0	78.19	0.7989	99.0	98.37
72	9.5	7.64	15	24.5	20.01	26	39.5	32.90	53	54.5	46.75	10	69.5	61.92	13	84.5	78.79	62	99.5	99.22
66	10.0	8.04	10	25.0	20.43	19	40.0	33.35	43	55.0	47.24	0.8898	70.0	62.44	0.8498	85.0	79.40	0.7938	100.0	100.00
60	10.5	8.45	04	25.5	20.85	11	40.5	33.79	32	55.5	47.72	86	70.5	62.97	83	85.5	80.01	(Absolute alcohol)		
55	11.0	8.86	0.9699	26.0	21.27	03	41.0	34.24	21	56.0	48.21	74	71.0	63.51	68	86.0	80.62			
49	11.5	9.27	94	26.5	21.69	0.9495	41.5	34.68	11	56.5	48.69	62	71.5	64.04	53	86.5	81.23			
44	12.0	9.67	88	27.0	22.11	87	42.0	35.13	0.9199	57.0	49.18	50	72.0	64.59	38	87.0	81.85			
38	12.5	10.08	83	27.5	22.54	79	42.5	35.58	89	57.5	49.69	37	72.5	65.13	22	87.5	82.47			
33	13.0	10.49	77	28.0	22.96	70	43.0	36.03	78	58.0	50.16	24	73.0	65.68	05	88.0	83.12			
27	13.5	10.90	72	28.5	23.38	62	43.5	36.48	67	58.5	50.65	11	73.5	66.22	0.8389	88.5	83.74			
22	14.0	11.31	66	29.0	23.81	54	44.0	36.93	55	59.0	51.16	0.8798	74.0	66.77	72	89.0	84.39			
17	14.5	11.72	60	29.5	24.23	45	44.5	37.39	43	59.5	51.66	85	74.5	67.31	57	89.5	85.01			

COMPOSITION OF ALLOYS

Type of Alloy	Name	Approximate Composition in Parts per 100
Copper Alloys	Brass	Cu (67–80); Zn (20–33); Pb½ ±
	Bronze	Cu 84; Pb 1 ±; Sn 5; Zn 10
	Aluminum bronze	Cu (90–95); Al (5–10)
	Manganese bronze	Cu 70; Mn 30
	Gun metal	Cu 90; Sn 10
	Bell metal	Cu 75; Sn 25
	Dutch metal	Cu 80; Zn 20
	Copper coins	Cu 95; Sn 4; Zn 1
	5-cent piece	Cu 75; Ni 25
	German Silver	Cu (55–60); Zn 20; Ni (20–25)
	Manganin	Cu 84; Ni 4; Mn 12
Iron Alloys	Steel	Fe (99–99.75); C (0.25–1)
	Nickel steel	Fe 94; C 1 ±; Ni 5 ±
	Chrome steel	Fe (94½–96½); C (1–1½); Cr (2½–4 ±)
	Tungsten steel	Fe 94; C 1 ±; W 5 ±
	Manganese steel	Fe 91; C 1 ±; Mn 8 ±
	Molybdenum steel	Fe (89–94); C 1 ±; Mo (5–10)
Lead-Tin Alloys	Solder	Pb (33–66); Sn (33–66)
	Pewter	Sn 75; Pb 25
	Shot metal	Pb 99; As 1 ±
	Type metal	Pb 75; Sb 20; Sn 5
	Britannia metal	Sn 85; Sb 10; Cu 5
Silver Alloys	Sterling silver	Ag 92.5; Cu 7.5
	Coin silver	Ag 90; Cu 10
Gold Alloys	U. S. gold coinage 21.6 K	Au 90; Cu 10
	18 K gold	Au 75; Cu 25
	14 K gold	Au 58.3; Cu 41.7
	10 K gold	Au 41.6; Cu 58.4
Fusible Alloys	Wood's metal (melts at 60.5° C.)	Bi 50; Pb 25; Cd 12.5; Sn 12.5
	Rose's metal (melts at 94° C.)	Bi 82; Pb 9; Sn 9
Miscellaneous Alloys	Babbitt's antifriction metal	Zn 69; Sn 19; Pb 5; As 4; Sb 3
	Magnalium	Al 98; Mg 2

PER CENT OF SUGARS IN SOME COMMON FRUITS

Fruit	Sucrose	Glucose and Fructose	Fruit	Sucrose	Glucose and Fructose
Pineapple	11.3	2.0	Pears	0.3	8.4
Apricots	6.0	2.7	Gooseberries	0.0	6.4
Strawberries	6.3	5.0	Cherries	0.0	10.0
Oranges	4.2	4.3	Figs	0.0	11.5
Apples	5.3	8.7	Grapes	0.0	17.3

INDIVIDUAL APPARATUS

One set of the articles marked (*) may be used by all classes with a considerable saving in the cost of equipment.

A liberal reserve stock of glassware should be kept on hand to provide for breakage.

3 Beakers, 75, 150, 250 cc.
1 Blowpipe
1 Bottle, gas generating, to fit No. 5 stopper
1 *Burette clamp
1 *Burner and 2 ft. gas tubing, ¼ in. inside diameter
1 *Calcium chlorid tube, 15 cm.
1 *Clay triangle, 6 cm. side
1 Cobalt glass plate, 5 × 5 cm., double thickness
1 *Combustion spoon, ½ in. shallow bowl
1 *Crucible, porcelain, 26 cc.
1 Delivery tube, glass
1 Evaporating dish, porcelain, 100 cc.
1 Flask, 250 cc.
1 Funnel, 7.5 cm.
3 *250 cc. gas collecting jars and glass covers
1 Graduate, cylindrical, 25 cc.
1 Ignition tube, 15 cm.
1 L-tube, glass
1 Mortar and pestle
1 *Platinum wire, 22 gauge, 7 cm. long, sealed in glass tube
1 Plug, glass
1 *Ring stand with 3 rings
1 Rubber tube, connector, 4 in. long, 3/16 in. inside diameter
2 Rubber stoppers, 2-hole, Nos. 4 and 5
1 *Slotted disk (like weights of a platform scale)
1 Stirring rod, glass, 20 cm.
6 Test-tubes, 10 cm.
6 Test-tubes, 15 cm.
1 Test-tube, 20 cm.
1 Test-tube brush
1 Test-tube clamp
1 *Test-tube rack
1 Thistle tube
2 Watch glasses, 6½ and 8 cm.

GENERAL APPARATUS

The quantities are estimated for a class of 20.

10 Burettes, 50 cc. in tenths
10 Condensers, glass, Liebig or spiral
10 Dropping funnels, 60 cc.
1 Eudiometer
6 Flasks, volumetric, 2 each 250, 500, and 1000 cc.
20 Gas measuring tubes, 100 cc. in fifths
6 Graduates, 2 each 250, 500, and 1000 cc.
1 Hofmann electrolysis apparatus
2 Hydrometers, 1 each 0.700 to 1.000 and 1.000 to 2.000
10 Pipettes, 10 cc.
12 Pipettes, 2 each 1, 2, 5, 25, 50, and 100 cc.
10 Retorts, glass, 250 cc., tubulated
5 Thermometers, reading − 10° to 250° C.
4 Balances, chemical, sensitive to 1 mg.
1 Barometer
1 Boyle's law apparatus
1 Charles' law apparatus
2 lb. Copper wire, #16
1 Diffusion apparatus
4 Electric cells, dry or storage
1 Hammer
1 Hessian crucible, #5
1 Induction coil, 6 mm. spark
20 Iron dishes, or 10 × 10 cm. iron sheets, or old tablespoons
1 Lamp chimney, large
5 Magnifiers, magnification 24 diameters
1 Sodium capsule or spoon
1 Thread, spool, #20
1 Tile, 10 × 10 cm.
5 Trip balances with set weights 10 g. to 500 g.
10 Water baths, copper

SPECIAL APPARATUS

1 Blast lamp and foot bellows, or large size Meker burner
1 Camera
2 Drying ovens
6 Enameled stew pans, 1 qt.
1 Extraction apparatus, Soxhlet's or Wiley's
25 Extraction shells (1 box)
1 Iron kettle, 1 gal.
2 Kjeldahl flasks, 250 cc. and connecting bulbs
5 Lactometers, Quevenne's
1 Microscope
1 Milk scale, Richmond's
1 Milk tester, Babcock's
10 Photographic trays
1 Touchstone and set of gold needles

REAGENTS AND SOLUTIONS

Figures in solution ratio column indicate cubic centimeters of liquids and grams of solids. The (*) indicates special directions for preparing solutions.

Name	Formula	Mol. Wt.	Amount for a Class of 20–25		Solution Ratio	Cheapest Quantity	Remarks
			Required Exps.	Optional Exps.			
Acid, acetic	HCO_2CH_3	60.03	3 lb.	2 lb.	—	5 lb.	99.5%, glacial, C. P.
Acid, acetic	make up from glacial acid				2 : 5		30%, dilute
Acid, acetic, anhydrid	$(CH_3CO)_2O$	102		4 oz.			
Acid, boric	H_3BO_3	62		1 oz.	1 : 20	5 lb.	"boracic" acid, C. P.
Acid, butyric	$HCO_2C_3H_7$	88.1		4 oz.	—	1 lb.	50%
Acid, chloroplatinic	H_2PtCl_6	409.5			1 : 50		
Acid, citric	$C_3H_4(OH)(COOH)_3.H_2O$	210		4 oz.		1 lb.	powd. C. P.
Acid, formic	HCO_2H	46		4 oz.		1 lb.	Sp. gr. 1.20 85% pure
Acid, hydrochloric	HCl	36.46	24 lb.	12 lb.	—	6 lb.	Sp. gr. 1.20 39% conc. C. P.
Acid, hydrochloric	make up from concentrated				1 : 4	—	dilute
Acid, hydrochloric	make up from concentrated				1 : 40		1%
Acid, nitric	HNO_3	63.01	7 lb.	14 lb.	—	7 lb.	52% conc. C. P.
Acid, nitric	make up from concentrated				1 : 5	—	dilute
Acid, oxalic	$(COOH)_2.2\ H_2O$	0.92	4 oz.			1 lb.	pure cryst.
Acid, phosphoric	H_3PO_4	98.04	1 lb.			5 lb.	85%, sirupy
Acid, picric	$C_6H_2OH(NO_2)_3$	185.02		1 oz.	1 : 200		tech.
Acid, pyrogallic	$C_6H_3(OH)_3$	126.1	1 oz.				resublimed
Acid, salicylic	$C_7H_6O_3$	138.1		1 oz.			U. S. P.
Acid, sulfanilic	$C_6H_4SO_3H.NH_2.2\ H_2O$			1 oz.	*		
Acid, sulfuric	H_2SO_4	98	36 lb.	18 lb.	—	9 lb.	Sp. gr. 1.84 conc.
Acid, sulfuric	make up from concentrated				1 : 6	—	dilute
Acid, tannic	$C_{14}H_{10}O_9$	322.1		4 oz.			
Acid, valerianic	$C_5H_{10}O_2.Aq$	102.1		4 oz.			
Albumen				4 oz.			egg soluble
Alcohol, ethyl	C_2H_5OH	46	½ pt.	1 pt.	—	1 gal.	95% Obtained tax free for schools in 25 gal. lots
Alcohol, methyl	CH_3OH	32	¼ pt.	1 qt.	—	1 gal.	
Alcohol, methyl	deodorized methyl alcohol			1 qt.	—	1 gal.	Columbian spirits
Alizarin			4 oz.	—	1 : 4		25% paste
Alum, potassium	$K_2SO_4.Al_2(SO_4)_3.24\ H_2O$	949.2	4 oz.			5 lb.	technical
Aluminum	Al	27.1	8 oz.				scraps or #18 wire
Aluminum sulfate	$Al_2(SO_4)_3.18\ H_2O$	666.7	4 oz.		1 : 10	5 lb.	tech.
"Ammo"			1 lb.				washing powder
Ammonium carbonate	$NH_4HCO_3+NH_4NH_2CO_2$	157.1		1 lb.	*	1 lb.	C. P.
Ammonium chlorid	NH_4Cl	53.5		1 lb.	1 : 10	1 lb.	C. P.
Ammonium chlorid			1 lb.	4 oz.		5 lb.	tech.
Ammonium hydroxid	NH_4OH		20 lb.	10 lb.		4 lb.	Sp. gr. 0.90 concentrated 28%
Ammonium hydroxid	make up from concentrated				1 : 2		10% dil. Sp. gr. 0.96
Ammonium hydroxid	make up from concentrated				1 : 4		
Ammonium molybdate	$(NH_4)_6Mo_7O_{24}.4\ H_2O$	268.08	2 oz.	2 oz.	*		
Ammonium nitrate	NH_4NO_3	80.1	1 oz.	8 oz.	—	1 lb.	pure
Ammonium oxalate	$(NH_4)_2C_2O_4.H_2O$	142.10	8 oz.	1 lb.	1 : 25	1 lb.	C. P. use warm water

Name	Formula	Mol. Wt.	Amount for a Class of 20–25		Solution Ratio	Cheapest Quantity	Remarks
			Required Exps.	Optional Exps.			
Ammonium polysulfid	$(NH_4)_2S_x$				*		
Ammonium sulfate	$(NH_4)_2SO_4$	132.2	1 lb.			5 lb.	tech.
Ammonium sulfate				1 oz.	1 : 4	1 lb.	C. P.
Amyl acetate	$C_5H_{11}(CO_2CH_3)$	130.1		4 oz.	—	1 gal.	tech.
Amyl alcohol	$C_5H_{11}OH$	88.1		4 oz.	—		tech.
Aniline	$C_6H_5NH_2$	93.1		8 oz.		1 lb.	tech.
Antimony	Sb	120	1 oz.		—	1 lb.	tech. powder
Antimony chlorid	$SbCl_3$	226.4	2 oz.	2 oz.	1 : 20	1 lb.	C. P. sol. in conc. HCl
Antimony sulfid	Sb_2S_3	336.2		4 oz.	—	1 lb.	tech., red
Apatite				8 oz.		5 lb.	tech.
Arsenious chlorid	$AsCl_3$	181.4	2 oz.		1 : 20	8 oz.	
Arsenious oxid	As_2O_3	198		8 oz.		1 lb.	tech.
Asbestos			4 oz.				shredded
Asbestos paper			2 sq. ft.	4 sq. ft.			
Asphaltum				1 lb.		1 lb.	chips
Barium chlorid	$BaCl_2.2\ H_2O$	244.3	2 oz.	8 oz.	1 : 40	1 lb.	C. P.
Barium hydroxid	$Ba(OH)_2$	315.6		8 oz.	1 : 20	1 lb.	pure
Barium nitrate	$Ba(NO_3)_2$	261.5		8 oz.	1 : 20	1 lb.	C. P.
Benzol	C_6H_6	78.1	4 oz.	1 lb.	—	1 gal.	90%
Benzine	C_8H_{18} and C_9H_{20}			1 lb.		1 gal.	
Beta-naphthol	$C_{10}H_7OH$	144.06		1 oz.	1 : 2000		Use also equal weight of NaOH
Bismarck brown				1 oz.	1 : 1000		
Bismuth nitrate	$Bi(NO_3)_3.5\ H_2O$	484.7		4 oz.	1 : 20	1 lb.	C. P. Dissolve in dilute HNO_3
Bismuth chlorid	$BiCl_3$	314.9		4 oz.	1 : 20	1 lb.	C. P. Dissolve in dilute HCl
Bone ash				1 lb.		5 lb.	tech.
Borax	$Na_2B_4O_7.10\ H_2O$	382.16	1 lb.		1 : 10	5 lb.	pure
Bromin	Br_2	159.84	4 oz.	4 oz.	sat. sol.		
Cadmium bromid	$CdBr_2.4\ H_2O$	344.2		4 oz.	1 : 10	1 lb.	C. P.
Cadmium nitrate	$Cd(NO_3)_2.4\ H_2O$	308.6		4 oz.	1 : 10	1 lb.	C. P.
Calcium	Ca	40.00	1 oz.				
Calcium carbid	CaC_2	64		4 oz.	—	1 lb.	tech.
Calcium carbonate	$CaCO_3$	100		8 oz.		1 lb.	C. P.
Calcium carbonate			8 oz.			5 lb.	tech.
Calcium carbonate				5 lb.		5 lb.	marble chips
Calcium chlorid	$CaCl_2$	111	2 lb.			5 lb.	anhydrous, granular
Calcium chlorid	$CaCl_2.6\ H_2O$	219	8 oz.		1 : 10	1 lb.	
Calcium fluorid	CaF_2	78	1 lb.			5 lb.	native
Calcium hydroxid	$Ca(OH)_2$	74	1 lb.			5 lb.	hydrated lime
Calcium hydroxid					sat. sol.		limewater
Calcium oxid	CaO	56	5 lb.	5 lb.		5 lb.	quicklime
Calcium nitrate	$Ca(NO_3)_2.4\ H_2O$	236.2		4 oz.		1 lb.	C. P.
Calcium sulfate	$CaSO_4.2\ H_2O$	172.1	4 oz.	4 oz.	sat. sol.	1 lb.	C. P.
Calico			2 yd.				figured
Camphor gum				2 oz.			
Candles			2 doz.				paraffin
Carbon tetrachlorid	CCl_4	153.8	2 oz.	1 lb.		1 lb.	
Carbon disulfid	CS_2	76.1	1 lb.			5 lb.	C. P.
Carbon disulfid			1 lb.			5 lb.	pure
Cement				5 lb.			Portland
Chlorin water					sat. sol.		Keep in opaque bottles
Charcoal			1 lb.			5 lb.	lump

REAGENTS AND SOLUTIONS — *Continued*

Name	Formula	Mol. Wt.	Amount for a Class of 20–25		Solution Ratio	Cheapest Quantity	Remarks
			Required Exps.	Optional Exps.			
Charcoal			2 lb.			5 lb.	powder
Charcoal			1 lb.			5 lb.	animal
Cheese-cloth				10 yd.			
Chlorid of lime	$CaCl(OCl)$		1 lb.	2 lb.		5 lb.	
Chloroform	$CHCl_3$	119.4		2 lb.		5 lb.	U. S. P.
Clay			5 lb.				
Cloth, woolen			1 yd.				nun's veiling
Cloth, cotton			5 yd.				muslin
Cobalt chlorid	$CoCl_2.6\ H_2O$		4 oz.		1 : 10	1 lb.	pure
Cobalt nitrate	$Co(NO_3)_2.6\ H_2O$	291.2		4 oz.	1 : 40	1 lb.	C. P.
Cochineal			4 oz.				
Cocoanut oil			1 lb.			5 lb.	pure
Congo red			1 oz.				
Copper, foil	Cu	63.6	4 oz.				
Copper, turnings			2 lb.	1 lb.			or copper shot
Copper nitrate	$Cu(NO_3)_2.\ 6\ H_2O$	295.8		4 oz.		1 lb.	C. P.
Copper oxid	CuO	79.6	8 oz.	8 oz.		1 lb.	C. P. powder
Copper sulfate	$CuSO_4.5\ H_2O$	249.7	2 lb.	1 lb.		5 lb.	crystals, pure
Cream of tartar	$KHC_4H_4O_6$	188.1	8 oz.			5 lb.	
Damar, gum				4 oz.			
Developer			2 tubes				M. Q.
Ether	$(C_2H_5)_2O$			2 lb.		5 lb.	U. S. P.
Extraction shell				6			S. and S.
Ferric chlorid	Fe_2Cl_6	324.7		1 lb.	1 : 10	1 lb.	C. P.
Ferric sulfate	$Fe_2(SO_4)_3.9\ H_2O$	562.4		1 lb.			
Ferrous sulfate	$FeSO_4.7\ H_2O$	278.2		1 lb.	*	1 lb.	C. P.
Ferrous sulfate			1 lb.			5 lb.	pure
Filters			10 pkg.				9 cm. qualitative
Filters				1 pkg.			ashless
Fuchsin				1 oz.			
Gasoline	C_6H_{14} and C_7H_{16}		1 qt.	1 qt.		1 gal.	
Gelatin			4 oz.			1 lb.	
Glue			4 oz.			1 lb.	
Glycerin	$C_3H_5(OH)_3$	92.1	1 lb.			5 lb.	U. S. P.
Gold ore				1 lb.		10 lb.	
Gypsum	$CaSO_4.2\ H_2O$	172.1		1 lb.		5 lb.	native lump
Hydrogen peroxid	H_2O_2	34.01	1 lb.			1 lb.	3%
Indigo			1 oz.	4 oz.			paste
Iodin	I_2	126.85	2 oz.	2 oz.			resublimed
Iron, nails			2 lb.				brads
Iron, powder			1 lb.			1 lb.	by alcohol
Iron, wire			1 lb.				#18
Iron, picture wire			8 oz.			1 lb.	
Iron alum	$Fe_2(SO_4)_3(NH_4)_2SO_4.\ 24\ H_2O$	392.2		8 oz.		1 lb.	
Iron oxid	Fe_2O_3	160	8 oz.			1 lb.	
Iron sulfid	FeS	88.1	2 lb.			5 lb.	granular
Kerosene	$C_{10}H_{22} - {}_{16}H_{34}$			1 gal.		1 gal.	
Labels							assorted
Lactose	$C_{12}H_{22}O_{11}.12\ H_2O$	360.2	1 lb.			5 lb.	milk sugar
Lead acetate	$Pb(C_2H_3O_2)_2.3\ H_2O$	379.2	1 lb.	1 lb.	1 : 10	5 lb.	U. S. P.
Lead acetate					189.5 per L.		
Lead nitrate	$Pb(NO_3)_2$	331	8 oz.			5 lb.	C. P.

Name	Formula	Mol. Wt.	Amount for a Class of 20–25		Solution Ratio	Cheapest Quantity	Remarks
			Required Exps.	Optional Exps.			
Lead monoxid	PbO	223	1 lb.			5 lb.	tech. litharge
Lead dioxid	PbO_2	239	8 oz.			5 lb.	tech.
Lead tetroxid	Pb_3O_4	685	8 oz.			5 lb.	tech. minium
Litmus			1 oz.		*		pure soluble
Logwood				4 oz.	*	1 lb.	
Magnesia mixture					*		
Magnesium, metal	Mg	24.4	2 oz.				ribbon
Magnesium, powd.			4 oz.				
Magnesium oxid	MgO	40.4	8 oz.			5 lb.	
Magnesium sulfate	$MgSO_4.7\ H_2O$	246.6	1 lb.			1 lb.	C. P.
Malachite green			1 oz.				
Manganese dioxid	MnO_2	87	8 oz.			1 lb.	C. P.
Manganese dioxid			1 lb.			5 lb.	tech, granular
Manganese dioxid			2 lb.			5 lb.	tech. powder
Manganese sulfate	$MnSO_4.4\ H_2O$	223		4 oz.		1 lb.	C. P.
Mercuric chlorid	$HgCl_2$	271.2		8 oz.	1 : 20	1 lb.	C. P.
Mercuric sulfid	HgS	232.4		8 oz.		1 lb.	cinnabar
Mercurous nitrate	$Hg_2(NO_3)_2$	524.7		8 oz.		1 lb.	C. P.
Methyl orange			2 oz.		*		
Methyl violet			1 oz.		1 : 1000		
Methylene blue			1 oz.		1 : 1000		
Molasses			2 qt.				Karo sirup
Naphtha				1 qt.		1 gal.	benzine
Naphthalene	$C_{10}H_8$	128.1	1 lb.			5 lb.	moth balls
Naphthylamine-α	$C_{10}H_7NH_2$	143.1	1 oz.		*		
Nickel sulfate	$NiSO_4.7\ H_2O$	280.9	4 oz.			1 lb.	C. P.
Nitrobenzene	$C_6H_5NO_2$	123.1		8 oz.		1 lb.	
Oxone				8 oz.		1 lb.	fused copper sulfate and sodium peroxid
Paraffin			2 lb.	1 lb.		5 lb.	
Phenolphthalein			1 oz.		*		
Phosphorus, red	P	31	8 oz.			1 lb.	
Phosphorus, white	P	31	4 oz.			1 lb.	
Phosphorus sesquisulfid	P_4S_3	220.2		8 oz.			
Plaster of Paris	$(CaSO_4)_2.H_2O$	290.2	1 lb.			5 lb.	
Potassium metal	K	39.15	1 oz.				
Potassium bromid	KBr	119.1	8 oz.			1 lb.	U. S. P.
Potassium carbonate	K_2CO_3	138.3	1 lb.			5 lb.	pure cryst.
Potassium chlorate	$KClO_3$	122.6	2 lb.			5 lb.	pure cryst.
Potassium chlorate			1 lb.			1 lb.	C. P.
Potassium dichromate	$K_2Cr_2O_7$	294.5	1 lb.	1 lb.		5 lb.	cryst.
Potassium ferrocyanid	$K_4Fe(CN)_6.3\ H_2O$	422.9		8 oz.	3 : 100	1 lb.	C. P.
Potassium ferricyanid	$K_3Fe(CN)_6$			8 oz.	3 : 100	1 lb.	C. P.
Potassium hydroxid	KOH	56.2	2 lb.	3 lb.		5 lb.	pure sticks
Potassium iodid	KI	166	8 oz.			1 lb.	U. S. P.
Potassium nitrate	KNO_3	101.2		1 lb.		5 lb.	pure
Potassium nitrite	KNO_2	85.2		8 oz.		1 lb.	C. P.
Potassium permanganate	$KMnO_4$	1 lb.	1 lb.	1 lb.		1 lb.	pure
Potassium sulfate	K_2SO_4	174.4	1 lb.			5 lb.	cryst.
Potassium sulfate				1 lb.		1 lb.	C. P.
Potassium sulfocyanate	$KCNS$	97.3		8 oz.		1 lb.	C. P.
Primulin				4 oz.			
Pumice		1	5 lb.			5 lb.	lump

Name	Formula	Mol. Wt.	Amount for a Class of 20–25		Solution Ratio	Cheapest Quantity	Remarks
			Required Exps.	Optional Exps.			
Resorcin	$C_6H_6O_2$	110.1	1 oz.				
Rochelle salt	$KNaC_4H_4O_6.4\ H_2O$	282.3		1 lb.		1 lb.	C. P.
Rock candy	$C_{12}H_{22}O_{11}$		1 lb.			5 lb.	crystals
Rosin			1 lb.			5 lb.	
Sand			2 lb.				clean
Seltzer water			1 qt.				
Shellac			1 lb.			1 lb.	white
Silver nitrate	$AgNO_3$	170	4 oz.	4 oz.			C. P.
Soap, castile			1 lb.				
Sodium, metal	Na	23.05	8 oz.				
Sodium acetate	$NaC_2H_3O_2.3\ H_2O$	136.7	1 lb.			1 lb.	fused
Sodium acetate					sat. sol. & 10%		C. P.
Sodium arsenate	$Na_2HAsO_4.12\ H_2O$	402.4		1 lb.		1 lb.	C. P.
Sodium arsenite	Na_2HAsO_3	170.1		1 lb.		1 lb.	C. P.
Sodium bicarbonate	$NaHCO_3$	841	5 lb.			5 lb.	pure
Sodium carbonate	Na_2CO_3	106.1	5 lb.	5 lb.		5 lb.	pure
Sodium chlorid	$NaCl$	58.5	10 lb.			5 lb.	tech.
Sodium dichromate	$Na_2Cr_2O_7$	262		1 lb.		5 lb.	pure
Sodium hydroxid	$NaOH$	40.1	5 lb.	5 lb.		5 lb.	pure sticks
Sodium nitrate	$NaNO_3$	85.1	1 lb.			5 lb.	pure
Sodium nitrite	$NaNO_2$	69.1	1 lb.			1 lb.	C. P.
Sodium phosphate	$Na_2HPO_4.12\ H_2O$	358.4		1 lb.		5 lb.	tech.
Sodium silicate	Na_2SiO_3	122.5	5 lb.			5 lb.	water glass
Sodium sulfate	$Na_2SO_4.10\ H_2O$	322.4	5 lb.			5 lb.	Glauber's salt
Sodium thiosulfate	$Na_2S_2O_3.5\ H_2O$	248.3	5 lb.			5 lb.	"Hypo"
Sodium bisulfite	$NaHSO_3$	104.07	1 lb.			5 lb.	
Starch	$(C_6H_{10}O_5)_n$		1 lb.			1 lb.	
Starch	Various kinds 1 oz. each						
Strontium, nitrate	$Sr(NO_3)_2$	211.6	1 lb.			1 lb.	tech.
Sugar, brown			1 lb.				
Sugar, cane	$C_{12}H_{22}O_{11}$	342.2	1 lb.	4 lb.		5 lb.	
Sugar, glucose	$C_6H_{12}O_6$	180.1		5 lb.		5 lb.	
Sulfur, flowers			5 lb.			5 lb.	
Sulfur, roll			10 lb.	10 lb.		5 lb.	brimstone
Tartar emetic	$KSbOC_4H_4O_6$			1 lb.		1 lb.	pure
Thionine				1 oz.			
Thread			1 spool				
Tin			1 lb.			1 lb.	mossy
Tin chlorid	$SnCl_2.2\ H_2O$	225.4	1 lb.			1 lb.	C. P.
Tin chlorid	$SnCl_4$	260.3	1 lb.			1 lb.	C. P.
Tin oxid	SnO	134.5	1 lb.			1 lb.	
Toluene diamid				1 oz.			
Turpentine	$C_{10}H_{16}$	136.2	1 lb.	1 lb.		5 lb.	
Wool blue				1 oz.			
Wool, yarn			1 lb.	1 lb.			white
Zinc, metal	Zn	65.4	10 lb.	10 lb.		5 lb.	mossy
Zinc, sheet			2 sq. ft.				
Zinc, dust			5 lb.			5 lb.	
Zinc chlorid	$ZnCl_2$	136.3		1 lb.		1 lb.	
Zinc oxid	ZnO	81.4	1 lb.			1 lb.	
Zinc sulfate	$ZnSO_4.7\ H_2O$	287.6	1 lb.			1 lb.	pure
Zinc sulfate				1 lb.		1 lb.	C. P.

SPECIAL SOLUTIONS

1. Ammonium Acetate.

Add 500 cc. of NH_4OH (sp. gr. 0.90) to 450 cc. of glacial acetic acid and neutralize the mixture with more of either reagent as the case requires. Dilute the mixture to sp. gr. 1.092 for a 50% solution, and to 1.022 for a 10% solution.

2. Ammonium Carbonate.

Add 250 g. of ammonium carbonate and 100 cc. of NH_4OH (sp. gr. 0.90) to 1 L. of distilled water. Keep the bottle closed

3. Ammonium Molybdate.

50 g. of ammonium molybdate, dissolved in 200 cc. of hot distilled water, are mixed with a 250 cc. solution of equal parts by volume of distilled water and concentrated HNO_3. Add 25 g. of NH_4NO_3. On standing a yellow precipitate of ammonium phosphomolybdate may deposit if pure materials are not used.

4. Ammonium Persulfid (poly).

Mix 200 cc. of concentrated commercial H_2SO_4 with a liter of water and about 360 g. of FeS. Pass the evolved gas into 1 liter of NH_4OH (sp. gr. 0.96). Add 25 g. of flowers of sulfur to the latter.

5. Chromic Acid Cleaning Solution.

Dissolve 25 g. of commercial sodium dichromate in 150 cc. of water, then add 100 cc. of concentrated commercial sulfuric acid. The same solution can be used repeatedly.

6. Cobalt Chlorid Test Paper.

Dissolve 20 g. of $CoCl_2.6\,H_2O$ in about 200 cc. of water. Wet filter paper with the solution and dry it. Cut into strips. Dry it over a flame before using it.

7. Ferrous Sulfate.

Dissolve 200 g. of $FeSO_4.7\,H_2O$ with 100 g. of $(NH_4)_2SO_4$ in 500 cc. of water with 50 cc. of concentrated H_2SO_4. Solution will keep for several months.

8. Halphen's Reagent.

Dissolve 2 g. of sulfur flowers in 125 cc. of carbon disulfid and mix with 125 cc. of amyl alcohol.

9. Haine's Solution, Test for Glucose.

Dissolve 10 g. of $CuSO_4.5\,H_2O$ in 875 cc. of water and add 45 g. of KOH sticks. Add 100 cc. of glycerin. This is a single solution reagent and will keep for a year.

10. Iodin Solution.

Dissolve 5 g. of iodin crystals and 10 g. of KI in 235 cc. of water.

11. Standard Lead Subacetate Solution.

Boil 430 g. of lead acetate, 130 g. of lead oxid (PbO) in 1 liter of water for 30 minutes. Filter the mixture, cool, and dilute it to 1.25 sp. gr.

12. Litmus Solution Indicator.

Powdered litmus should first be heated in alcohol to the boiling point. Filter the undissolved solid from the liquid and allow it to remain for several hours in cold water to remove alkaline impurities. Finally boil the solid residue with about 5 times its weight of water to make the solution for use. Preserve the solution by adding a little chloroform or thymol.

13. Logwood.

Dissolve logwood extract in hot water to make a solution having quite a dark color.

14. Low's Reagent.

Mix 4 volumes of glacial acetic acid with 1 volume of concentrated sulfuric acid.

15. Magnesia Mixture.

Dissolve 90 g. of $MgCl_2.6\,H_2O$ and 240 g. of NH_4Cl in 1 liter of water, then add 50 cc. of NH_4OH of sp. gr. 0.90.

16. Methyl Orange Indicator.

Mix 0.4 g. of methyl orange powder with 30 cc. of 95% ethyl alcohol and 170 cc. of water.

17. Millon's Reagents.

Dissolve mercury in its weight of concentrated nitric acid and dilute with an equal volume of water. Should be prepared as wanted.

18. Nessler's Reagent.

Add, slowly and with constant stirring, a 5% solution of $HgCl_2$ to a 10% solution of KI until the last drop leaves a slight permanent precipitate of HgI_2. Add to the mixture an equal volume of 10% KOH and filter. Keep in rubber-stoppered bottles.

19. Nickel Hydroxid Solution.

Dissolve 5 g. of nickel sulfate in 100 cc. of water and add a solution of NaOH until all the nickel is precipitated as hydroxid. Wash it well and dissolve it in 25 cc. of concentrated NH_4OH and 25 cc. of water. This solution dissolves silk at once, and reduces the weight of vegetable fibers only $\frac{1}{2}$% and of wool only $\frac{1}{3}$%.

20. Test for Nitrites.

Dissolve 0.5 g. of sulfanilic acid in 150 cc. of dilute $HC_2H_3O_2$. Boil 0.2 g. of solid α-naphthylamine with 20 cc. of water and pour off the colorless solution (throwing away the residue) and add 150 cc. of dilute $HC_2H_3O_2$. Combine the two solutions. If the solution becomes reddish, slake it with zinc dust and filter.

21. Phenolphthalein Indicator.

Dissolve 0.4 g. of phenolphthalein in 120 cc. of 95% ethyl alcohol and add 80 cc. of distilled water.

McDesm

22. Schweitzer's Reagent.

Precipitate hydrated copper oxid from a 2% solution of $CuSO_4$ by a slight excess of NaOH solution. Wash the precipitate free from alkali with water containing a little glycerin. This precipitate can be preserved unchanged in a 10% solution of glycerin in stoppered bottles. To use wash the oxid free from the glycerin and dissolve in 20% ammonia water.

23. Turmeric Paper.

Digest turmeric with several changes of water and filter. Discard the filtrate. Dry the turmeric and digest it with about six times its weight of ethyl alcohol. Dip filter paper in the solution and dry it. Cut into strips for use.

24. Acid Proof Wood Stain.

I. Dissolve 125 g. each of copper sulfate and potassium chlorate in the same liter of water.

II. 150 g. of anilin oil mixed with 180 g. of concentrated hydrochloric acid and a liter of water.

Apply 2 coats of the boiling hot solution I with a brush. Allow each coat to dry. Apply 2 coats of solution II in the same manner. When the wood is dry, wash with hot soapsuds. Finish with raw linseed oil or hot liquid paraffin, and refinish when the tables become dingy.

GENERAL RULES FOR SOLUBILITY

Certain generalizations can be made concerning compounds shown in the table on page 241. The exceptions to these generalizations are few and unimportant.

1. All **sodium**, **potassium**, and **ammonium** compounds are **soluble** in water.
2. All **nitrates**, **chlorates**, and **acetates** are **soluble** in water.
3. All **chlorids** are **soluble**, except those of silver, mercury (mercurous), and lead (lead slightly soluble).
4. All **sulfates** are **soluble**, except those of barium, lead, and calcium (calcium slightly soluble). The silver and the mercurous sulfates are only moderately soluble.
5. All **carbonates** are **insoluble**, except those of sodium, potassium, and ammonium.
6. All **oxids** and **hydroxids** are **insoluble**, except those of ammonium, sodium, potassium, and barium; calcium hydroxid is slightly soluble.

ACTIVITY SERIES (ELECTRO-CHEMICAL SERIES)

Metals		Non-Metals
Cæsium	**Hydrogen**	Silicon
Rubidium	Arsenic	Carbon
Potassium	Copper	Boron
Sodium	Antimony	Nitrogen
Lithium	Bismuth	Selenium
Barium	Mercury	Phosphorus
Strontium	Silver	Sulfur
Calcium	Palladium	Iodin
Magnesium	Platinum	Bromin
Aluminum	Gold	Chlorin
Manganese	Iridium	Oxygen
Zinc	Rhodium	Fluorin
Chromium	Osmium	
Cadmium		
Iron		
Cobalt		
Nickel		
Tin		
Lead		