SCIENCE TEXT-BOOKS

# EXPERIMENTAL CHEMISTRY

NEWELL

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# EXPERIMENTAL CHEMISTRY

BY

LYMAN C. NEWELL, PH.D. (JOHNS HOPKINS)

PROFESSOR OF CHEMISTRY, BOSTON UNIVERSITY,
BOSTON, MASS.

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# PREFACE.

THE purpose of this book is to promote the more efficient teaching of chemistry by modern methods. The choice and arrangement of subject-matter is based on the author's extended experience with students of varied ability. The book as a whole is the outcome of a desire to provide a course in chemistry which shall be a judicious combination of the inductive and deductive methods.

The concise directions and specific questions in the experiments aim to obviate the difficulties and perplexities which hinder the progress of beginners in chemistry. The experiments are representative. No important ones have been omitted, and no senseless or dangerous ones have been included. Every experiment need not be done by each pupil. Indeed, many experiments may be advantageously performed by the teacher either on the lecture table or in the laboratory; others may be done by the more skilful students and the results utilized in the class-room. Abundant material has been purposely provided to meet varied conditions. The TEACHER'S SUPPLEMENT suggests different uses of many experiments, and the author invites correspondence regarding the adaptation of the book to special demands.

Two supplementary sets of questions arise from most completed experiments. One class may and often must be answered in the laboratory where necessary data are available. The LABORATORY EXERCISES aim to meet this demand. The second class of questions concerns the vast amount of descriptive and historical matter indirectly connected with the experiments; this material is needed for a broad knowledge of the subject, but its presence would often destroy the real value of the experiments by precluding original thought on the part of the student. Hence, at appropriate points, CLASS-ROOM EXERCISES are inserted. They are not to be considered in the laboratory, but may be used as the basis of lectures, recitations, examinations, quizzes, or reviews. Both kinds of exercises may be omitted in whole or in part, if the teacher so desires. The BIBLIOGRAPHY contains titles of books helpful for a discussion of these exercises, and additional information may be found in the Teacher's Supplement. Despite the variety and amount of matter presented in these exercises, the ordinary text-books will meet the student's needs in most cases except those involving the recent discoveries and applications. A judicious selection and distribution of the points actually needed will obviate the immediate necessity of an extensive library. Facts not accessible to students may be easily supplied by the teacher. It should not be forgotten, however, that reference books are indispensable adjuncts to laboratory work.

Opinions differ widely regarding the kind of experiments which should constitute a course in chemistry. There is a growing belief, however, that selected experiments involving accurate measurements possess fundamental disciplinary value. This book contains several quantitative experiments, most of which have been repeatedly performed in the author's classes under a variety of conditions. The details have been worked out with the greatest care. It will be noticed that each quantitative experiment illustrates one or more of the fundamental principles of chemistry, hence the results permit conclusions which furnish a sound and attractive basis for the treatment of chemical theory. Teachers are earnestly urged to give these experiments a trial and to utilize the results in class work.

An unusually large number of problems has been intentionally incorporated to allow a generous distribution among several classes or sections. It is not intended that each student shall solve all the problems. The author will be grateful for information of errors or ambiguity detected in these problems. Only those who have prepared similar work for publication can realize the countless sources of error. The Teacher's Supplement contains answers to all problems.

The Appendixes contain directions for manipulation which could not be conveniently incorporated in the text, a complete treatment of the subject of gases, and miscellaneous data not easily found elsewhere.

Material intended exclusively for teachers has been put in a Teacher's Supplement. It is believed that the contents and the manner of its presentation will not only lessen their labors, but guide them in the choice of experiments, apparatus, etc. Copies can be obtained only by ordering directly of the publishers.

The entire manuscript has been read by —

Mr. Fred L. Bardwell, Assistant Professor of Chemistry, Institute of Technology, Boston, Mass.

Dr. William B. Shober, Instructor in Organic Chemistry, Lehigh University, Bethlehem, Pa.

Mr. Irving O. Palmer, Instructor in Chemistry and Physics, High School, Newton, Mass.

The proof has been read by —

Dr. Edward C. Franklin, Professor of Chemistry, University of Kansas, Lawrence, Kansas.

Mr. F. F. Coburn, Principal of the State Normal School, Lowell, Mass.

Dr. Charles A. Pitkin, Professor of Chemistry, Thayer Academy, South Braintree, Mass., and Professor of General Chemistry, Tufts College Medical School, Boston, Mass.

Mr. William H. Snyder, Master of Science, Worcester Academy, Worcester, Mass.

Dr. A. M. Muckenfuss, Professor of Chemistry, Millsaps College, Jackson, Miss.

Dr. H. G. Shaw, Instructor in Chemistry, High School, Melrose, Mass.

Dr. A. J. Hopkins, Assistant Professor of Chemistry, Amherst College, Amherst, Mass.

Dr. J. E. Bucher, Professor of Chemistry, College of Agriculture and Mechanic Arts, Kingston, R. I.

Dr. C. E. Boynton, Instructor in Chemistry, North Division High School, Chicago, Ill.

Dr. M. D. Sohon, Instructor in Physics, Mixed High School, New York.

Dr. Milo S. Walker, Instructor in Chemistry, West Division High School, Chicago, Ill.

Mr. Ernest A. Congdon, Professor of Chemistry, Drexel Institute, Philadelphia, Pa.

Mr. George W. Benton, Instructor in Chemistry, High School, Indianapolis, Ind.

Dr. J. Elliott Gilpin, Associate in Chemistry, Johns Hopkins University, Baltimore, Md.

The chapter on "Some Common Organic Compounds" was read in manuscript by Dr. James F. Norris, Instructor in Organic Chemistry, Institute of Technology, Boston, Mass.

The author is profoundly grateful to these teachers for their helpful and intelligent criticism, and he gladly takes this opportunity to express his appreciation of their kindness.

Numerous courtesies have been received from the L. E. Knott Apparatus Co., and for these the author hereby tenders his thanks.

To his former pupils and colleagues in the English High School at Somerville, Mass., the author is under obligation for generous assistance.

Suggestions for a short course may be found on page 398.

It is believed that the intelligent and careful examination of both manuscript and proof by the above teachers has produced a book relatively free from errors; nevertheless, for such errors as may be found the author willingly assumes responsibility.

L. C. N.

LOWELL, MASS.,
June, 1900.

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# CHEMISTRY.

## INTRODUCTION.

Much of the apparatus used in this book can be constructed in part by the student. The directions for such work here given are explicit, and the student is earnestly advised not only to acquire at the outset reasonable skill in glass working and other operations performed in the laboratory, but also to make constant use of the hints and directions concerning general manipulation. More detailed directions may be found in Appendix A.

### BUNSEN BURNER.

**The Bunsen Burner** is used as a source of heat in most chemical laboratories. A common form of this burner is shown in Fig. 1. It is attached to the gas cock by a piece of rubber tubing. When the gas is turned on, the current of gas draws air through the holes at the bottom of the tube, and this mixture, when lighted, burns with an almost colorless, *i.e.* non-luminous, flame. It is a hot flame and deposits no soot. The burner should be lighted by turning on a full current of gas and holding a lighted match in the gas about 5 centimeters (two inches) above

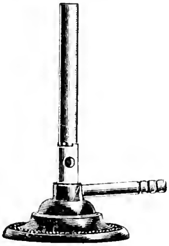

Fig. 1.— Bunsen burner.

the top of the burner; the height of the flame should then be regulated to meet the special demand.

The air supply may be reduced or entirely cut off by turning the ring at the bottom of the burner so that the holes in the tube are partly or completely closed. As the air supply is lessened, the flame gradually becomes yellow, and finally is luminous, like an ordinary gas flame. The luminous flame deposits soot and is not hot enough for most laboratory work. The non-luminous flame, which is often called "the Bunsen flame," or simply "the flame," should be used in all experiments, unless the directions otherwise state.

The parts of a Bunsen burner are shown in Fig. 68. The theory of the burner and structure of the flame are considered in Chapter XIII.

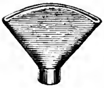

FIG. 2.—Wing-top burner.

A wing-top burner is shown in Fig. 2. This movable attachment slips over the top of the tube (see Figs. 4, 5, 7) and thereby produces a flat flame, which is used to heat through a longer space than the Bunsen flame permits, *e.g.* in heating glass tubing which is to be bent.

## GLASS WORKING.

**Glass Tubing.** — Soft glass tubing is used for all purposes except those in which intense heat is employed. It melts easily, and can be bent and blown into numerous shapes.

**Cutting.** — Glass tubes and rods are cut as follows: Determine the length needed, lay the glass on the table, and with a forward stroke of a triangular file make a short but deep scratch where the glass is to be cut. Grasp the tube in both hands, one on each side of the mark, and hold the thumbs together behind the scratch. Now push gently with the thumbs, pull at the same time with the hands, and the tube will break at the desired point. Hard glass tubing is cut in the same manner as soft tubing, though the scratch must be deeper. See Fig. 3.

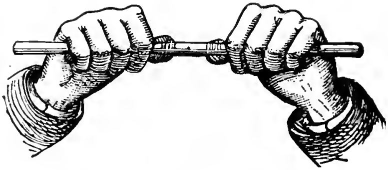

FIG. 3. — Cutting a glass tube.

The sharp ends should be rounded by rotating the end of the glass slowly in the flame until a yellow color is distinctly seen. See Fig. 4.

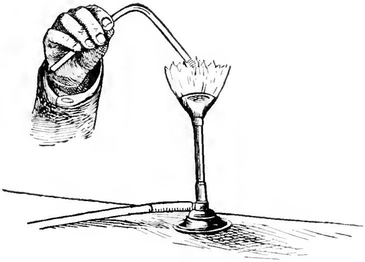

FIG. 4. — Rounding the sharp ends of a glass tube — "fire-polishing."

Certain precautions must be observed in cutting glass.

(1) If the glass does not break with *slight* pressure, do not push hard, but make a deeper scratch and try again.

(2) If the tube is moderately large or of hard glass, wrap it in a piece of cloth before attempting to break it. Never try to break large tubes by this method. Ask the teacher for directions.

**Bending — General Directions.** — (1) Use a wing-top burner, which produces a flat flame.

(2) Use a slightly yellow flat Bunsen flame for thin-walled tubes, and the colorless flame for thick-walled, or hard tubes.

(3) Never put a glass tube into the flame suddenly, nor a hot tube on a cold surface, lest it crack.

(4) Heat evenly, and cool gradually. All bends when made should be held at the top of the large yellow flame until covered with soot. This deposit protects the outside from the cooler air, and thus prevents sudden or unequal cooling. This operation is called "annealing."

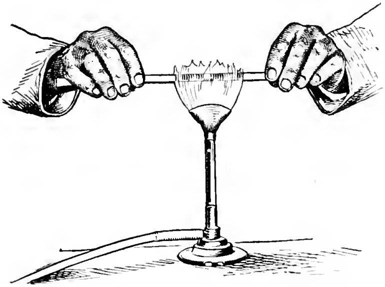

FIG. 5.—Bending a tube into a right angle—I.

(5) If two or more bends are made in the same tube, they should all lie in the same plane, *i.e.* all parts of the tube should touch a level surface.

**Operation 1. To bend a tube into a right angle.**—Hold the tube in the flame as shown in Fig. 5, and slowly rotate it between the thumbs and forefingers in order to heat it evenly. When it feels soft and ready to yield, take it out of the flame and slowly bend it into a right angle, as shown in Fig. 6. Compare the angle with the corner of a square block of wood, or with any other right angled object, and if not correct, repeat the operation with another tube.

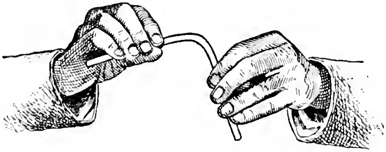

FIG. 6.—Bending a tube into a right angle—II.

**Operation 2. To bend a tube into an oblique angle.**—Hold the tube in the flame, as shown in Fig. 7. Rotate it, and at the same time move it slowly to the right and left so that from five to eight centimeters are evenly heated. When soft, remove it and bend into the desired angle, as shown in Fig. 8. It is helpful to have this angle sketched

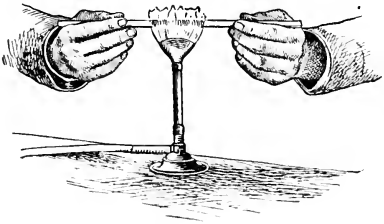

FIG. 7.—Bending a tube into an oblique angle—I.

roughly on a piece of paper near by, so that the tube may be held over it while being bent.

**Precautions.** — (1) Tubes to be bent at an oblique angle should be heated through about twice the space required for a right angle. A very slight bend, however, is often made by heating a short space.

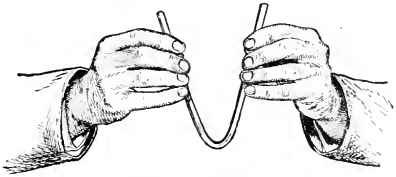

FIG. 8. — Bending a tube into an oblique angle — II.

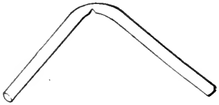

FIG. 9. — A worthless bend.

(2) The heat should be applied continuously, and the tube never bent at a low temperature, otherwise the curves may be flattened or creased, or may even collapse, as shown in Fig. 9. Such bends are brittle, and therefore worthless.

## CUTTING AND BENDING GLASS.

**Experiment 1.** — Cut into halves a glass tube about 30 centimeters long. Bend one piece into a right angle and the other into an oblique angle. Submit each to the teacher for criticism.

**Experiment 2.** — Bend a glass tube, about 35 centimeters long and 5 millimeters in external diameter, into a double right angle, so that each arm is about 10 centimeters long. See Fig. 10. Preserve this tube, as it is used in many experiments.

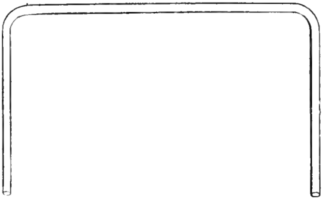

FIG. 10. — A double right angle bend.

**Experiment 3.** — Bend a glass tube, 40 to 45 centimeters long, into the shape shown in Fig. 11. The numbers indicate the approximate length of the different portions. The shorter arm is made by holding the tube at that point (7 centimeters from the end) at almost right angles to the flame, and then bending slowly, regulating the pressure by the cooling of the tube. Preserve this delivery tube for future use.

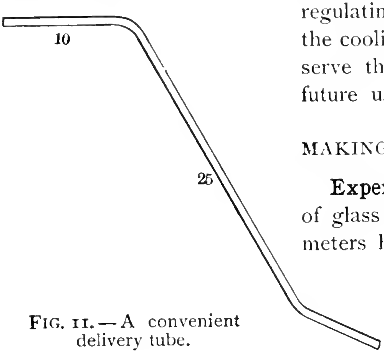

FIG. 11.—A convenient delivery tube.

## MAKING STIRRING RODS.

**Experiment 4.**—A piece of glass rod about 25 centimeters long and 5 millimeters in diameter is needed. Heat the rod in the middle in the ordinary—not flat—Bunsen flame, and when soft draw it out slightly into the shape shown in Fig. 12. Cut it into two rods by making a slight scratch at the desired point of the narrow portion. It is convenient to have one rod rather blunt for stirring hot metals and the other somewhat pointed for general use. All rough points should be melted off in the flame. A completed rod is shown in Fig. 13.

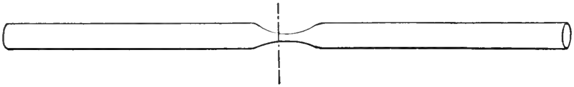

FIG. 12.—Stirring rod partially made.

## CLOSING TUBES.

**Experiment 5.**—Small tubes are closed, or sealed, by heating one end of the tube in the Bunsen flame; the tube should be constantly rotated to insure even heating. Large tubes are closed as described in Exp. 6. Close a small tube about 10 centimeters long, and submit it to the teacher for criticism.

Directions for working with hard glass tubing may be found in Appendix A.

FIG. 13.—A stirring rod—exact size.

### MAKING IGNITION TUBES. — OPTIONAL.

**Experiment 6.** — See APPENDIX A, § 1.

### CLASS-ROOM EXERCISE. I.

1. Why does thick glass crack when suddenly heated?
2. Why is glass brittle, if it has been suddenly cooled?
3. Why does a hot tube crack, if laid on a cold surface or if wet with water?

### HEATING.

Effective use of the Bunsen burner flame is learned only by experience, but a few preliminary hints may prove serviceable. The burner should always be lighted before any piece of apparatus is held over it, or before it is placed under the piece of wire gauze which supports a dish (see App. A, § 2) — sudden heat may crack the apparatus. A small flame is more easily controlled than a large one. Usually a flame from 5 to 10 centimeters (2 to 4 inches) high gives sufficient heat, and is not easily blown about by sudden drafts.

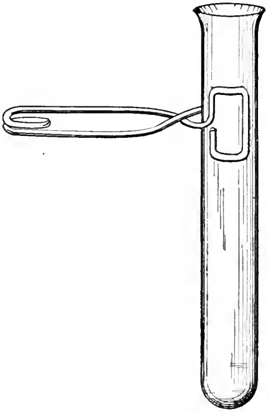
FIG. 14. — Test tube and holder.

Chemical glassware is made of uniformly thin glass, and if heated with certain precautions will not crack. A wet test tube should never be put in or over the bare flame. The temperature of dry test tubes, or of those containing only solids, should be raised gradually by moving them in and out of the flame or by holding them in the flame and rolling them slightly between the thumb and forefinger. If the test tube contains a solid, special care must be taken to distribute the heat evenly; if the test tube contains a liquid, it should be slightly inclined so that the greatest heat is not on the bottom, which is often thinner than the sides; the flame should never come above the surface of the liquid, or it will heat the tube and then the liquid touching this spot will cause the tube to crack. In most experiments test tubes can be held between the thumb and forefinger without discomfort. If they are too hot to handle, a test tube holder should be used. See Fig. 14.

## FILTRATION.

A finely divided solid suspended in a liquid may be separated from the liquid by filtration or filtering. A circular piece of porous paper (Fig. 15) is folded so that it exactly fits a funnel, and when the mixture is poured upon this paper, the solid — the *residue* or *precipitate* — is retained, while the liquid — the *filtrate* — passes through the paper. The paper is prepared for the funnel by folding it successively into the shapes shown in Figs. 16 and 17, and then opening as shown in Fig. 18, so that three thicknesses are on one side and one on the other. (See App. A, § 4.)

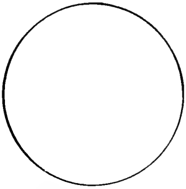

Fig. 15. — Piece of filter paper.

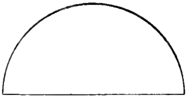

Fig. 16.—Folded paper — first stage.

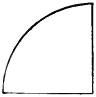

Fig. 17.—Folded paper — second stage.

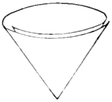

Fig. 18. — A filter paper folded ready for the funnel.

## MISCELLANEOUS SUGGESTIONS.

**Smelling and Tasting.** — A safe rule to follow is never to smell or taste unfamiliar substances, unless so directed, and even then with the utmost caution. Never inhale a gas vigorously, but waft it gently toward the nose. Taste by touching a minute portion of the substance to the tip of the tongue, and as soon as the sensation is detected, reject the solution at once — never swallow it.

**Cleanliness.** — Successful laboratory work is largely measured by general cleanliness. All apparatus should be clean before use, and should be washed as soon as it has been used. In quantitative work it is absolutely necessary to have balance, weights, crucibles, forceps, and every other piece of apparatus perfectly clean.

### WEIGHING AND MEASURING.

Exact relations by weight or volume are found by weighing or measuring, or by both. Such operations require a working knowledge of the system of weights and measures used in science and a familiarity with the chemical balance and other instruments of precision.

**The Metric System** of weights and measures is used in science. For a brief outline of the system, with tables, see Appendix A, § 18.

The relation of the metric system to the weights and measures in common use is shown by the following: —

Table of Equivalents.

| | | |
|---|---|---|
| * 1 meter | = | 39.37 inches |
| 1 liter | = | 1.056 quarts (liquid) |
| 1 liter | = | 0.908 quarts (dry) |
| 1 gram | = | 15.432 grains |
| * 1 kilogram | = | 2.2 pounds (avoir.) |
| * 1 inch | = | 2.54 centimeters |
| 1 mile | = | 1.6 kilometers |
| 1 cubic inch | = | 16.39 cubic centimeters |
| 1 quart (liquid) | = | 0.9465 liters |
| 1 pound (avoir.) | = | 0.4536 kilograms |

The equivalents marked with a * should be learned. Other equivalents can be easily deduced from the above. (See App. A, § 18, Table of Transformation.)

A convenient relation (true only in the case of water) to remember is 1 l. = 1 kg. = 1 cu. dm. = 1000 cc. = 1000 gm. = 2.2 lb.

The customary abbreviations of the most common denominations are as follows:—

Meter, m.
Decimeter, dm.
Centimeter, cm.
Liter, l.
Kilogram, kg. or Kg.
Decigram, dg.
Centigram, cg.
Milligram, mg.
Cubic centimeter, cc.

The preferable abbreviation for gram is gm. Occasionally cu. cm. is used for cc. The same abbreviation may be used for the plural and singular, *e.g.* 1 gm., 4 gm., .04 gm.

### PROBLEMS. I.

1. Express:—
   (*a*) 1 km. in millimeters.
   (*b*) 1 cm. in decimeters.
   (*c*) 127.5 cm. in inches.

2. Add 1 km., 2 m., 4 dm., 6 cm., 9 mm., and express the result in centimeters.

3. Subtract 1 mm. from 4 cm., and express the answer in centimeters.

4. How many tubes 1.5 cm. long can be made from 249 m. of tubing, allowing 1.5 m. for loss?

5. Express in cubic centimeters:—
   (*a*) 1 l. (*b*) 27 cu. dm. (*c*) 1721 l. (*d*) 24.8 l.

6. How many liters in a tank 75 m. long, 27 cm. deep, and 3 dm. wide?

7. What is the capacity in cubic centimeters of a box 2 m. long, 8 dm. wide, and 7 cm. deep?

8. How many grams in 1.647 kg.?

9. How many cubic centimeters in 721 l.?

10. Add 2 gm., 9 cg., 14 dg., and 237 mg., and express the sum in grams.

11. How many pounds in 25 kg.?

12. How many milligrams in 5 cc. of water at 4° C.?

13. The standard pressure at which a gas is measured is 760 mm. Express the same in inches.

14. How many millimeters in 1.575 m.? In 0.5 m.? In 0.005 m.?

15. How many meters in 100.056 dm.? In 1.5 dm.? In 50 dm.?

16. A box is 0.5 m. deep, 25 dm. long, and 1.2 m. wide. What is the volume in cubic centimeters?

17. If a cylinder holds 141.3 gm. of water, and its height is 20 cm., what is its diameter?

18. How many grams in 2000 dg.? In 1800 cg.?

19. How many centigrams in 21 gm.? In 1900 mg.? In 1565 dg.?

20. How many cubic centimeters in 100 l.? In 0.5 l.? In 15 cu. dm.? In 0.5 cu. m.?

21. How many grams of water in a cylinder 9 dm. high and 5 cm. in diameter?

22. How many millimeters, centimeters, and meters are respectively contained in 0.437 dm.?

23. How many grams in 1.725 kg.?

24. How many centigrams in 2.567 kg.?

25. A piece of platinum foil measuring 10.5 cm. by 1.5 cm. weighs 0.723 gm. Into how many pieces each weighing one decigram may it be divided?

26. A circular piece of filter paper is 10 cm. in diameter. What is its area?

27. A cistern is 2 m. long, 1.5 m. wide, and 1 m. deep. How many liters of water will it contain?

28. A cylindrical gas holder is 1 m. in diameter and 1.5 m. high. How many liters of gas will it hold?

29. A wire 255 mm. long weighs 0.172 gm. What length of this wire is necessary to make a rider weighing 1 cg.?

30. Sulphuric acid is 1.8 times heavier than water. How many grams of acid will a liter flask hold?

31. Alcohol is 0.8 as heavy as water. What is the weight of 1200 cc of alcohol?

32. Express in meters and in millimeters:—

(*a*) 6 in. (*b*) 1.25 ft. (*c*) 2.6 yd.

33. How many pints in 2.5 liters?

34. How many liters in 2 gal.?

35. How many grams in 150 grains?

36. How many pounds in 7 kg.?

**A balance** sufficiently accurate for the exact experiments of this book is shown in Fig. 19. It is often called a "horn pan balance," because the pans are made of that material. Some of the experiments give satisfactory results, if the ordinary trip or platform scales are used. See Fig. 20. All rough or approximate weighing should be done on the trip scales.

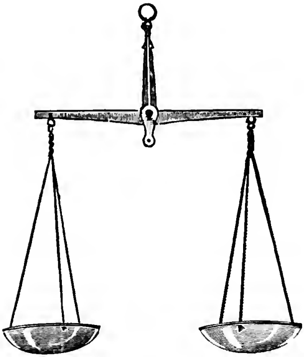

FIG. 19. — Horn pan balance.

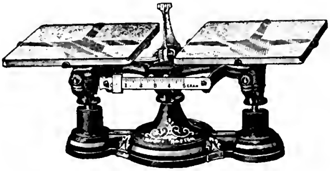

FIG. 20. — Trip or platform scales.

Both of the balances above mentioned must be counterpoised before use. Directions for counterpoising and for weighing may be found in Appendix A, §§ 19, 21.

### WEIGHING.

**Experiment 7.** — (*a*) Weigh a small object such as a bottle, stone, or piece of lead on the trip scales. Weigh to a decigram, and express the result in grams and a decimal fraction of a gram. Record the weight in the proper place below.

(*b*) Verify the result by weighing the same object on an accurate balance, if one is available; if one is not available, repeat the weighing on the same scales. Record the weight in the proper place below.

(*c*) Results: —

| | GRAMS. |
|---|---|
| Weight of object on trip scales . . . . . . | |
| Weight of object on balance . . . . . . . | |
| Error in weighing . . . . . . . . . . | |

**Form of Record of Experiments.** — No one form of record is adapted to every experiment, but all forms should include at least (*a*) a short description of the operation, (*b*) all weights and measurements, and (*c*) answers to all questions. The following will serve as a preliminary model: —

### WEIGHING.

**Experiment 7.** — A small bottle was weighed on the trip scales and then on a balance. The following results were obtained: —

| | GRAMS. |
|---|---|
| Weight of bottle on trip scales . . . . . . | |
| Weight of bottle on balance . . . . . . . | |
| Error . . . . . . . . . . . . . | |

Every record must, of course, include the title and number of the experiment. It is often helpful, also, to know the date when a particular experiment was performed. The notes, or record, should be written, as far as possible, during the performance of the experiment.

### WEIGHING AND MEASURING.

**Experiment 8.** — (*a*) Weigh a small, dry, clean, empty bottle on the trip scales. Record the weight at once in the note-book.

(*b*) Fill the bottle to the neck with water and weigh again. Be sure that the outside of the bottle, including the bottom, is dry, before standing it on the scale pan. Record the result as shown below.

(*c*) Transfer all the water to an empty graduated cylinder. Pour the water into the cylinder without spilling a drop. This can be done by pouring the water down a glass rod (see App. A, § 12), or by a preliminary movement quick enough to prevent the water from running down the outside of the bottle. Read the exact volume of water in the graduate. Record the result as shown below. What is the relation between the weight and volume of water?

**Form of Record of Experiment 8.** — The record of Exp. 8 should be somewhat as follows: —

## WEIGHING AND MEASURING.

**Experiment 8.** — A small bottle was weighed empty and then nearly full of water, and the volume of water measured.

Results: —

| | | GRAMS. |
|---|---|---|
| Weight of bottle and water . . . | | 115.2 |
| Weight of bottle, empty . . . . | | 65.4 |
| Weight of water . . . . . . . | | 49.8 |
| Volume of water . . . | 50 cc. | |
| Weight of volume of water . . . | | 50.0 |
| Error . . . | | 0.2 |

The weight and volume are nearly equal.

(date)

### PROBLEMS. II.

*Formulas for the conversion of thermometric readings will be found in Appendix A,* § 23.

1. Convert into Fahrenheit readings the following readings on the centigrade scale: —

(*a*) 60.5
(*b*) 40
(*c*) 70
(*d*) −5
(*e*) −36.6
(*f*) 643
(*g*) 720
(*h*) 860
(*i*) 973
(*j*) 1000
(*k*) 0
(*l*) 100

2. Convert into centigrade readings the following readings on the Fahrenheit scale: —

(*a*) 207
(*b*) 180
(*c*) 114
(*d*) 0
(*e*) −30
(*f*) 674
(*g*) 100
(*h*) 463
(*i*) 3478
(*j*) 2941
(*k*) −31
(*l*) −76

3. What is the difference in centigrade degrees between 72° C. and 72° F.?

4. What temperature is expressed by the same number on each scale?

# CHAPTER I.

## PHYSICAL AND CHEMICAL CHANGES.

THE matter of the universe is constantly changing. Sometimes the change temporarily modifies the special properties of the matter under examination, but often the change is permanent and another substance or kind of matter is the result. When the properties of a given portion of matter are so changed that a different kind of matter is formed, then the change is called a *chemical change*. If the properties are temporarily changed, then the substance has undergone a *physical change*. Most chemical changes are accompanied by physical changes, sometimes as causes, sometimes as effects. So closely related are they that it is often impossible to distinguish them, even for study.

**Chemistry** is largely a study of these chemical and physical changes, and, especially, their causes, products, and attendant phenomena.

**Preliminary Definitions.** — Many of the substances used in experimental chemistry have exceedingly simple properties, so simple that it is believed that these substances cannot be subdivided into simpler ones, and they are therefore placed in a class by themselves, called **elements**. Some are familiar and others will become familiar as the experiments proceed. The common **metals** are elements. Thus copper and zinc have properties which are so simple that chemists regard these metals as elements. Platinum and magnesium may not be familiar, but they resemble other metals in their luster, hardness, power to conduct heat and electricity, as well as in their chemical behavior toward other elements. All metals do not have the same properties, but the best known have an undeniable resemblance to the familiar shining substances called metals.

Elements contain no other substances. Zinc is only zinc, hence, if zinc changes into a substance having different properties, the change is due to something else than the zinc. This fact must be remembered in interpreting the experiments in this chapter. A broad knowledge of elements, whether metals or not, is obtained only by extensive study; hence the ideas gained in this preliminary investigation will be enlarged by additional experiments.

**Changes in Metals when Heated.** — A knowledge of some of the conditions accompanying chemical and physical changes is necessary as a foundation for the study of chemistry, especially of experimental chemistry. An introductory study should be simple and limited, and the present chapter will be confined to an examination of the changes which metals undergo when they are heated in the air.

### GENERAL CHANGES.

**Experiment 9.** — *Supplies:* Platinum wire, copper wire, magnesium ribbon, sheet zinc, forceps.

Examine successively a piece of platinum, copper, zinc, and magnesium carefully enough so that they could be detected with certainty, if seen again under similar conditions. Then take each successively in the forceps and hold it in the upper part of the Bunsen burner flame long enough to produce a definite change. If the change occurs suddenly, remove the metal and observe the character of the change. Try each metal until there is a definite result. Look for simple things, such as change of color, new substances, etc. In examining the zinc hold the burner at an angle, so that the melted zinc will not drop inside. When definite results have been obtained, answer the following: —

(1) Are all the metals changed?

(2) Are all changed permanently? Give reasons for the conclusion.

(3) Examine each product and state briefly how they differ, if they do, from the original metal.

(4) Since nothing in the metal assisted the heat in causing the change, what did?

Further study is necessary to verify any theory of the auxiliary cause of the changes in Exp. 9.

### EFFECT OF HEATING A METAL WHEN UNCOVERED AND WHEN COVERED.

**Experiment 10.** — *Supplies:* Porcelain crucible, pronged tripod or similar support, pointed glass rod (blunt form), lead, magnesium oxide.

(*a*) Put a small piece of lead in a porcelain crucible and stand the crucible on a tripod as shown in Fig. 21. If such a support is not available, use a triangle and ring of an iron stand. Heat the crucible for a few minutes with a low flame, — about 5 cm. high, — then gradually increase the heat until the bottom of the crucible is hot. Occasionally scrape aside the greenish product with the blunt glass stirring rod. The heat must not be sufficient to melt this product.

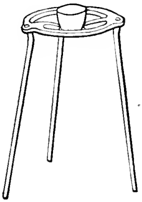

FIG. 21. — Crucible supported by a tripod.

Continue to heat until there is enough of the new substance to examine. While the lead is still molten, firmly grasp the crucible near the edge with the forceps, and pour out the contents of the crucible upon an iron pan, or a piece of asbestos board, or a block of wood. Compare the product with the unchanged lead, stating briefly the differences. How does the change in the lead resemble the changes in Exp. 9? Is it a similar chemical change? (If the crucible is not clean, heat it where the lead sticks, and scrape off the softened lead with a file.)

(*b*) In the same crucible put another piece of lead and cover it completely with magnesium oxide, pressing down the white, infusible substance with the finger. Heat the crucible as in (*a*), but do not disturb the contents. The lead should be completely covered during the whole operation. Heat until it is certain that the lead has been melted several minutes — long enough at least to produce the greenish powder, if any can form. Pour out the contents as before. Has any greenish powder formed? Why? What, then, assists heat in the chemical changes in Exp. 9 and Exp. 10 (*a*)? Does this conclusion verify your theory drawn from Exp. 9? If not, modify your notes accordingly.

The result of Exp. 10 gives a clew to the nature of the permanent changes which metals undergo when heated in the air, but it does not prove conclusively whether by the change the metals undergo a loss or a gain. This question is considered in the next experiment.

### RESULT OF HEATING A KNOWN WEIGHT OF A METAL IN THE AIR.

**Experiment 11.**—*Supplies:* Powdered iron, zinc dust, crucible and support, glass rod used in Exp. 10, crucible block.

(*a*) Clean and dry a porcelain crucible and weigh it to a decigram on the trip scales. Slide powdered iron into the crucible from a narrow strip of smooth paper creased in the middle, until about 3 gm. have been added. (See App. A, § 12.) Do not spill the iron on the scale pans. The iron should not coat the inside of the crucible, but be in a heap on the bottom. Wipe off any iron which is on the outside of the crucible or on the inside above the heap, before weighing the crucible and contents. The iron need not weigh exactly 3 gm., but the exact weight, whatever it is, must be known. Record the weights in the note-book, as soon as made, thus: —

| | GRAMS. |
|---|---|
| Weight of crucible . . . . . . . . . . . . . | 12.2 |
| Weight of iron and crucible . . . . . . . . . | 15.4 |
| Weight of iron . . . . . . . . . . . . . . | 3.2 |

In carrying a crucible to and from the scales, it should not be carried in the hand nor by the forceps, but placed in the crucible block which is shown in Fig. 22. Support the crucible as before and heat it for about ten minutes. Heat with a low flame at first and gradually increase the heat. Finally stand the burner under the crucible, taking care, however, that the flame is *below* the top of the crucible. Touch the metal occasionally with the pointed glass rod and if it is hard, pierce the crust cautiously in several places to expose a fresh surface. If iron clings to the rod, hold the pointed end over the crucible and tap the rod gently against the crucible so that the particles will fall back. Avoid losing anything from the crucible. After the crucible is cool enough to bear the hand without discomfort, weigh again. If the weight has changed, record thus: —

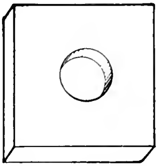
FIG. 22. — Crucible block.

| | GRAMS. |
|---|---|
| Weight of crucible and contents before heating . . | 15.4 |
| Weight of crucible and contents after heating . . | |
| Gain or loss in weight . . . . . . . . . | |

If there is no change in weight, heat again, and then weigh; if still no change, consult the teacher. What is proved by the change in weight?

(*b*) The general result obtained in (*a*) may be verified by heating zinc dust. Weigh the crucible and weigh in it about 3 gm. of zinc dust, observing the same precautions as in (*a*). Record thus: —

| | GRAMS. |
|---|---|
| Weight of crucible . . . . . . . . . . . . . . | 13.0 |
| Weight of zinc and crucible . . . . . . . . . | 15.9 |
| Weight of zinc . . . . . . . . . . . . . . . | 2.9 |

The zinc must be heated with the utmost care to avoid loss. Use a low flame, not more than 5 cm. high, during the entire heating. Do not stand the burner under the crucible. and if the zinc glows or gives off a white smoke, remove the burner at once. This smoke weighs something, and its loss means a loss in weight which cannot be estimated. If a crust forms on the zinc, break it as in (*a*). Avoid losing any zinc or any of the white product. Heat the crucible for about 10 minutes, then cool and weigh. If there is a change in weight, record thus: —

| | GRAMS. |
|---|---|
| Weight of crucible and contents before heating . . | 15.9 |
| Weight of crucible and contents after heating . . | |
| Gain or loss in weight . . . . . . . . . | |

If there is no change, heat again and then weigh; if still no change, consult the teacher. What is proved by the change in weight?

If magnesium, lead, copper, or mercury, in fact, almost any metal, is heated as in Exps. 10 and 11, similar results are produced. The nature of the interacting substance, as far as the above experiments show, is unknown, but whatever it is, it has combined with them in such a way as to change their properties. It forms chemical compounds of which the metal is one part and this "unknown substance" is another part. If the process could be reversed by decomposing one of these compounds and getting this "unknown substance," it could be studied, and its nature would throw some light on the real cause of the changes which metals undergo when heated in the air. None of the compounds — the products of Exps. 10 and 11 — is suited to such an experiment. Fortunately we have available the very compound used by the chemist who discovered this "unknown substance." It is a red powder and is formed by heating the metal mercury in the air, just as the white powder was formed from zinc, the black film from copper, and the greenish powder from lead. When heated, this red powder decomposes into a well-known metal and a gas which is the "unknown substance."

### PRODUCTS OBTAINED BY HEATING THE RED POWDER.

**Experiment 12.** — *Supplies:* Apparatus shown in Fig. 23, pneumatic trough, small bottle, iron stand and clamp, the red powder, splinter of soft wood, ignition tube.

The delivery tube for this experiment was made in Exp. 3. If the ignition tube is not available, it may be made according to the directions in Appendix A, § 1. If the ignition tube is not large, it may be attached to the delivery tube as shown in Fig. 23; if it is rather large, then it must be attached to the delivery tube by putting the rubber connector inside the ignition tube, the upper end of the delivery tube itself remaining, as before, within the connector. A cross section of such a connection is shown in Fig. 24.

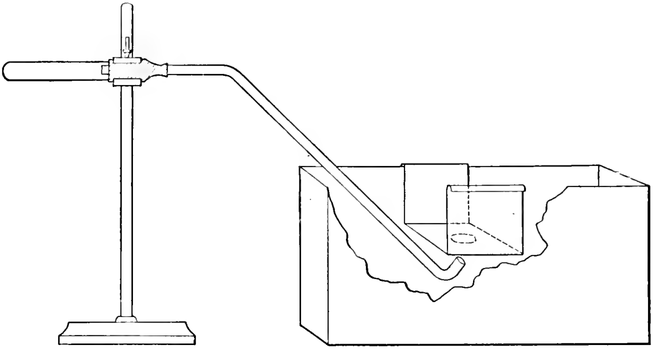

FIG. 23. — Apparatus for heating the red powder and collecting the gaseous product.

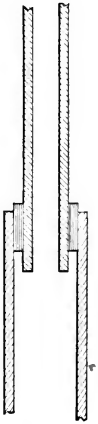

FIG. 24. — Section of an inside connector.

Slip a little of the red powder from a paper into the ignition tube and connect with the delivery tube. (See App. A, § 11.) Clamp the tube near the top so that the apparatus shall have about the position shown in Fig. 23. Tap the tube gently to spread out the powder in the tube. The outer end of the delivery tube reaches just below the shelf of a pneumatic trough arranged to collect the gas over water. Fill the bottle full of water, invert it, and stand it on the shelf near the hole. (See App. A, § 15.)

Heat the whole ignition tube with a low flame at first, gradually increasing the heat where the red powder is located. As the heat increases, bubbles pass up through the water. Reject the first few, which are air, then slip the bottle over the hole in the shelf. If the powder "crawls" up the ignition tube, tap the tube gently until the contents slips back. If the evolution of gas ceases, heat more strongly. Do not wave the burner back and forth, but apply the heat steadily by a low flame which envelopes that part of the tube containing the powder.

Avoid heating the end of the tube, as it is often so thin that it melts.

Collect the bottle at least half full of gas and remove it thus: Slip the bottle from the shelf, still keeping the mouth under water, and cover the mouth with a piece of moistened filter paper. Keep hold of the paper with one hand and with the other invert the bottle and stand it, still covered, on the desk. Remove the delivery tube *at once* from the trough to prevent the water from being drawn up into the hot apparatus.

The gas from the red powder is the "unknown substance" which was taken from the air by a certain metal when heated in the air. Examine the deposit on the inside of the ignition tube. If its nature is doubtful, pour it out on a block of wood. What is it? Test the gas thus: Light a splinter of wood, let it burn a few seconds, then blow it out and plunge the glowing end into the gas in the bottle. What happens? Remove the splinter, if there is any decided change, and once more plunge it, while glowing, into the gas. Answer the following: —

(1) Is the gas air? Why not?
(2) Where must the gas have come from originally?

A substance is usually identified by comparing its properties with those of some known substance with which it is suspected to be identical.

### IDENTIFICATION OF THE GAS OBTAINED FROM THE RED POWDER.

**Experiment 13.** — Obtain a bottle of oxygen from a gas holder and plunge a glowing splinter into it. Has oxygen the same characteristic property as the "gas obtained from the red powder"? What one conclusion can be drawn from this experiment?

**Definitions.** — An **Oxide** is a compound of oxygen and *one* other element, such as iron, zinc, or lead. **Chemical action** is a term applied to all classes of chemical changes, such as the addition of oxygen to iron, or the decomposition of the red powder into oxygen and mercury. The cause, accompaniments, and products of chemical action will be constantly studied.

### LABORATORY EXERCISE. I.

1. Summarize the results of Exps. 9 to 13 into a concise argument.
2. Name all the oxides studied or used in this chapter.
3. What is the chemical name of the red powder?

### CLASS-ROOM EXERCISE. V.

1. Discuss the states and properties of matter.
2. Physical and chemical changes.
   - (*a*) Give three illustrations of both occurring in every-day life.
   - (*b*) What is supposed to cause many chemical changes? Is change constant or intermittent?
   - (*c*) What aids and what retards chemical change?
3. Review the Metric System.
4. Manipulation.
   - (*a*) Precautions in heating hard glass and porcelain.
   - (*b*) Most effective use of Bunsen flame.
   - (*c*) Give the successive steps in collecting gases.
5. Give evidences of chemical action observed in Exps. 9 to 13.

# CHAPTER II.

## OXYGEN

THE experiments in Chapter I. teach that one important constituent of the air is oxygen. Subsequently, air will be examined with special reference to its composition. We shall now study oxygen, which is the most important of the few simple substances to which the matter of the universe may be reduced.

### PREPARATION AND PROPERTIES OF OXYGEN.

**Experiment** 14. — *Supplies:* Retort, ring and stand, pneumatic trough, six bottles (about 250 cc. capacity), filter paper, splinter of soft wood, deflagrating spoon, roll sulphur, piece of charcoal fastened to a wire, short piece of wire picture cord unwound at one end, magnesium ribbon, 15 gm. of crystallized potassium chlorate, and 15 gm. of coarsely powdered manganese dioxide.

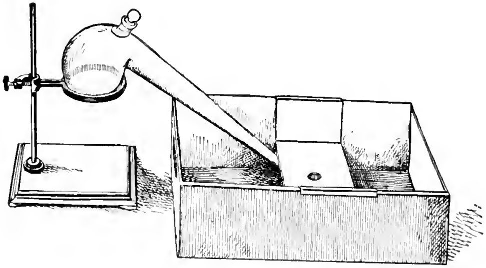

FIG. 25. — Apparatus for the preparation of oxygen.

Procure all the supplies before attempting to prepare the oxygen. Arrange the apparatus as shown in Fig. 25, and collect the gas over water. (See App. A, § 15.) The potassium chlorate and manganese dioxide should be free from pieces of cork, paper, straw, or wood; the retort must be dry and clean and provided with a tightly fitting glass stopper.

Mix the potassium chlorate and manganese dioxide on a piece of smooth paper, and pour the mixture into the retort as shown in Fig. 101. Invert the bottles full of water in the trough, and have the filter paper where it can be conveniently reached. Adjust the apparatus, and heat the bare retort carefully with a low flame. Direct the heat upon the part of the retort containing the mixture. Do not collect the gas before it bubbles freely through the water. If the gas is evolved too rapidly, lessen the heat; if not at all, then examine the stopper for a leak. Collect six bottles of gas, covering each when full with a piece of wet filter paper, as soon as taken from the trough.

When the gas has been collected, remove the neck of the retort immediately from the water, lest the cold water be drawn up into the hot bulb of the retort, as the gas contracts.

Proceed at once to study the properties of oxygen by the following experiments: —

(*a*) Thrust a glowing splinter of wood into one bottle. Remove the splinter, if there is any change, and repeat as many times as possible. Describe the result. What property of oxygen does this experiment show? What is the essential difference observed between burning in air and in oxygen?

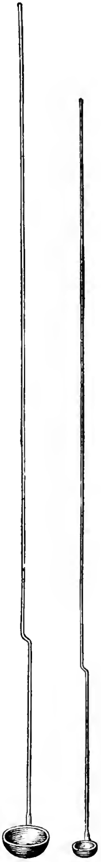
FIG. 26. — Deflagrating spoons.

(*b*) Put a small piece of sulphur in the deflagrating spoon (Figs. 26 and 27), hold the spoon in the flame until the faint blue flame of the burning sulphur can be seen, then lower the spoon into a bottle of oxygen. Describe and explain any change in the flame. Smell the vapor cautiously by wafting a little toward the nose, and state of what the odor reminds you. To what class of compounds, in all probability, does this vapor belong? (If the odor of the burning sulphur is disagreeable, plunge the spoon into water, or remove it to the hood.)

(*c*) Hold the charcoal in the flame long enough to produce a glow, then lower it into a bottle of oxygen. Describe and explain any change in the appearance of the charcoal. How does the chemical change resemble that in (*b*)?

(*d*) Dip the unwound end of the picture cord into the deflagrating spoon containing the melted sulphur used above. If the sulphur on the end is not burning, light it, and quickly lower it into a bottle of oxygen. The picture cord should take fire and burn brilliantly. A picture cord is chemically the same as iron. Can the change be explained similarly to that in (*b*) and (*c*)? Is there any visible, material proof of change besides the vapor from the burned sulphur? Does the product of the chemical change belong to the same class as the products in (*b*) and (*c*)?

(*e*) Hold one end of a piece of straightened magnesium ribbon about 10 cm. long in the forceps, light the other end, and instantly lower it into a bottle of oxygen. In what general ways does the result resemble those obtained in the other experiments with oxygen? What conclusive evidence is there that the magnesium was chemically changed? How does the product resemble that formed by burning magnesium in the air? Are the products identical?

Write a short account of this experiment in your notebook, including a sketch of the apparatus used for the preparation of oxygen. Arrange the notes somewhat in the order of the above directions, using the same letters, and answering *all* questions.

FIG. 27.—Combustion cup (or deflagrating spoon) made from a piece of copper wire and a piece of crayon having a cavity at the upper end.

LABORATORY EXERCISE. II.

1. Summarize the properties of oxygen.
2. What is its most characteristic property?
3. What oxides were formed in Exp. 14?
4. From what class of substances might oxygen be prepared?
5. From what substances have you previously prepared it?
6. What chemical change probably takes place when a substance burns?

7. If air contains something besides oxygen, what must be the general properties of this other constituent?

It was shown in Exp. 11 that when a metal is heated in the air it gains in weight, and in Exps. 12 and 13 that the gain is actually due to the formation of a new substance, composed of the metal and oxygen. The experiments with oxygen, also, show that this gas unites easily and vigorously with metals, like iron and magnesium, and with other substances, like sulphur and charcoal (chemically known as carbon), forming oxides. These oxides belong to a large class of substances called **Chemical Compounds**. The experiments thus far performed show that chemical compounds are produced under special conditions, and that they possess specific properties. Additional facts may be learned about chemical compounds by studying the mathematical relation of oxygen to two of its compounds.

### THE COMBINATION OF OXYGEN WITH MAGNESIUM.

**Experiment 15.** — *Method:* Heat a weighed quantity of magnesium in a covered crucible at a low temperature, weigh the product, and the gain will be the weight of oxygen which has combined with the known weight of magnesium.

*Supplies:* Porcelain crucible and cover, powdered magnesium, forceps, pronged tripod or similar support, crucible block.

*Process:* Clean and dry the crucible and cover, and weigh both together. Record the weight as shown below. Weigh in the crucible from 0.4 to 0.5 gm. of magnesium, taking care to weigh to a centigram the exact amount. Record the weights thus: —

| | Grams. |
|---|---|
| Weight of crucible, cover, and magnesium . . | 13.32 [1] |
| Weight of crucible and cover . . . . . . . . | 12.81 |
| Weight of magnesium . . . . . . . . | 0.51 |

---

[1] These and similar results which will be given are solely for illustration.

Stand the crucible on the tripod, as shown in Fig. 28, and heat for five minutes with a flame which just touches the bottom of the crucible. Grasp the cover firmly by the ring with the clean forceps, cautiously lift it, and if the magnesium glows, cover the crucible instantly. Repeat this operation at frequent intervals, until the glow is dull red and ceases to spread through the mass; then adjust the cover so that a small opening is left between the cover and the crucible, and heat strongly for ten or fifteen minutes. If the contents has ceased to glow, heat the crucible, uncovered, for five minutes. Take care not to upset the cover by accident or insecure handling with the forceps. At no time should the flame touch the cover of the crucible; roughly speaking, the flame should reach as high outside as the magnesium does inside. The operation is complete when the product is gray or white. Cool the crucible gradually. When cool enough to handle, it is cool enough to weigh. Weigh, and record thus: —

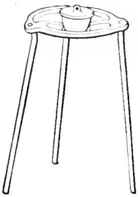

FIG. 28. — Covered crucible supported by a tripod.

| | GRAMS. |
|---|---|
| Weight of crucible, cover, and contents, after heating . . | 13.66 |
| Weight of crucible, cover, and contents, before heating . . | 13.32 |
| Weight of oxygen which has combined with the magnesium | 0.34 |

Heat the uncovered crucible again strongly for five minutes, cool, and weigh again as before. If the weight is not the same, continue until the last two weights are approximately the same.

*Calculation:* Since 0.34 gm. of oxygen combined with 0.51 gm. of magnesium, the ratio in which they combined may be expressed by the proportion —

$$0.34 : 0.51 :: 1 : x$$

$$\therefore\ x = 1.5$$

Class average =

**Discussion of Experiment 15.** — This result means that for every gram of oxygen one and one-half grams (approximately) of magnesium are necessary to form a chemical compound. The truth of this statement is more vivid when the class average is known; for if the work has been done accurately, the class average will agree closely with the theoretical result. More important conclusions will be drawn later from this experiment, but it is sufficient at present to grasp the idea of *definite relations by weight* which exist between the components of a chemical compound.

### THE RELATION OF OXYGEN TO POTASSIUM CHLORATE.

**Experiment** 16. — *Method:* Heat a weighed quantity of potassium chlorate in a covered crucible until all the oxygen is evolved, and the loss will be the oxygen which was in chemical combination in the potassium chlorate.

*Supplies:* Crucible and cover, forceps, support (as in Exp. 15), dry, powdered potassium chlorate, crucible block.

*Process:* Weigh the crucible and cover as in Exp. 15, and weigh in the crucible about 1.5 gm. of dry, powdered potassium chlorate. Weigh to a centigram the exact amount taken, and record the weights thus: —

| | Grams. |
|---|---|
| Weight of crucible, cover, and potassium chlorate . . | 15.34 |
| Weight of crucible and cover . . . . . . . . . . | 14.00 |
| Weight of potassium chlorate . . . . . . . . . | 1.34 |

Arrange the apparatus as shown in Fig. 28, and proceed with the heating somewhat as in Exp. 15. The potassium chlorate first melts, and as the oxygen, which is liberated, bubbles through the molten mass, it may spatter a little potassium chlorate up on the inside of the cover. Regulate the heat, if possible, to avoid this difficulty. Remove the cover with the clean forceps occasionally, and if it is coated with potassium chlorate, lay it (ring side down) on a block of wood, deftly loosen the thin layer with a pin or sharp rod, then hold the cover over the crucible and gently scrape the pieces off into the crucible with a sharp rod. If any fall upon the desk and are touched with the moistened finger, they will cling to the finger and may be returned thereby to the crucible. Regulate the flame so that the bubbling will be slow. As the temperature increases, the mass solidifies, and the action apparently stops. Increased heat, however, drives off more oxygen. As soon as the second evolution of oxygen is passed, remove the cover, lay it (ring side down) on a block of wood where it will be safe, and then heat the crucible strongly for five or ten minutes. The decomposition is complete when the product is perfectly white and non-crystalline, and not changed in appearance by a flame nearly as high as the top of the crucible. Cool gradually, and when cool, weigh the crucible, cover, and contents, as before. Record the weights thus: —

| | GRAMS. |
|---|---|
| Weight of crucible, cover, and contents before heating . | 15.34 |
| Weight of crucible, cover, and contents after heating . | 14.82 |
| Weight of oxygen lost . . . . . . . . . . . | 0.52 |

Heat again strongly for ten minutes, then cool and weigh as before. If the weight is the same as after the first heating, proceed with the calculation; if not the same, heat and weigh until two successive weights are the same, *i.e.* heat to constant weight.

*Calculation:* From 1.34 gm. of potassium chlorate 0.52 gm. of oxygen was obtained, or —

$$1.34 : 0.52 :: 100 : x$$

$$x = 39.3 \text{ per cent.}$$

Class average =          per cent.

**Discussion of Experiment 16.** — This result means that a definite proportion of oxygen by weight can always be obtained from potassium chlorate. Individual experiments may produce slightly varying results, but an average of several results will be substantially identical with the theory. As in Exp. 15, the full significance of the result is not yet apparent, and later the experiment will be discussed in its widest application. This experiment is intended to teach, at this stage of the experimental work, that *definite weight relations* are essential to chemical compounds.

**The Weight of a Liter of Oxygen** will be serviceable, later, in explaining fundamental facts. Before it can be determined, however, the contents of Appendix B should be mastered.

### PROBLEMS. III., IV., V., VI.

See Appendix B, §§ 2, 5, 7.

### CLASS-ROOM EXERCISE. II., III., IV.

See Appendix B, §§ 5, 6, II.

### WEIGHT OF A LITER OF OXYGEN.

**Experiment 17.** — *Method:* Weigh the volume of water displaced by generating a known weight of the gas.

*Apparatus:* The apparatus is shown in Fig. 29. *A* is a test tube containing a mixture of potassium chlorate and manganese dioxide; it is attached to the bent tube *F* by a small rubber stopper. *B* is a bottle to be filled with water; it is provided with a two-hole rubber stopper, through which pass *F* and *C*, the latter being the glass portion of the whole delivery tube, *CC′*, while *C′* is rubber. A Hofmann screw is attached at the point *E*. A large bottle, *D*, serves to catch the water forced over from *B* through *CC′* by the oxygen generated in *A*. The bent tube *F* retains any moisture driven over from *A*, and is weighed with *A*. All joints must be air tight.

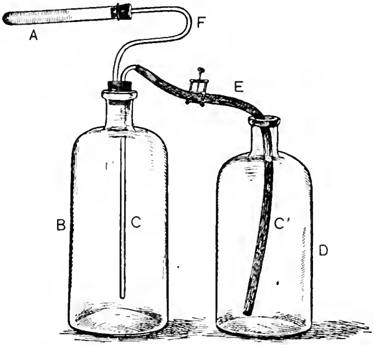

FIG. 29. — Apparatus for determining the weight of a liter of oxygen.

*Process:* (1) Fill *A* two-thirds full of a mixture of manganese dioxide (2 parts) and potassium chlorate (3 parts). Each substance must be powdered and free from organic matter (*e.g.* paper, cork, straw). The mixture should be dried by heating it in an oven to about 110° C., on a radiator or on some convenient heated object. Push a little glass wool, or shredded asbestos (previously ignited to a red heat), into the tube to hold the contents in place. The tube must be free, inside and out, from loose particles of manganese dioxide, or other matter which might be rubbed off after the weighing — *clean*, in other words. Weigh *AF* to a centigram. Weigh the empty, dry, clean bottle, *D*, to a decigram on the trip scales.

(2) Fill *B* with water nearly to the neck. Fill *CC′* with water and tighten the Hofmann screw to prevent the water from running out. Insert *AF* into the stopper of *B*. Push the stopper into the bottle, slowly at first, then hard; if water rises in *F*, loosen the screw at *E* slightly, remove *A*, and blow gently into *F* to force the water back into *B*. When properly adjusted, the water should be in *B* and *CC′*, but not in *F*. Replace *A*, taking care not to crush the thin glass by pushing it too hard upon its stopper; open the screw at *E*. If the apparatus is tight, the water will not flow out. It should be adjusted until air tight. Leave the screw open.

(3) Heat *A* gently with a low flame, beginning at the closed end and keeping the flame back of any moisture which might be driven out of the mixture. The liberated oxygen will force the water from *B* into *D*. Heat *A* just hot enough to cause a gentle flow of water into *D*. When *D* is about three-fourths full, decrease the heat gradually. If moisture appears near the top of the mixture, heat behind it, and either drive it over into *F* or cool *A* so gradually that the moisture will not shatter the tube. While *A* is cooling sufficiently to weigh, stand a thermometer in *D*; also read the barometer. When *A* is cold, raise *B* until the water is at the same level in *B* and *D*, pinch *C′* tight and remove it from *D*. Read and remove the thermometer. Dry *D* on the outside, if necessary, and then weigh it, using the same large weights as before; the gain in weight (in grams) of *D* gives the volume (since 1 gm. of water = 1 cc.) of oxygen liberated. Weigh *AF*; its loss in weight is the weight of the oxygen which passed into *B*.

*Calculation:* The observed volume must be reduced to the volume it would occupy, if it were at 0° C., 760 mm., and in the dry state, *i.e.* reduced to standard conditions. (See App. B, § 7.) This is done by the formula —

$$V = \frac{\text{V}' (\text{P}' - a)}{760\ (1 + .00366\,t)}.$$

Substitute the proper values in this formula, and solve for *V* — the corrected volume of oxygen liberated.

$$V = \qquad \text{cc.}$$

Since 1 l. contains 1000 cc., then

$$\frac{V}{1000} = \qquad \text{l. of oxygen liberated.}$$

The weight of oxygen liberated is found thus: —

| | GRAMS. |
|---|---|
| Weight of $AF$ before liberation of oxygen . . | |
| Weight of $AF$ after liberation of oxygen . . . . . . | |
| Weight of oxygen liberated . . . . . . . . | |

Therefore the weight of 1 l. of oxygen equals

$$\frac{\text{Wt. of oxygen liberated}}{\text{Vol. of oxygen liberated}} = \frac{\qquad}{\qquad} = \qquad \text{gm.}$$

Substitute in this formula the proper values. Tabulate the total result of the experiment thus: —

WEIGHT OF 1 L. OF OXYGEN.

| Found. | Class Average. | Correct Weight. |
|---|---|---|
| | | |

**Definitions. — Combination** may mean a process or its product. Thus the process employed in Exp. 15 is called combination, and the product, magnesium oxide, is also called a combination of magnesium and oxygen. The context, however, usually eliminates any ambiguity. A synonym of the process, combination, is **Synthesis**, which literally means "a putting together." **Decomposition** means the separation of a chemical compound into its parts. These parts are called decomposition products, or, better, **Components**. And as decomposition may be partial or complete, these components may be elements or compounds. A synonym of decomposition is **Analysis**, which literally means "a breaking up." Thus in Exp. 16, the potassium chlorate was actually broken up into the element, oxygen, and the compound, potassium chloride. A **Determination** is an accurate chemical examination. It may be synthetical, analytical, gravimetric — involving weights, or volumetric — involving volume. Experiment 16 was a gravimetric determination of the per cent of oxygen in potassium chlorate. A **Residue** is usually the product remaining after something has been removed from a compound or a mixture. Thus in Exp. 16 the residue is the potassium chloride remaining in the crucible. A **Chloride** belongs to a class of compounds, each member of which consists of the element chlorine and one other element. Chlorides bear the same relation to chlorine as oxides bear to oxygen.

**Oxidation.** — Oxides were mentioned in Chapter I. The process of forming an oxide, or in general of adding oxygen to an element or compound, is called **Oxidation.** Thus Exp. 15 was an oxidation of magnesium. The metals heated in the experiments performed in Chapter I. were also oxidized. Compounds which yield oxygen easily are called oxidizing agents; thus nitric acid, a compound to be studied later, decomposes easily, and one of the decomposition products is oxygen, which often immediately combines with some other element or some compound. Oxidation is an important chemical operation.

**Law of Definite Proportions by Weight.** — Experiment 15 showed that when magnesium is heated so that it has access to oxygen the two elements combine in the ratio of 1 to 1.5. This also means that the product, magnesium oxide, always contains 1.5 times as much magnesium as oxygen by weight. Experiment 16 shows that potassium chlorate always yields about 39 per cent of oxygen. A vast number of experiments has verified these, and similar results, in the case of other compounds. The fact of definite proportions by weight revealed by these two experiments (15 and 16), and extended to cover many others, is the result of a law known as the "Law of Definite Proportions by Weight." It was established as the outcome of a controversy between two chemists, Proust and Berthollet, which lasted from 1799 to 1806, though isolated suggestions of the law had previously appeared. The law may be stated thus: —

*A chemical compound always contains the same elements in the same proportions by weight.*

Or,

*A chemical compound has a definite composition by weight.* This law is one of the foundation stones of chemistry.

**Compounds and Mixtures.** — Proust, in his discussion with Berthollet, showed conclusively the distinction between mixtures and compounds. The "gradual changes" found in the substances analyzed by Berthollet were shown by Proust to be due to the fact that Berthollet analyzed mixtures — *masses containing ingredients in varying proportions.* Mixtures are the opposite of compounds; they may have any composition; their ingredients are not held together by that form of energy called chemical affinity, or attraction. A mixture of sulphur and iron can be easily separated by mechanical means, but in a compound of iron and sulphur the components are held together by a force which does not yield to the influence of a solvent or a magnet, nor to any other purely physical agent.

### CLASS-ROOM EXERCISE. VI.

1. Additional study of oxygen.
    - (*a*) Occurrence and distribution.
    - (*b*) Chemical relation to plants and animals.
    - (*c*) Liquid oxygen.
    - (*d*) Uses.
    - (*e*) Preparation on a large scale.
2. Historical.
    - (*a*) Discovery of oxygen.
    - (*b*) Essential facts of Priestley's life.
    - (*c*) Scheele's and Lavoisier's actual contribution to the discovery.
    - (*d*) Names of oxygen, by whom so called, and why.
3. Combustion.
    - (*a*) Theories of Becher and Stahl.
    - (*b*) Lavoisier's work.
    - (*c*) Present views.
4. Manipulation.
    - (*a*) Rules for weighing.
    - (*b*) Heating and cooling porcelain.
    - (*c*) General precautions in accurate experiments.
    - (*d*) Sources of error in Exps. 15 and 16, how detected, avoided, and remedied.
5. Elements and compounds.
    - (*a*) Definition of each.
    - (*b*) Enumerate all the elements and compounds thus far used or studied.
    - (*c*) What compounds have been made in the previous experiments? What compounds have been decomposed?
    - (*d*) Define and illustrate: a mixture, combination, decomposition, determination, residue, oxide, oxidation, chloride, component.
6. Law of Definite Proportions by Weight.
    - (*a*) Historical development.
    - (*b*) Statement in several different forms.
    - (*c*) Illustrations.

## PROBLEMS. VII.

1. How many grams of magnesium will exactly combine with 16 gm. of oxygen? With 32 gm. of oxygen? With 79 gm. of oxygen?

2. How many grams of oxygen will exactly combine with 24 gm. of magnesium? With 48 gm. of magnesium? With 90 gm. of magnesium?

3. What weight of oxygen gas could be obtained by decomposing 200 gm. of potassium chlorate? What weight of oxygen, if the potassium chlorate had contained 10 per cent of an impurity?

4. What weight of oxygen gas could be obtained from 100 gm. of pure potassium chlorate? If the chlorate used contained 12 per cent of its weight of potassium chloride, what would be the difference in the amount of oxygen obtained?

5. If a liter of oxygen weighs 1.43 gm., what will be the weight of oxygen in a room 25 m. long, 17 m. wide, and 15 m. high?

6. How many liters of oxygen can be obtained from 122.5 gm. of potassium chlorate? From 255 gm. of potassium chlorate, allowing 2 per cent for impurities in the potassium chlorate?

7. A quantity of potassium chlorate weighing 721 gm. is decomposed, the oxygen collected, and magnesium is burned in the oxygen until no gas remains.

(*a*) What weight of oxygen is produced?
(*b*) What volume of oxygen is produced?
(*c*) What weight of magnesium is required?
(*d*) What weight of magnesium oxide is produced?

# CHAPTER III.

## HYDROGEN.

THE element hydrogen is a gas under ordinary conditions. It was recognized as a "peculiar and distinct substance" in 1766 by the English chemist Cavendish. It is a component of many compounds, and is the standard of several scientific measurements.

Hydrogen was first prepared by the interaction of an acid and a metal.

### PREPARATION OF HYDROGEN BY METALS AND AN ACID.

**Experiment 18.** — *Supplies:* Test tubes, granulated zinc, iron filings, dilute sulphuric acid, dilute hydrochloric acid.

(*a*) Put a few pieces of granulated zinc in a test tube and cover them with dilute sulphuric acid. A gas will bubble through the liquid. After a minute or two test the gas by holding a lighted match at the mouth of the test tube. Describe the result.

(*b*) Repeat (*a*), using dilute hydrochloric acid. Is the final result the same? If not, repeat both (*a*) and (*b*).

(*c*) Put a little iron filings in a test tube and cover them with dilute sulphuric acid. Test the gas evolved, as in (*a*). Is it the same gas?

(*d*) Repeat (*c*), using dilute hydrochloric acid, and test the gas, as before. Is it the same gas?

What is the source of the hydrogen? What are its most characteristic properties?

The slight explosion heard in some, perhaps all, of the above experiments is due to impure hydrogen. What, in all probability, is the impurity? Does the observation suggest any danger in working with hydrogen? What is the simplest way to avoid the danger?

The specific nature of hydrogen may be learned from its preparation on a larger scale than in Exp. 18 and a study of its properties.

### PROPERTIES OF HYDROGEN.

**Experiment 19.**—*Supplies:* Hydrogen generator, 25 gm. of granulated zinc, dilute sulphuric acid, pneumatic trough, five bottles (each 250 cc.), taper.

*Apparatus:* The generator is shown in Fig. 30. *A* is a 500 cc. flask provided with a two-hole rubber stopper. *B* is a safety tube. *C* is the double right angle bend which was made in Exp. 2. *D* is a rubber connector about 4 cm. long. *E* is a delivery tube of such length that its end is level with the bottom of the flask. The flask stands on the desk, and *E* dips into the pneumatic trough.

*Precautions:* All joints must be tight. The stopper must fit perfectly, and both *B* and *C* must turn with difficulty after the stopper has been pushed well into the flask. Acid must be left in the bend of the safety tube to about the point indicated by the dotted line. (See App. A, § 5.) *All flames, large and small, should be extinguished in the vicinity during this experiment, because mixtures of air and hydrogen, when ignited, explode violently.* Save the contents of the generator for Exp. 20.

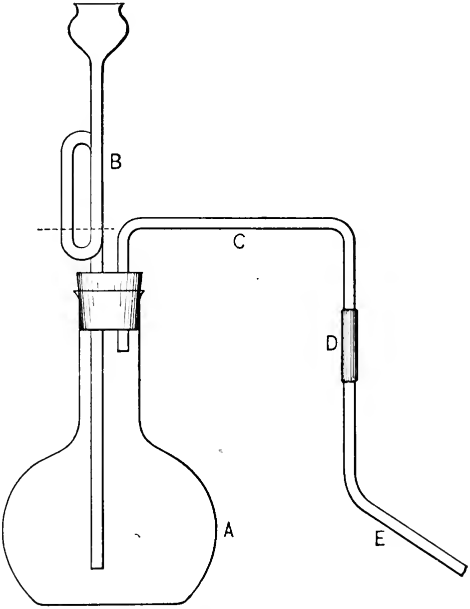

FIG. 30.—Hydrogen generator.

*Process:* Incline the flask and slide the zinc into it from a paper. Do not drop it in — the bottom is thin and may crack. Insert the stopper with its tubes by holding the flask by the neck in one hand and with the other working the stopper slowly into the flask. Adjust the flask and delivery tube so that the end of *E* is on the bottom of the pneumatic trough and under the hole in the shelf. Have the bottles filled with water and inverted in the trough, and filter paper ready to cover the bottles of gas.

Pour enough dilute sulphuric acid through the safety tube into the flask to cover the zinc, taking care to leave some acid in the lower bend of the safety tube. This latter precaution prevents the gas from flowing out the back of the apparatus; if at any time the gas should flow backward, pour a little more acid into the bend. If the acid will not run down the safety tube, loosen the stopper for an instant. Remember, however, that this operation may admit air, and the gas evolved for the next few minutes will be impure. As soon as the acid and zinc begin to interact, hydrogen will be rapidly evolved. Allow the action to proceed several minutes to free the apparatus from air, then collect five bottles of gas over water, as in the case of oxygen. As fast as the bottles are full, remove and cover them with wet filter paper. When the gas has been collected, remove the generator to the hood. Proceed *at once* to study the properties of hydrogen as follows: —

(*a*) Uncover a bottle for an instant and then drop a lighted match into the bottle. Describe and explain the result.

(*b*) Remove the paper from a bottle of hydrogen and allow it to remain uncovered for three minutes — by the clock. Then show the presence or absence of hydrogen by dropping a lighted match into the bottle. Explain the result.

(*c*) Verify the result in (*b*) thus: Hold a bottle of air over a bottle of hydrogen, remove the paper from the hydrogen bottle, then bring the bottles into the relative positions shown in Fig. 31. Hold them there for a minute or two, then stand the lower bottle on the desk, cover the upper one with wet filter paper and stand it on the desk. Drop a lighted match into each bottle. What has

FIG. 31. — Pouring hydrogen.

become of the hydrogen? What property of hydrogen is shown by (*b*) and (*c*)?

(*d*) *Read the directions carefully and watch all parts of the apparatus for definite results.* Invert a bottle of hydrogen, remove the paper, and thrust a lighted taper up into it. Withdraw the taper slowly and then insert it again. Does the hydrogen burn? If so, where? Does the taper burn when in the bottle? When out of the bottle? Feel of the neck of the bottle, and describe and explain. Repeat with the remaining bottle of hydrogen. What *three* properties of hydrogen are shown in (*d*)?

### LABORATORY EXERCISE. III.

1. Sketch in your note-book the generator used to prepare hydrogen in Exp. 19.
2. Summarize the properties of hydrogen.
3. What is its most characteristic property?
4. Why is there danger of an explosion in generating hydrogen? How may the danger be averted?
5. From what class of substances may hydrogen be prepared?

**Chemical Reaction.** — The mutual chemical action which results when two or more elements or compounds are added to each other, is called a **Reaction**. Thus when zinc and sulphuric acid come together there is chemical action, apparent so far only by the formation of hydrogen. Inasmuch as every experiment involves one or more reactions, an examination of the reaction which takes place in the preparation of hydrogen will serve as a preliminary study of reactions. Not every fact about a reaction can be proved at the outset. Certain fundamental facts are always utilized by chemists. For our present purpose, we must regard zinc as an element and sulphuric acid as a compound of hydrogen, sulphur, and oxygen. Now in this reaction the zinc and hydrogen exchange places, and the result is the formation of hydrogen — already shown — and zinc sulphate. Furthermore, we can prove that a definite relation by weight exists between the zinc used and hydrogen produced. The present problem, then, has two steps: (1) the identification of zinc sulphate as the second product of the reaction, and (2) the actual determination of the relations by weight between the interchanging zinc and hydrogen.

**Tests.** — The identity of an element or a compound is established by a **Test**, *i.e.* by subjecting it to the action of various chemicals and comparing the results with the recorded properties of the substance under examination. Sometimes the test is a single operation, sometimes several different examinations are necessary. The test for zinc sulphate involves two distinct steps: (1) the identification of the group to which it belongs — the sulphate group, and (2) the determination of the kind of sulphate, *i.e.* the establishment of those properties which distinguish *zinc* sulphate from all other members of the sulphate group. If the substance in question does not show these properties, then, of course, it is not zinc sulphate, and other tests must be applied to determine its nature.

**Crystallization.** — Most substances to be tested must first be purified. A convenient method of purification is crystallization. Most solid substances are more soluble in hot than in cold water. Hence a hot liquid on cooling will often deposit some of the dissolved solid, especially if the liquid contains much solid in solution. If the hot liquid is cooled slowly, the solid is deposited in masses having a more or less definite geometrical form, called crystals. This operation is called **Crystallization**, and the substance is said to **Crystallize**. If a solution evaporates slowly in the air, crystals will form when the solution reaches the same condition as a cooled solution, *i.e.* a point where the liquid cannot hold all the solid in solution. This point must be found by trial, hence no rules can be given to obtain well-shaped crystals, except perhaps (1) slow evaporation in the air, if time permits, or (2) frequent cooling, if the evaporation is caused by heating. When a substance crystallizes, most of the impurities remain behind in solution, hence the crystallized solid is usually pure. Repeated crystallization — recrystallization — is a common method of purification, whenever it is applicable. Usually every chemical substance has one or more distinct geometrical forms in which it crystallizes. This crystal form is also a means of identification, but the chemical tests are more decisive.

### CRYSTALLIZATION OF ZINC SULPHATE.

**Experiment 20.** — Fold a filter paper according to the directions given on page 8, fit it into a glass funnel, and moisten the paper with water to hold it in place. Support the funnel by a ring or filter stand, and pour the contents of the hydrogen generator gradually upon the filter, taking care that no liquid comes above the edge of the paper. (See App. A, § 4.) If the contents of the generator consists partly of a white solid, add a little hot water and shake the mass until the solid is dissolved. The filtrate should be caught in a casserole, or evaporating dish. Wash any remaining zinc with water to remove the acid, and preserve for future use. When the filtrate measures about 100 cc., place the casserole on a gauze-covered tripod, or similar support, and heat the liquid for ten or fifteen minutes, but do not allow it to boil violently. (See App. A, §§ 2, 3.) Then cool it slowly, and crystals of zinc sulphate will be deposited. If no crystals separate after a few hours, evaporate again until a thin film appears on the surface when the solution has cooled. Then if none are deposited from the cold solution, consult the teacher. If for any reason the filtrate cannot be immediately evaporated, let it remain undisturbed, and crystals will be deposited as the water evaporates. This liquid, of course, should not be boiled.

In any case remove the crystals and dry them between filter paper. Describe them, giving color, luster, approximate shape and size, and any other striking property. Put any remaining crystals into the stock bottle marked **Zinc Sulphate.**

### TEST FOR ZINC SULPHATE.

**Experiment 21.** — Dissolve some of the crystals of zinc sulphate obtained in Exp. 20, in a test tube half full of water, divide the solution into halves, and proceed as follows: —

(*a*) *Test for zinc:* To one part add a little sodium hydroxide solution. A white jellylike solid will be formed, but it will disappear if considerable sodium hydroxide solution is added and the mixture well shaken. The insoluble jellylike compound is zinc hydroxide, which is changed by the excess of sodium hydroxide into soluble sodium zincate. Now add a little hydrogen sulphide solution, and a white solid is formed. It is zinc sulphide. Compare its color and texture with those of zinc hydroxide. Shake the test tube containing the zinc sulphide to suspend the solid in the liquid, and divide into halves. To one add dilute hydrochloric acid, and to the other add acetic acid. Shake each tube and observe the results. Zinc sulphide is decomposed by hydrochloric acid, but not by acetic acid. These facts are characteristic of compounds containing zinc, but not true of compounds of other metals. Hence zinc must be one of the components of the solid product of the reaction between zinc and sulphuric acid.

(*b*) *Test for a sulphate:* To the other half of the original solution add a little barium chloride solution. A white solid is formed, called barium sulphate, which will be found to be insoluble, in both hydrochloric and acetic acids. Try it. Barium sulphate is always formed when barium chloride is added to the solution of a sulphate, or to sulphuric acid, which is closely related to all sulphates. Hence the zinc compound formed in the hydrogen generator must be a sulphate.

(*c*) Draw a final conclusion from the results obtained in (*a*) and (*b*).

**Definitions.** — The test made in Exp. 21, which showed that zinc sulphate is the compound formed by the reaction which liberates hydrogen, is called a **Qualitative** test, since it determined only the quality of the compound, not the quantity by weight or volume of each component. The latter operation is called a **Quantitative** test. The replacement of one element or group of elements by another element or group of elements is called **Substitution.** Thus, the formation of zinc sulphate is the substitution of zinc for hydrogen in sulphuric acid. So also the formation of barium sulphate consists in the substitution of barium for zinc in the zinc sulphate; the zinc, as could be easily proved, takes the place left by barium in the barium chloride. Substitution is a broad, complex subject, which has a deeper meaning than mere replacement. It will be clearer as the experiments proceed. A **Sulphate** is a compound derived from sulphuric acid by substituting a metal for the hydrogen in the acid. A **Sulphide** is a compound of sulphur and *one* other element or group acting like an element. Sulphides correspond to oxides and chlorides — all being binary compounds, *i.e.* compounds of two elements or groups acting like elements.

## QUANTITATIVE EXAMINATION OF THE REACTION BETWEEN ZINC AND SULPHURIC ACID.

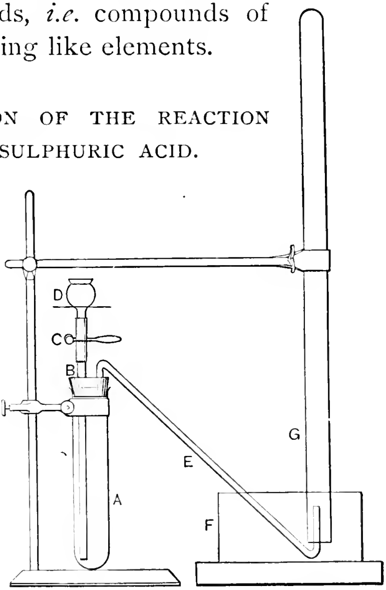

FIG. 32. — Apparatus for determining the quantitative relation between zinc and hydrogen.

**Experiment 22.** — *Method:* Measure the hydrogen liberated from sulphuric acid by a known weight of zinc.

*Apparatus:* The construction and arrangement of the apparatus is shown in Fig. 32. *A* is a large test tube provided with a two-hole rubber stopper. The tube *B* passes through one hole to the bottom of the test tube and is connected by a short piece of pressure tubing with the cup *D*. A Mohr's pinchcock, *C*, controls the flow of acid from *D* into *A*. If the pinchcock is not strong enough to close the pressure tubing, then ordinary rubber tubing will answer, if fresh. The part *DB* is made by cutting a thistle tube into the requisite lengths. The ends of *B* and *D* are as close together as the pinchcock will permit. *F* is a shallow dish, and *G* is a 100 cc. graduated tube. The modification shown in Fig. 33 is more convenient, since it is not easily broken and requires no support.

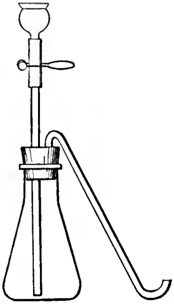

FIG. 33. — Alternate apparatus.

*Process:* Select a single piece of zinc weighing not less than .22 nor more than .24 gm., weigh it exactly, wind a short piece of platinum wire around it and drop it into *A*. Fill *A* full of water and insert the stopper with its tubes. Fill the *whole* apparatus with water thus: Fill *D* and admit water repeatedly until *all* air is forced out of *A*, *B*, and *E*, taking care never to let the water in *D* fall below the point indicated by the dotted line. Fill *F* nearly full, fill *G full* and invert it in *F* over the end of *E*. All water used in this experiment must have stood in the laboratory at least a day.

Fill the cup *D* with hot dilute sulphuric acid, and let the acid run slowly into *A*, but never below the dotted line. Introduce in this way and at this time about 50 cc. of acid. Do not add acid after the action begins unless absolutely necessary, as the zinc may be forced over into *F*. Hydrogen will be liberated and will pass through *E* up into *G*. Let the gas accumulate in *A* at the top, and there will be little danger of losing the zinc. Let the action continue until the zinc disappears, then force over into *G* any gas in the apparatus by admitting water, observing the precautions previously given.

Unclamp *G*, close the end with the second finger, remove it to a tall jar of water which has been standing long enough to assume the temperature of the room, and clamp it so that the water is about the same height within and without the tube. Stand a thermometer in the water and allow the whole to remain undisturbed for at least fifteen minutes, and longer if possible. Then, *without touching the tube where it contains gas*, adjust the height so that the meniscus just touches the surface of the water in the jar. (See App. A, § 22. (1), (2).)

Read the gas volume, the thermometer, and the barometer, observing all the precautions mentioned in App. B, II., (3) and (4). Record numerical results as shown below. Perform the calculation as directed below, and, if possible, before the tube $G$ is used again. The latter precaution may obviate a repetition of the experiment.

*Calculations:* (The figures recorded here are only for illustration. Those actually obtained will probably be different.)

Data: —

Weight of zinc, 0.24 gm.
Volume of hydrogen, 88.5 cc.
Temperature, 14° C.
Barometer, 758 mm.
Aqueous tension, 11.9.

Reduce the observed volume of hydrogen to the volume it would occupy at 0°, 760 mm., and in the dry state by the formula —

$$V = \frac{V' \ (P' - a)}{760 \ (1 + .00366t)}.$$

If the proper values are substituted in this formula, the corrected volume of hydrogen is found to be 82.64 cc. A liter (1000 cc.) of dry hydrogen under standard conditions weighs .0896 gm. Hence the weight of 82.64 cc. is found by solving the proportion —

$$1000 : 82.64 :: .0896 : x$$

$$x = .0074 \text{ gm.}$$

Weight of corrected gas volume = .0074 gm.

Since .0074 gm. of hydrogen is replaced by .24 gm. of zinc, the weight of zinc which will replace 1 gm. of hydrogen is found by the proportion

$$.0074 : .24 :: 1 : x$$

$$x = 32.43.$$

Therefore 32.43 gm. of zinc are chemically equivalent to 1 gm. of hydrogen.

Class average =

*Final Directions:* Wash the whole apparatus free from acid. Preserve the platinum wire for future use. If the final result is incorrect, first examine the arithmetical work, then repeat the readings (gas volume, pressure, and temperature), and recalculate, if blunders were made.

**Discussion of Experiment 22.** — The result found in this experiment should be approximately 32.5. The number, 32.5, is called the equivalent of zinc. The **Equivalent** of an element is the weight of that element which will replace one part by weight of hydrogen. The equivalent of zinc is the same whatever the acid used. The fact that an average of independent results, such as would be obtained by a class, agrees within the limits of experimental error shows that definite relations by weight exist between hydrogen and zinc, when zinc and sulphuric acid interact. This definite and fixed relation has a deeper meaning than can be considered at this stage of the subject. Its significance will be explained after more facts have been accumulated. Just now, it serves as an introduction to the quantitative study of chemical equations.

**Burning Hydrogen.** — It was shown in Exp. 19 that hydrogen burns, but the product was not examined. It was shown in Chapter I. that when metals are burned in air, they combine with oxygen.

### PRODUCT OF BURNING HYDROGEN.

*Perform this experiment with the utmost precaution, since a mixture of air and hydrogen explodes violently, if ignited. Proceed exactly according to the directions.*

**Experiment 23.** — *Method:* Generate hydrogen, and examine the product from a burning jet.

*Supplies:* Hydrogen generator used in Exp. 19, U-tube with two one-hole stoppers, platinum tip, small bottle, calcium chloride.

*Apparatus:* The construction and arrangement is shown in Fig. 34. The lower portion of the delivery tube is removed and the generator is then connected directly with the U-tube, which is filled about two-thirds full of lumps of calcium chloride separated from the stoppers by plugs of cotton. The delivery tube is bent and attached to the U-tube as shown in the figure. The platinum tip, which is made by the method given in App. A, § 13, is attached to the delivery tube by a short, gas-tight, rubber connector. All joints and connections must be gas tight.

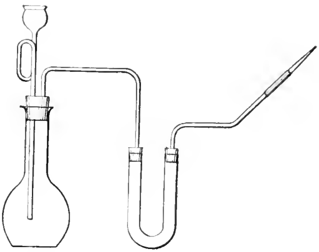

FIG. 34.—Apparatus for burning hydrogen.

*Process:* Pour slowly but continuously through the safety tube enough (about 50 cc.) dilute sulphuric acid upon at least 25 gm. of granulated zinc to produce a steady current of hydrogen gas for about eight minutes. It is advisable to use considerable zinc and a moderate amount of acid. Acid must not be added after the evolution of gas begins, unless, of course, the experiment is begun anew. Let the gas bubble through the acid for at least *two minutes* by actual observation, then attach the platinum tip to the rubber connector at the end of the delivery tube, leaving a short space between the ends of the two glass tubes so that the rubber tube may be compressed suddenly, if necessary. Let the gas run for *another full minute.* This latter precaution is to drive all air out of the tip. Light the hydrogen, and observe *at once* the nature of the flame, its color, heat (by holding a match, or copper, or platinum wire over it), and any other striking property Then hold a small dry bottle over the flame in such a position that the flame is just inside the bottle. When conclusive evidence of the product of burning hydrogen is seen inside the bottle, remove the bottle, and extinguish the flame *at once* by pinching the rubber connector. Remove the generator to the hood, and if the evolution of hydrogen is still brisk, dilute the acid by pouring water through the safety tube. Examine the inside of the bottle. What in all probability is the deposit ? Explain its formation. What previous experiments does this one resemble ? What experiment would have enabled you to predict the result in this experiment ?

Sketch in your note-book the apparatus used in this experiment.

*Final Directions:* When the apparatus is taken apart, replace the stoppers of the U-tube with corks to preserve the contents for further use. Insert the glass end of the platinum tip into the hole at the smaller end of a rubber stopper or cork and insert it into a test tube. This device will protect the platinum end of the tip. Pour off the acid from the zinc, wash the zinc with water to remove all acid, and preserve for further use.

**Discussion of Experiment 23.**—The product of burning hydrogen is water, though the final proof cannot be made until the composition of water is studied.

### LABORATORY EXERCISE. IV.

1. Calcium chloride absorbs water. Explain the use of the U-tube in Exp. 23.
2. Is water an oxide? Why?
3. How does the heat of the hydrogen flame compare with its luminosity?

**Weight of a Liter of Hydrogen.**—Hydrogen is the lightest known substance. A liter at 0° C. and 760 mm. weighs only .0896 gm., while a liter of oxygen under the same conditions weighs 1.43 gm., as was found by Exp. 17. The number .0896 must be remembered. Hydrogen is usually taken as the standard of density in the case of gaseous elements and compounds, though air is sometimes so regarded. Thus the density of oxygen on the air standard is 1.105, and on the hydrogen standard is approximately 16, —

$$.0896 : 1.43 :: 1 : 16.$$

The significance of this relation will subsequently appear.

### CLASS-ROOM EXERCISE. VII.

1. Additional study of hydrogen.
    - (*a*) Distribution.
    - (*b*) Compounds of hydrogen thus far studied or used.
    - (*c*) Uses.
    - (*d*) Liquid hydrogen.
    - (*e*) Solid hydrogen.

2 History of hydrogen.

(*a*) Discovery.
(*b*) Essential facts of Cavendish's life.
(*c*) Names of the gas, by whom so called, and why ?

3. Sources of error in Exp. 22 ; how detected, avoided, and remedied.

(*a*) Weighing the zinc.
(*b*) Reading the graduated instruments.
(*c*) Blunders in the calculation.

4. Compound blow-pipe.

(*a*) Structure (with outline sketch).
(*b*) Action.
(*c*) Properties of its flame.
(*d*) Precautions in using it.
(*e*) Uses.

5. Lavoisier's experiment.

Lavoisier passed steam through a red-hot tube containing iron filings. A combustible gas and a black solid were obtained.

(*a*) What is the name of the gas ?
(*b*) What in all probability are the components of the black solid?
(*c*) Devise an apparatus for this experiment.
(*d*) Read the historical value of this experiment in a history of chemistry.

6. Impurities are often removed from hydrogen gas by passing it through a solution of potassium permanganate. This compound contains a relatively large proportion of oxygen. which under certain conditions is easily liberated. What must be the general character of the change in the removal of the impurities?

7. Crystallization and precipitation.

(*a*) Favorable conditions for crystallization.
(*b*) Why are perfect crystals rare ?
(*c*) How does precipitation differ from crystallization ?
(*d*) The distinguishing characteristics of the six crystal systems.
(*e*) Define "crystallized," "amorphous," "dimorphous," "pseudomorph."
(*f*) Define "precipitate" and "precipitation."

8. Miscellaneous.

(*a*) What is effervescence, and how does it differ from boiling? What experiments in this chapter illustrate effervescence?

(*b*) What is meant by the "diffusion of gases"? What part of Exp. 19 illustrates diffusion?

(*c*) Describe a generator, a safety tube, a delivery tube, a tripod.

(*d*) What synthesis was made in one or more of the experiments with hydrogen? What oxidation? What determination? What chemical compounds additional to those in previous experiments were made or used? What new elements (free or combined) were prepared or used?

## PROBLEMS. VIII.

1. Calculate the equivalent of zinc from the following data actually obtained by a student: —

Weight of zinc .12 gm.
Corrected volume of hydrogen 40.89 cc.

2. Calculate the equivalent of zinc from —

Weight of zinc .50 gm.
Observed volume of hydrogen 183 cc.
Pressure 748 mm.
Temperature 9° C. ($a = 8.57$.)

3. What is the weight of 500 cc. of dry hydrogen at 0° C. and 760 mm.? Of 1800 cc.? Of 9 l.?

4. How many grams of zinc will liberate 100 gm. of hydrogen from sulphuric acid? How many grams of zinc will liberate 100 l.?

5. What is the weight of a liter of hydrogen measured *over water* at 50° C. and 790 mm.? ($a = 91.98$.)

6. How much zinc (in grams) is necessary to liberate from sulphuric acid 100 l. of dry hydrogen at 91° C. and 800 mm.?

7. How many times heavier than a liter of hydrogen is a liter of oxygen, both being dry and under standard conditions?

## CHAPTER IV.

### SYMBOLS — FORMULAS — CONSERVATION OF MATTER — CHEMICAL EQUATIONS — QUANTITATIVE INTERPRETATION OF EQUATIONS — PROBLEMS BASED ON EQUATIONS — PROBLEMS.

**Symbols.** — A symbol is an abbreviation of the name of a chemical element. It is usually the first letter of the name of the element. Thus O is the symbol for oxygen, H for hydrogen, N for nitrogen; as more than one element has the same initial letter, another letter in some cases is added. Thus B is the symbol of boron, but Ba of barium, Bi of bismuth, Br of bromine, etc. The symbol of several metals is derived from their Latin name. Thus, the symbol of iron is Fe (*ferrum*), of lead is Pb (*plumbum*), of sodium is Na (*natrium*), of potassium is K (*kalium*), etc. These symbols denote one atom of the element, *i.e.* H means the smallest particle of hydrogen which can take part in a chemical change. They also represent a certain number, called the atomic weight, which belongs to that atom. They are sometimes loosely used to mean any convenient amount of the element. A list of symbols is given in Appendix C, Table I. If more than one atom is to be designated, the proper numeral is placed before the symbol. Thus: —

2 O means 2 atoms of oxygen.
3 H means 3 atoms of hydrogen.
4 P means 4 atoms of phosphorus.

But if the atoms are in chemical combination, either with themselves or other atoms, then a small numeral is placed after and a little below the symbol. Thus: —

$H_2$ means 2 atoms of hydrogen in combination.
$N_3$ means 3 atoms of nitrogen in combination.
$P_4$ means 4 atoms of phosphorus in combination.

**Formulas.** — A formula is a group of symbols which expresses the composition of a compound. Thus, $KClO_3$ is the formula of potassium chlorate, and $MnO_2$ of manganese dioxide. Formulas represent single molecules. Thus $KClO_3$ means one molecule of potassium chlorate containing one atom each of potassium and chlorine and three atoms of oxygen. They also represent the molecular weight of the compound. More than one molecule is designated by the proper numeral placed before the formula, thus: —

$2\ KClO_3$ means 2 molecules of potassium chlorate.
$3\ H_2O$ means 3 molecules of water.
$4\ H_2SO_4$ means 4 molecules of sulphuric acid.

Groups of atoms acting like a single atom in reactions are enclosed in a parenthesis, or separated by a period. Thus $(NH_4)NO_3$ represents a molecule of ammonium nitrate, in which the group of atoms $(NH_4)$ bears such relations to the $NO_3$ that the group is represented as a unit; and $C_2H_5 \cdot OH$ similarly represents a molecule of alcohol. A group of atoms to be multiplied is also enclosed in a parenthesis. Thus, $Pb(NO_3)_2$ is the formula of lead nitrate, and means that the $(NO_3)$ group bears to the atom of lead (Pb) such relations that the group must be multiplied by two. Many facts similar to the last will be clearer as the experiments proceed.

Symbols and formulas have a deeper significance, which will be considered as the necessary facts accumulate.

### CLASS-ROOM EXERCISE. VIII.

1. Give the symbol of each of the following elements, all of which have been used or studied in the preceding experiments: —

Oxygen, Hydrogen, Sulphur, Carbon, Iron, Magnesium, Zinc, Mercury, Platinum, Lead.

2. Name the elements which correspond to the following symbols: —

Fe, Pb, Pt, Zn, Hg, Na, K, Cu.

3. Give the formula of each of the following compounds: —

Magnesium Oxide, Mercuric Oxide, Potassium Chlorate, Potassium Chloride, Sulphuric Acid, Hydrochloric Acid, Zinc Sulphate, Sulphur Dioxide, Carbon Dioxide.

**Conservation of Matter.** — Most experiments transform matter, but they never destroy or create it; no weight is gained or lost. The total weight of matter involved in a reaction is the same before and after. This vast conception was first demonstrated by Lavoisier, and, largely as a result of his initial work, it has become a fundamental law of chemistry. It is called the Law of the Conservation of Matter, and is often stated thus: —

*No weight is lost or gained in a chemical reaction.*

**Chemical Equations.** — The mathematical significance of symbols and formulas, together with the conception of the conservation of matter, allows reactions to be represented quantitatively by equations. Chemical equations differ from algebraic equations in one vital respect: they are the result of experiment. Thus in Exp. 15 it was shown that magnesium unites with oxygen to form magnesium oxide. These facts are represented in the simplest way, thus:—

$$\underset{\text{Magnesium}}{\mathrm{Mg}} \quad + \quad \underset{\text{Oxygen}}{\mathrm{O}} \quad = \quad \underset{\substack{\text{Magnesium}\\ \text{Oxide}}}{\mathrm{MgO}} \qquad (1)$$

This equation means that one atom of magnesium unites with one atom of oxygen, and forms one molecule of magnesium oxide In Exp. 16 it was shown that potassium chlorate when heated yields oxygen and a residue (called potassium chloride). This reaction is represented in the simplest way by the equation:—

$$\underset{\substack{\text{Potassium}\\ \text{Chlorate}}}{\mathrm{KClO_3}} \quad = \quad \underset{\text{Oxygen}}{3\mathrm{O}} \quad + \quad \underset{\substack{\text{Potassium}\\ \text{Chloride}}}{\mathrm{KCl}} \qquad (2)$$

This equation means that one molecule of potassium chlorate yields by decomposition three atoms of oxygen and one molecule of potassium chloride. It has been shown as the result of several experiments that the interaction of zinc and sulphuric acid produces hydrogen and zinc sulphate. This fact is represented by the equation:—

$$\underset{\text{Zinc}}{\mathrm{Zn}} \quad + \quad \underset{\substack{\text{Sulphuric}\\ \text{Acid}}}{\mathrm{H_2SO_4}} \quad = \quad \underset{\text{Hydrogen}}{\mathrm{H_2}} \quad + \quad \underset{\substack{\text{Zinc}\\ \text{Sulphate}}}{\mathrm{ZnSO_4}} \qquad (3)$$

This equation means that one atom of zinc interacting with one molecule of sulphuric acid produces two (combined) atoms of hydrogen and one molecule of zinc sulphate.

These equations have not been completely proved by the particular experiments, but the work of others supplements the results obtained and permits important conclusions.

It should be noted —

(1) That these equations are expressions between atoms and molecules, not between the quantities actually seen in action.

(2) That some substances are not included in the equation. Thus in equation (3) water is not represented because it takes no chemical part in the reaction; it simply serves to dissolve the zinc sulphate from the surface of the zinc, and thereby allow the reaction to proceed. Likewise in equation (1) no nitrogen appears. Nitrogen combines with magnesium only when the latter is at a red heat in the absence of oxygen — obviously impossible in Exp. 15.

(3) That only the beginning and end of reactions are represented. Thus, in Exp. 9, certain compounds are formed and decomposed again as the temperature increases; but these facts do not appear in the equation, because here we are concerned not with phases, but only with the actual ultimate decomposition of potassium chlorate into oxygen and potassium chloride.

(4) That chemical equations *prove* nothing. They simply help us interpret an experiment. They are not a sufficient record of an experiment, and they should never be used unless they have been proved in whole or in part.

The atoms or molecules entering into the initial stage of a reaction are called *factors*, those present in the final stage are called *products*. Thus in equation (3) those atoms and molecules at the left of the equality sign are factors, those at the right are products. Chemical equations are read from left to right, and are very rarely reversible.

**Quantitative Interpretation of Equations.** — Since matter cannot be destroyed, all weight entering a reaction can be accounted for. Much more information, however, lurks in an equation. It was found by Exp. 15 that when magnesium is heated in air, the ratio of the weights in which magnesium and oxygen combine is 1.5 : 1. Equation (1) might, therefore, be written —

$$\begin{array}{c} Mg + O = MgO \\ 1.5 + 1 = 2.5 \end{array}$$

Any number might be substituted for 1, provided, of course, the ratio 1.5 : 1 is not destroyed. Hence the above equation might be written —

$$\begin{array}{c} Mg + O = MgO \\ 24 + 16 = 40 \end{array}$$

This is the simplest kind of an equation, but the facts similar to those revealed here may be discovered about every equation. Equation (3) might be written —

$$\begin{array}{c} Zn + H_2SO_4 = H_2 + ZnSO_4 \\ 65 + \quad 98 \quad = 2 + \quad 161 \end{array}$$

Experiment 22 proved that 32.5 (approximately) gm. of zinc correspond to 1 gm. of hydrogen. If the work had been extended, it would have been found that 65 gm. of zinc correspond to 161 gm. of zinc sulphate. The equation would then become, as the result of experiment, —

$$\begin{array}{c} Zn + H_2SO_4 = H_2 + ZnSO_4 \\ \underbrace{65 + \quad 98}_{163} \quad = \underbrace{2 + \quad 161}_{163} \end{array}$$

Finally, in Exp. 16 it was shown that potassium chlorate yields about 39 per cent of oxygen. If we select for 3 O the number 48 (equal to 3 × 16), then we can form the proportion

$$48 : 39 :: x : 100$$

$$x = 123$$

If the experiment should be conducted with the utmost care and precaution, the proportion would be slightly modified, so that instead of equalling 123, $x$ would have equalled 122.5, a number nearer the true relation existing between potassium chlorate and the oxygen it yields. Equation (2) may now be written —

$$KClO_3 = 3\,O + KCl$$

$$122.5 = 48 + 74.5$$

From the above discussions it is evident that equations, accurately evolved and correctly interpreted, reveal the exact proportions by weight in which compounds react in the particular chemical change expressed by the equation. Other facts of fundamental importance are contained in chemical equations, and these will be considered in the proper place.

**Other Chemical Equations Studied.** — Experiment 12 was a qualitative experiment, but if the factors and products had been weighed, the final result might have been expressed thus: —

$$HgO = Hg + O \qquad (4)$$

$$216 = 200 + 16$$

This means not only that mercuric oxide decomposes into mercury and oxygen, but that 216 parts by weight of mercuric oxide produce 200 parts of mercury and 16 parts of oxygen, since careful experiments show that these numbers represent the correct proportions.

Again, when sulphur and charcoal were burned in oxygen in Exp. 14, the final results might be expressed thus: —

$$\begin{array}{c} S + O_2 = SO_2 \\ \underbrace{32 + (2 \times 16)}_{64} = \underbrace{32 + (2 \times 16)}_{64} \end{array} \qquad (5)$$

$$\begin{array}{c} C + O_2 = CO_2 \\ \underbrace{12 + (2 \times 16)}_{44} = \underbrace{12 + (2 \times 16)}_{44} \end{array} \qquad (6)$$

Equation (5) is correct, because many experiments prove that when 32 parts of sulphur are heated in oxygen, they combine with 32 parts by weight of oxygen, and, furthermore, the compound formed, called sulphur dioxide, yields by decomposition the same proportion of sulphur and oxygen by weight. Equation (6) is true for similar reasons. The experiment has been tried repeatedly, and the results show that carbon burning in an excess of oxygen always combines with 32 parts of oxygen for every 12 parts of carbon. The product of the reaction, called carbon dioxide, has been analyzed many times, and when analyzed always yields carbon and oxygen in the ratio of 12 to 32 parts by weight.

**Problems Based on the Foregoing Equations.** — In the experiments which led to the above equations, no special weight of matter was used. The equation simply states the proportions which participate in the reaction. Obviously any convenient weight of magnesium might be heated in the air, or any weight of potassium chlorate might be decomposed, if necessity or choice prevailed, but the proportions of the factors and products are always the same as the proportions which accurate experiments show are represented by the numbers given in the equation. Thus in equation (1),

$$\begin{aligned} Mg + O &= MgO, \\ 24 + 16 &= 40 \end{aligned}$$

if 12 gm. of magnesium are taken, then 8 gm. of oxygen will be necessary to convert all the magnesium into magnesium oxide, which will weigh 20 gm. — the sum of the weights of the two factors. Hence from the weight of one substance participating in a reaction we can calculate by proportions the actual weights of all other participants, if we know the equation corresponding to the reaction.

Suppose 17 gm. of magnesium are to be converted into magnesium oxide, and it is desired to know (*a*) how much oxygen is necessary and (*b*) how much magnesium oxide will be formed. The questions are answered by solving the following proportions: -

$$24 : 16 :: 17 : x.$$

$x = 11.3$ gm. of oxygen.

$$24 : 40 :: 17 : x.$$

$x = 28.3$ gm. of magnesium oxide.

Similarly, when 45 gm. of zinc interact with sulphuric acid, the weights of (*a*) acid required, (*b*) hydrogen formed, and (*c*) zinc sulphate produced are found by the proportions:—

$$65 : 98 :: 45 : x.$$

$x = 67.8$ gm. of sulphuric acid.

$$65 : 2 :: 45 : x.$$

$x = 1.38$ gm. of hydrogen.

$$65 : 161 :: 45 : x.$$

$x = 111.4$ gm. of zinc sulphate.

## PROBLEMS. IX.

1. How many grams of oxygen can be prepared from 122.5 gm. of potassium chlorate? From 245 gm.? From 421 gm.?

2. How much hydrogen can be prepared from 65 gm. of zinc? From 130 gm.? From 297 gm.?

3. How much zinc is needed to prepare 2 gm. of hydrogen? 4 gm.? 17 gm.?

4. How much zinc sulphate can be prepared from 98 gm. of sulphuric acid? From 196 gm.? From 427 gm.?

5. How much mercury and how much oxygen is obtained by heating 10 gm. of mercuric oxide?

6. How much oxygen can be prepared from 50 gm. of potassium chlorate?

7. A certain weight of potassium chlorate was heated and 298 gm. of potassium chloride remained. What weight of potassium chlorate was heated and what weight of oxygen was formed?

8. A certain weight of potassium chlorate was heated until completely decomposed, and the residue weighed 20.246 gm. What weight of potassium chlorate was heated, and how much oxygen was evolved?

9. If 60 gm. of mercuric oxide are completely decomposed, what volume of oxygen is obtained at 91° C. and 380 mm.?

10. If 400 gm. of potassium chlorate are completely decomposed, what volume of oxygen is obtained at 27° C. and 760 mm.?

11. If 100 gm. of zinc interact with sulphuric acid, what volume of hydrogen is evolved?

12. A balloon holds 132.74 kg. of hydrogen. How much zinc and sulphuric acid are needed to produce the gas?

13. How much potassium chlorate is needed to prepare 36.48 l. of oxygen at 15° C. and 750 mm.?

14. How many liters of oxygen can be obtained from a kilogram of potassium chlorate, if the gas is measured at 10° C. and 755 mm.?

15. How much zinc and sulphuric acid (in kg.) will yield hydrogen enough to fill a 350 cc. flask at 15° C. and 735 mm.?

16. If 12 gm. of carbon are burned in the oxygen obtained by decomposing 122.5 gm. of potassium chlorate, what weight of carbon dioxide is formed, and what weight of oxygen, if any, remains?

17. A lump of carbon weighing 24 gm. is burned in air.
    (*a*) What weight of carbon dioxide is formed?
    (*b*) What weight of oxygen is needed?
    (*c*) If a liter of oxygen weighs 1.43 gm., what volume of oxygen is needed?

18. What weight of carbon dioxide may be obtained by burning 112 lbs. of coal containing 15 per cent of impurities?

19. A sulphuric acid maker accidentally set fire to a storehouse containing 8794 kg. of sulphur. A rough estimate showed that 67 per cent had burned. What weight of sulphur dioxide was formed?

20. Calculate the weight of oxygen necessary to burn 731 gm. of sulphur containing 15 per cent of impurities.

21. A lump of sulphur weighing 32 gm. is burned in air. Calculate—
    (*a*) The weight of oxygen required.
    (*b*) The weight of sulphur dioxide formed.

# CHAPTER V.

## GENERAL PROPERTIES OF WATER.

WATER is the most abundant of all chemical compounds, and on account of its remarkable and varied properties it demands extensive study.

### GENERAL DISTRIBUTION.

**Experiment 24.** — *Supplies:* Test tubes, wood, meat, potato.

Heat successively in dry test tubes a small piece of wood, of meat, and of potato (or any fresh vegetable). Hold the test tube at such an angle that any moisture which may be liberated will not flow down upon the hot glass. Is there conclusive evidence of water? Is the amount unexpectedly large in any case?

These substances are types of animal and vegetable matter. Draw a general conclusion.

FIG. 35. — Potato showing (by shaded part) the proportion of water. Remainder is mainly starch and fiber.

FIG. 36. — Carrot showing (by shaded part) the proportion of water. Remainder is mainly starch, sugar, and fiber.

The water driven off by heating many organic substances may be simply mixed with other ingredients, or it may be due to the presence in them of hydrogen and oxygen in such proportions that these two gases combine to form water at the instant the compound is decomposed. Many examples of this latter kind are known and will be subsequently examined.

The wide distribution of water in vegetables is strikingly shown in Figs. 35 and 36. The human body is 70 per cent water. Water is essential to life.

**Water of Crystallization.** — Dry crystals deposited from the water solution of many solids often contain water, which seems to be an essential part of the chemical compound. This water is called *water of crystallization.* In some crystals the water passes off at the ordinary temperature, and in all crystals when they are heated.

### DETECTION OF WATER OF CRYSTALLIZATION.

**Experiment 25.** — (*a*) Heat a few small crystals of sodium carbonate or of alum in a dry test tube, inclining the test tube so that any water liberated may run out. What is the evidence that they contained water of crystallization? If there is any marked change in the appearance of the crystals, describe and explain it.

(*b*) Repeat, using a crystal of gypsum. Answer the question asked in (*a*).

(*c*) If time permits, examine successively crystals of the following compounds for water of crystallization, and describe the result in each case: Potassium nitrate, potassium dichromate, and magnesium sulphate.

(*d*) Heat two or three small crystals of copper sulphate in an evaporating dish which stands on a gauze-covered tripod. As the action proceeds, hold a dry funnel or glass plate over the dish. Is there conclusive evidence of escaping water of crystallization? Do the crystals change in color? In shape? Can the form of the crystals be changed by gently touching the mass with a glass rod? Continue to heat until the resulting mass is a bluish gray. Let the dish cool. Meanwhile heat a test tube two-thirds full of water. When the dish has cooled somewhat, pour the hot water slowly into the dish upon the copper sulphate. Explain the change in color, if any. If there are any lumps, crush them with a glass rod and heat the solution over the gauze until it is clear. Let the solution evaporate for several hours. Are crystals deposited? If not, heat a few minutes, and cool again; if so, why? Have they water of crystallization, and, if so, when did they get it?

The amount of water of crystallization in crystals is not arbitrary. It is constant in the same compound when crystallized under uniform conditions, but the amount varies between wide limits in different substances. Its necessity in some crystals and not in others, as well as its varying quantity, has never been explained.

### DETERMINATION OF WATER OF CRYSTALLIZATION IN BARIUM CHLORIDE.

**Experiment 26.** — *Method:* Heat to constant weight a weighed quantity of crystallized barium chloride in an uncovered crucible, and the loss will be the water of crystallization.

*Supplies:* Porcelain crucible and support, crystallized barium chloride, forceps, crucible block.

*Process:* Weigh accurately a clean, dry crucible. Weigh in it from 1.5 to 2 gm. of barium chloride, weighing *exactly* the amount taken. Record the results thus: —

| | Grams. |
|---|---|
| Weight of crucible and barium chloride . . | |
| Weight of crucible . . . . . . . . . . | |
| Weight of barium chloride . . . . . . | |

Stand the crucible on a pronged tripod or similar support, and heat with a low flame for about fifteen minutes, then gradually increase the heat, as in previous experiments, until the flame is just below the edge of the crucible. Continue the intense heat for ten minutes, then gradually lower the flame until it just touches the bottom of the crucible. Move the flame to and fro under the crucible, until all danger of cracking disappears. Stand the crucible in the crucible block, and weigh when cool enough to bear the hand. Record the results thus: —

| | GRAMS. |
|---|---|
| Weight of crucible and crystallized barium chloride . . | |
| Weight of crucible and barium chloride . . . . . . . | |
| Weight of water of crystallization . . . . . . . | |

To be sure that all water has been driven off, heat again strongly for five minutes, then cool and weigh, as before. If there is an appreciable loss, repeat until the last two weights are constant. If there is no appreciable loss upon the second heating, calculate the per cent of water lost.

Per cent of water lost =

Class average =

## PROBLEMS. X.

1. Calculate the per cent of water of crystallization in a sample of barium chloride, 1.5 gm. of which lost, on heating, 0.22 gm.

2. A student found that 2 gm. of barium chloride lost, on heating, 0.295 gm. What per cent of the sample was water of crystallization ?

**Definitions.** — Crystals which have lost water of crystallization are said to be **Dehydrated** or **Anhydrous.** Thus the bluish white product obtained by heating the blue crystallized copper sulphate is usually called "dehydrated copper sulphate." The words *dehydrated* and *anhydrous* have been extended to mean any substance from which water has been removed, as anhydrous alcohol or ether. The opposite term **Hydrated** is sometimes applied to a compound to emphasize the fact that it contains water of crystallization. The term **Dissociation** is applied to that special kind of decomposition in which the components may reunite under favorable conditions. If crystallized copper sulphate is heated, water, which is driven off by heat, readily reunites with the copper sulphate to form the blue crystallized compound, which, strictly speaking, is copper sulphate united with water, though it is loosely called copper sulphate. Crystallized copper sulphate, therefore, is said to dissociate when heated. There are many different varieties of this kind of chemical action.

**Efflorescence.** — Many crystals contain water of crystallization so loosely combined that the water passes off readily when the crystal is exposed to the air. This is called efflorescence, and the crystals are said to effloresce, or to be efflorescent.

### EFFLORESCENCE.

**Experiment 27.** — Put a fresh crystal of sodium carbonate and of sodium sulphate on a piece of filter paper and leave them exposed to the air for an hour or more. Describe any marked change. What does this change show about the air?

It is believed that the water in the crystals called efflorescent tends to pass off into the atmosphere. This tendency resembles the tendency of water to evaporate. And like water, this water of crystallization passes off only when the vapor tension of the crystal exceeds the tension of the water vapor in the atmosphere. Thus sodium sulphate at ordinary temperatures has a greater vapor tension than the atmospheric water vapor, but gypsum (crystallized calcium sulphate) has a lower vapor tension than the atmospheric water vapor. Hence sodium sulphate effloresces, but calcium sulphate does not. This is an easy way to distinguish the two compounds.

**Deliquescence.** — Many substances, crystallized and uncrystallized, absorb water when exposed to moist air, or often even to ordinary air, which always contains a little water vapor. This is called deliquescence, and such substances are said to deliquesce, or to be deliquescent.

### DELIQUESCENCE.

**Experiment 28.** — Place on a piece of glass or in an evaporating dish a small piece of calcium chloride, potassium hydroxide, sodium hydroxide, and potassium carbonate, and leave them exposed to the air for an hour or more. Describe any marked change which takes place. How does the action differ from that in Exp. 27? Where and why was this property of calcium chloride utilized?

Deliquescence is a property of substances very soluble in water. When such substances, potassium carbonate or calcium chloride for example, are exposed to the air, the water vapor forms with the substance a small quantity of a saturated solution. This saturated solution has a lower vapor pressure than that of the atmosphere, that is, the water is held by the substance, it does not tend to escape, hence more water vapor is added from the air, and finally the substance is entirely dissolved in this condensed vapor. Common salt or sodium chloride often appears to deliquesce, but the deliquescence is due to the very soluble magnesium and calcium chlorides which are usually mixed with commercial sodium chloride. Sodium nitrate is very soluble in water at the ordinary temperature, but potassium nitrate is only slightly soluble. Hence potassium nitrate, and not sodium nitrate, is used in the manufacture of gunpowder.

**Impure Water.** — Water, if impure, is usually contaminated with two kinds of impurities, — organic and inorganic. The former is, in general, decomposing animal and vegetable matter; the latter consists of various mineral substances dissolved from the earthy matter through which the water percolates. The term *impure* must be interpreted broadly, since many *mineral waters* are wholesome. "The rule generally adopted by chemists in dealing with water is to pronounce any water dangerous which is contaminated by sewage." (Remsen.)

### SIMPLE TESTS FOR IMPURITIES IN WATER.

**Experiment 29.** — (*a*) *Organic Matter:* Fill a clean test tube two-thirds full of distilled water and another with water containing a little dirt or a bit of paper. Add to each test tube a drop or two of concentrated sulphuric acid and sufficient potassium permanganate solution (made from distilled water) to color each liquid a light purple, as nearly alike as possible. Label one tube, and then heat gently nearly to the boiling point the tube containing the impure water. As soon as a definite change is seen, heat the other cautiously, as too sudden heat may cause the liquid to "bump out." Organic matter decolorizes potassium permanganate solution. Which tube shows the more organic matter?

(*b*) *Chlorides:* To a test tube half full of distilled water add a few drops of nitric acid, and then a few drops of silver nitrate solution. Do the same with faucet water or water known to contain a chloride in solution. What is the difference between the results? The cloudiness, or solid, is due to the precipitation of silver chloride, which is always formed when silver nitrate is added to hydrochloric acid or a chloride in solution. Silver chloride is soluble in ammonium hydroxide. Try it. This is the usual test for chlorides (and conversely for soluble silver compounds), and will hereafter be used without further description.

(*c*) *Sulphates:* Apply the usual test for sulphates to distilled water and to impure water known to contain sulphates (see Exp. 21 (*b*)).

(*d*) *Calcium Compounds:* Add a few drops of a fresh solution of ammonium oxalate to a test tube half full of clear lime water. Lime water is a solution of calcium hydroxide, and the white precipitate formed is calcium oxalate, which is soluble in hydrochloric acid but not in acetic acid. Try it. This is the test for calcium compounds, often called "lime" compounds, because lime, which is calcium oxide, is so well known. Apply this test to distilled water and to water known to contain calcium compounds, and compare the two results.

**Purification of Water.** — Water is purified by exposure to the air, by filtration, or by distillation. The last operation is convenient only with relatively small quantities, and is performed by means of a condenser, which is shown in Fig. 37 arranged for use. The condenser consists of an outer tube, *AA′*, provided with an inlet and outlet for a current of cold water, which surrounds an inner tube, *BB′*. These tubes are shown in Fig. 38. The vapor from the liquid boiling in the flask, *C*, condenses in the inner tube, owing to the decrease in temperature, and drops off the lower end of this tube, as the distillate, into the receiver, *D*. The condenser and receiver should, obviously, be scrupulously clean.

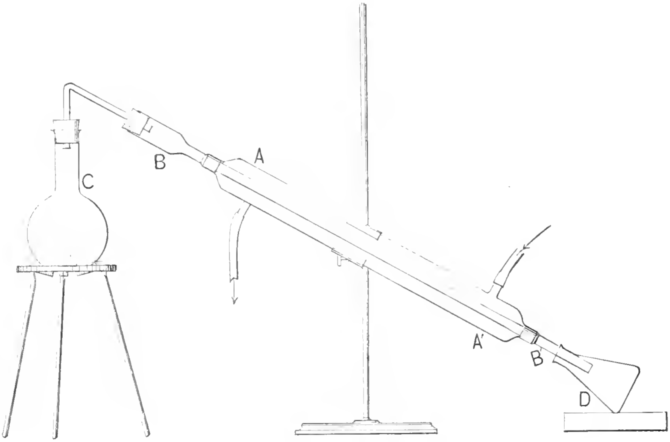

Fig. 37. — Condenser arranged for the purification of water by distillation.

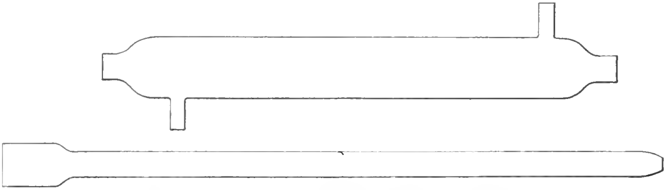

Fig. 38. — Inner and outer tubes of a condenser.

## DISTILLATION.

**Experiment 30.** — Fill the 500 cc. flask, *C*, half full of water known to contain the impurities mentioned in Exp. 29, add a few crystals (3 or 4) of potassium permanganate, and connect with the condenser as shown in Fig. 37. Attach the inlet tube to the faucet, fill the condenser slowly, and regulate the current so that a small stream flows continuously from the outlet tube into the sink or waste pipe. Heat the liquid in *C* gradually, and when it boils, regulate the heat so that the ebullition is not too violent. Collect about 25 cc. in the receiver, *D*. Test separate portions of this distillate for organic matter, chlorides, sulphates, and calcium compounds. Is organic matter found? Is mineral matter found? If the liquid in *C* had contained some volatile substance like alcohol, ammonia, or hydrochloric acid, where would this volatile matter have been at the end of the operation? Can water be separated from *any* solid by distillation? Sketch the apparatus used in this experiment.

## CLASS-ROOM EXERCISE. IX.

1. Additional study of water.
    - (*a*) Color, odor, and taste of pure water.
    - (*b*) Distribution of water in animal and vegetable matter.
    - (*c*) Rain water.
    - (*d*) Mineral waters: kinds, general character, medicinal value.
2. Water of crystallization.
    - (*a*) Relation to color and form of crystals.
    - (*b*) Additional examples.
    - (*c*) Theories.
3. Efflorescence and Deliquescence.
    - (*a*) Additional examples.
4. Drinking water.
    - (*a*) Organic matter: source, danger of its presence, how detected other than chemically, how removed.
    - (*b*) Significance of presence of sulphates and chlorides.
    - (*c*) Hard water: meaning of term, meaning of "temporary" and "permanent," how softened, how estimated, how detected other than chemically, why objectionable. What causes boiler scale? What causes the "furring" of a kettle?

(*d*) How is city water filtered? Of what value are large filters? Why should drinking water be exposed to the air?

(*e*) Why are lead compounds often found in water? How detected? Why dangerous?

(*f*) What is a "water analysis"?

5. Distillation.

(*a*) Literal meaning.

(*b*) Is it a recent process of purification?

(*c*) Is there any difference between water vapor and steam?

(*d*) Industrial applications.

6. Experimental.

Some substances have a moist surface and some have water mechanically enclosed within them. How would you distinguish this moisture from water of crystallization?

**Relation of Water to Temperature.** — The freezing point of water (or the melting point of ice) and the boiling point of water are the fixed points on a thermometer. They are easily found. They vary, however, under certain conditions.

### THE FREEZING AND BOILING POINTS OF WATER.

**Experiment 31.** — *Supplies:* Thermometer fitted loosely to a cork, iron stand and clamp, funnel, 250 cc. Erlenmeyer flask or similar vessel, tripod or iron ring, ice (or snow), salt.

(*a*) Crush the ice, if snow is not available, and put it in a funnel which stands in the flask. Clamp the thermometer around the cork, as shown in Fig. 39, and adjust the height so that the bulb is buried in the ice. After a short time, read the thermometer. Repeat the reading at brief intervals until the mercury is constant.

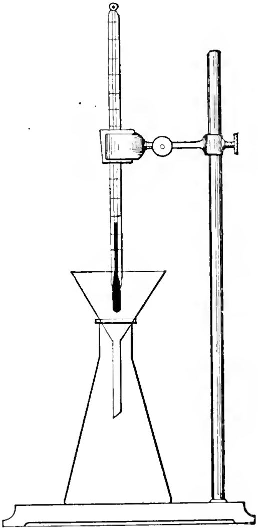

FIG. 39. — Apparatus for determining the melting point of ice.

Record the final reading. When the whole experiment is finished, sketch this apparatus in the note-book.

(*b*) Remove the thermometer, carefully transfer the ice to the flask, and fill the flask half full of water. Stand the flask on a gauze-covered tripod or ring. and clamp it loosely around the neck. Adjust the thermometer so that the bulb dips into the liquid. The mouth of the flask should be open. Read the thermometer. Heat gently and watch the thermometer. What happens? Repeat the readings until the mercury is constant. Record the final reading.

(*c*) Remove the thermometer and add from 10 to 20 gm. of sodium chloride — common salt. Adjust the thermometer as before, and find the boiling point of this solution. If not appreciably different, add more salt. Record the boiling point.

(*d*) Make a mixture of equal parts by weight of salt and crushed ice, and find, as in (*a*), the melting point of this mixture. Record the melting point.

(*e*) Summarize the results as follows: —

TABLE OF FREEZING AND BOILING POINTS.

| SUBSTANCE. | FREEZING POINT. | | | | BOILING POINT. | | | |
|---|---|---|---|---|---|---|---|---|
| | Found. | | Standard. | | Found. | | Standard. | |
| Water | C. | F. | C. | F. | C. | F. | C. | F. |
| | | | 0 | 32 | | | 100 | 212 |
| Water and Salt | | | Class Average. | | | | Class Average. | |
| | | | C | F. | | | C. | F. |
| | | | | | | | | |

## CLASS-ROOM EXERCISE. X.

1. Boiling.
   - (*a*) Exact definition.
   - (*b*) Effect of pressure on the boiling point.

- (*c*) Effect of a dissolved substance on the boiling point. Cause of this effect.
- (*d*) Other conditions affecting the boiling point.

2. Freezing.
   - (*a*) Exact definition.
   - (*b*) Effect of dissolved substance on the freezing point. Cause of this effect.
3. Miscellaneous.
   - (*a*) Explain the term *maximum density* of water.
   - (*b*) What is the scientific reason for using water to extinguish fires?
   - (*c*) What scientific measurements are based on water as a standard?
   - (*d*) State the changes in volume which occur when (1) ice melts. (2) water freezes, (3) water is heated from 0° C. to 15° C., (4) water is cooled from 15° C. to 0° C.

**Solution.** — Many solids, liquids, and gases disappear when put into water. This operation is called dissolving or putting into solution. The resulting liquid is called a solution of the substance used. The liquid in which the substance dissolves is called the **Solvent,** and the substance dissolved is called the **Solute.**[1] If the solute is not volatile, it may be recovered by evaporation or distillation of the solution. The degree of solubility is usually expressed by the terms *slightly soluble*, *soluble*, and *very soluble*. It is desirable, however, to state in all possible cases the exact proportions of solvent and solute. Substances which do not dissolve are called *insoluble*. A solution which contains a relatively small proportion of the solute is called a **Dilute** solution; one containing a relatively large proportion is called a **Concentrated** solution. Thus, dilute sulphuric acid contains usually one volume of sulphuric acid to three or more volumes of water, while concentrated sulphuric

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[1] See Glossary, Appendix C.

acid is nearly 98 per cent acid. The terms *dilute* and *concentrated* are loosely used in several senses, which will be clear as the experiments proceed. Other descriptive terms are applied to solutions, and these will be discussed in the proper place. Usually a solution means a solution in water, and the term is so used in this book, unless otherwise stated.

**Solutions of Gases.** — There is no general relation between the solubility of gases and their chemical composition. Their solubility varies between wide limits.

### SOLUBILITY OF GASES.

**Experiment 32.** — (*a*) Warm a little faucet water in a test tube. Is there *immediate* evidence of a previously dissolved gas? Is there evidence of much gas? What effect has increased heat?

(*b*) Warm slightly a few cubic centimeters of ammonium hydroxide in a test tube. Do the results resemble the observations in (*a*)? As soon as the result is obtained, pour the remaining liquid down the sink and flush well with water.

(*c*) Repeat (*b*), using a little concentrated hydrochloric acid. Do the results resemble those of (*a*) and (*b*)?

The volume of a moderately soluble gas which is dissolved by water is directly proportional to the pressure, if the temperature is constant. This relation is illustrated by the following: —

Table of Solubility of Carbon Dioxide Gas.

| Vol. of Water. | Vol. of Carbon Dioxide Measured under Normal Conditions. | Pressure. |
|---|---|---|
| 1 l. at 0° C. | 900 cc. | .5 atmosphere |
| | 1800 cc. | 1 atmosphere |
| | 3600 cc. | 2 atmospheres |
| | 7200 cc. | 4 atmospheres |

### CLASS-ROOM EXERCISE. XI.

1. What is "soda" water? Why should it be kept cold? Why kept in a strong vessel? Why so called?
2. Is oxygen soluble in water? Is hydrogen? Is air? (Consult a text-book regarding the solubility of hydrogen and of oxygen in water.)
3. What is carbonated water? Sulphur water?
4. Why are subterranean gases often easily dissolved by water?
5. Henry's Law of the solubility of gases.

**Solutions of Liquids.** — The solubility of liquids in water varies between wide limits. Some liquids, especially oils and organic liquids, are practically insoluble in water, and their insolubility is seen by the formation, after agitation, of two distinct layers of liquid.

### SOLUBILITY OF LIQUIDS.

*Alcohol, ether, carbon disulphide, and kerosene, which are used in this experiment, take fire easily. No flames should be near.*

**Experiment 33.** — (*a*) Measure into a test tube 10 cc. of water from a graduate or a pipette. If a pipette is used, proceed as follows: Rinse out the pipette with water, dip the pointed end into water and suck up the liquid nearly to the top, then remove the pipette from the mouth and quickly close the top with the forefinger, as shown in Fig. 40; now cautiously lessen the pressure of the finger enough to allow the liquid to sink slowly to the mark around the stem, then press down tightly, lower the end of the pipette into a dry test tube, and remove the finger. If it is a 10 cc. pipette, let all the water run out; if not, then let 10 cc. run into the test tube and reject the

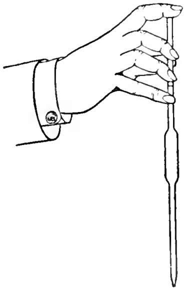

Fig. 40. — Using a pipette. This illustration shows the correct position of the hand after the pipette has been filled with liquid.

rest. Add to the water about half as much alcohol, and shake. Is there evidence of solution? Add a little more. and shake. Then add a third portion. Is there still evidence of solution? Draw a general conclusion regarding the solubility of alcohol in water.

(*b*) To 10 cc. of water add a few drops of ether. Observe the result and draw a conclusion. Add to the water about half its volume of ether, and shake. Observe the result.

From these two results make a general statement about the solubility of ether in water.

(*c*) Repeat (*b*), using successively kerosene, carbon disulphide, and glycerine. Observe the results and conclude accordingly.

Summarize the results as follows: —

TABLE OF SOLUBILITY OF SOME LIQUIDS.

| LIQUID. | SOLVENT. | |
|---|---|---|
| 1. Alcohol | 10 cc. | 1. |
| 2. Ether | of Water | 2. |
| 3. Kerosene | at Temperature | 3. |
| 4. Carb. disulph. | of | 4. |
| 5. Glycerine | Laboratory | 5. |

A rise of temperature in many cases increases the solubility of liquids in water.

**Solutions of Solids.** — The solubility of solids in water is a subject of vast practical importance, and recent investigations are giving it fundamental scientific value. Its principles can be mastered only after extensive study, but the following experiments will illustrate many of the fundamental facts.

### SOLUBILITY OF SOLIDS.

**Experiment 34.** — *Supplies:* About 20 gm. of powdered copper sulphate, 6 gm. of powdered potassium chlorate, 1 gm. of calcium sulphate.

(*a*) Label three test tubes I., II., III. Measure 10 cc. of water into each. To I. add 1 gm. of powdered copper sulphate, to II. add 1 gm. of powdered potassium chlorate, to III. add 1 gm. of calcium sulphate. Shake each test tube, and then allow them to stand undisturbed for a few minutes. Is there evidence of solubility in each case? Is there evidence of a varying degree of solubility? If III. is doubtful, carefully transfer a portion of the *clear* liquid to an evaporating dish by means of a pipette, or glass tube, and evaporate to dryness. Is there now conclusive evidence of solution? Draw a general conclusion from this experiment. Save solutions I. and II. for (*b*).

Tabulate the results of (*a*) as follows, using the customary terms to express the degree of solubility: —

TABLE OF SOLUBILITY OF TYPICAL SOLIDS.

| SOLUTE. | SOLVENT. | RESULTS. |
|---|---|---|
| 1. Copper sulphate | 10 cc. of Water at Temperature of Laboratory | 1. |
| 2. Potassium chlorate | | 2. |
| 3. Calcium sulphate | | 3. |

(*b*) Heat I. and add gradually 4 more grams of powdered copper sulphate. Does it all dissolve? Heat II. and add 4 more grams of powdered potassium chlorate. Does it all, or most all, dissolve? What general effect has increased heat on the solubility of solids? What is the difference between this general result and that in Exp. 32? Save the solutions for (*c*).

(*c*) Heat I. and II. nearly to boiling, and as the temperature increases add the respective solids. Do not boil the liquid away. If the volume of liquid becomes appreciably less, add a little water. Is there a limit to their solubility? Draw a general conclusion from these typical results.

**Saturation and Supersaturation.** — Some of the solutions obtained in Exp. 34 are **Saturated.** That is, the given volume of the liquid (or weight, since 1 cc. = 1 gm.) will dissolve no more solid at that temperature, although some of the solid remains undissolved. A saturated solution represents a state of stable equilibrium. There is no simple general relation between the temperature and the amount of solid dissolved. Since the maximum amount of solute usually depends upon the temperature, it is customary to say that a solution is saturated at such and such a temperature. Furthermore, the weight of the solvent must be given. Usually, solubility is stated as the number of grams soluble in 100 gm. of water at a certain temperature. Thus the number of grams of potassium nitrate which 100 gm. of water dissolves at —

0°C. is 13.3
20°C. is 31.2
50°C. is 85.0
100°C. is 246.0

In Exp. 34 an excess of undissolved solid was present in all cases when the solution was saturated. If, however, solutions can be brought to the point of saturation, either by cooling or by evaporation, frequently no solid will separate from the liquid. That is, solutions can exist which contain a larger amount of dissolved solid than is required for saturation at a given temperature. These solutions are called **Supersaturated.**

### SUPERSATURATED SOLUTIONS.

**Experiment 35.** — (*a*) Fill a test tube nearly full of crystallized sodium sulphate, and add not more than 5 cc. of water. Warm slowly. As solution occurs, heat gradually to boiling. Add sodium sulphate until no more will dissolve. Pour the solution into a warm, clean, dry test tube and let it stand until cool. Then drop in a small crystal of sodium sulphate and watch for any simple but definite change. What happens? Is the excess of solid large?

(*b*) Repeat with sodium thiosulphate, if time permits.

A supersaturated solution seems to be a case of unstable equilibrium.

## THERMAL PHENOMENA OF SOLUTION.

**Experiment 36.** — *Supplies:* Test tubes, thermometer, 5 gm. each of fused calcium chloride, potassium nitrate, ammonium nitrate, dehydrated copper sulphate, 1 gm. each of sodium hydroxide and potassium hydroxide, and a few cubic centimeters of concentrated sulphuric acid.

Measure 10 cc. of water into a test tube, take the temperature, add at one time 5 gm. of fused calcium chloride. As it dissolves, stir with the thermometer and observe the highest reading. Record in the proper place in the table below. Repeat successively with separate portions of water and the ammonium nitrate, potassium nitrate, dehydrated copper sulphate, potassium hydroxide, sodium hydroxide, and concentrated sulphuric acid (add the acid to the water). Tabulate the results as follows: —

SUMMARY OF THERMAL PHENOMENA OF SOLUTION.

| SUBSTANCES. | | SOLVENT. | | TEMPERATURE. | | THERMAL CHANGE. | |
|---|---|---|---|---|---|---|---|
| Name. | Amt. | Name. | Amt. | Initial. | Maximum or Minimum. | Maximum. | Kind. |
| 1. Calcium Chloride | | | | | | | |
| 2. Ammonium Nitrate | | | | | | | |
| 3. Potassium Nitrate | | | | | | | |
| 4. Copper Sulphate (dehy.) | | | | | | | |
| 5. Potassium Hydroxide | | | | | | | |
| 6. Sodium Hydroxide | | | | | | | |
| 7. Sulphuric Acid | | | | | | | |

In the act of dissolving, heat is necessary, hence solution is usually accompanied by a fall of temperature. If the dissolving body, however, combines with some of the liquid, heat is produced, as is the case in chemical combination. Sometimes the amount of heat produced is so large that it more than balances the slight fall of temperature. Thus when dehydrated copper sulphate is added to water, the dehydrated salt combines with water to form the blue compound, which is known to be copper sulphate combined with its water of crystallization. No doubt there was a slight fall of temperature due to the mere act of solution, but the heat due to the chemical combination is so much in excess that heat alone is observed as the result of solution. It has been proved that when sulphuric acid is dissolved in water, compounds of water and sulphuric acid — hydrates of sulphuric acid — are actually formed. So great is the heat produced with large quantities that the solution often boils, and sometimes so suddenly that the hot acid is spattered. Hence, the acid should always be poured slowly into water with constant stirring.

### SOLUTION AND CHEMICAL ACTION.

**Experiment 37.** — *Supplies:* A gram each of powdered tartaric acid, sodium bicarbonate, lead nitrate, potassium dichromate; mortar or evaporating dish, large vessel of water.

(*a*) Mix in a dry mortar or evaporating dish 1 gm. of powdered tartaric acid and an equal weight of sodium bicarbonate. Is there any decided evidence of chemical action? Pour the mixture into a large vessel of water. Is there conclusive evidence of chemical action?

(*b*) Repeat, using powdered lead nitrate and powdered potassium dichromate.

Describe the results in (*a*) and (*b*) and interpret the whole experiment from the standpoint of solution.

### CLASS-ROOM EXERCISE. XII.

1. Solvent power of water.
    - (*a*) Cleansing agent.
    - (*b*) Erosive agent.
    - (*c*) Solvent of drugs and medicines.
    - (*d*) Solvent of chemicals.
    - (*e*) Composition of sea water.
2. Relation of taste and solution.
3. Why are so many solutions used in the laboratory?
4. A liter of sea water was evaporated to dryness, and the residue weighed 36.4 gm. What per cent of the sea water was "salt"?
5. How is a cold saturated solution of sodium chloride influenced by rise of temperature?

# CHAPTER VI.

## COMPOSITION OF WATER.

WATER was thought to be an element until about the end of the eighteenth century. At that time its composition was shown by the efforts of Cavendish, Priestley, and Lavoisier, and, later, by Humboldt, Gay-Lussac, Nicholson and Carlisle, Dumas, and Davy. Their labors were so intermingled it will be impossible to follow their experiments chronologically. We shall, therefore, consider only the main facts which contributed to the discovery of the composition of water, together with some later work.

### DECOMPOSITION OF WATER BY ELECTRICITY.

Water was first decomposed by electricity in 1800 by Nicholson and Carlisle, and confirmed by Davy by a series of brilliant experiments extending through a period of six years. The decomposition of water by electricity is called **Electrolysis.**

### ELECTROLYSIS OF WATER.

**Experiment 38.** — Fill the Hofmann apparatus, Fig. 41, with water containing 10 per cent of sulphuric acid so that the water in the reservoir tube stands a short distance above the gas tubes after the stop-cock in each has been closed. Connect the platinum terminal wires with a battery. As the action proceeds, small bubbles of gas rise and collect at the top of each tube. Allow the current to operate until the height of the gas in the smaller tube is from 8 to 10 cm. Measure the height of each gas column. Assuming that the tubes have the same diameter, the volumes are in approximately the same ratio as their heights. How do the volumes compare?

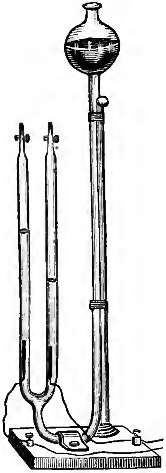

FIG. 41. — Hofmann apparatus.

Test the gases as follows: (*a*) Hold a glowing taper over the tube containing the smaller quantity of gas, cautiously open the stop-cock to allow the water (or air) to run out of the glass tip, and then let out a little gas upon the glowing taper. What is the gas? Repeat until the gas is exhausted. Care must be taken not to lose the gas by clumsy manipulation. It is advisable to have at hand several partially burned tapers or thin splints, in case the escaping water extinguishes the first one. (*b*) Open the other stop-cock long enough to force out the water in the glass tip; close the stop-cock, and slip a platinum tip (see App. A, § 13) over the glass tip. Open the stop-cock again, let out the gas slowly, and hold at the same time a lighted match at the end of the tip, then immediately thrust a taper into the small and almost colorless flame. What is the gas? Repeat until the gas is exhausted.

Draw a general conclusion from this experiment. Describe the whole experiment and sketch the apparatus.

### ELECTROLYSIS OF WATER. — *Optional Method.*

**Experiment 39.** — The Hofmann apparatus may be replaced by the one shown in Fig. 42. A large bottle, *A*, is cut into two parts by the method described in App. A, § 6. The upper part is used in this experiment and it is about 15 cm. deep. A bottle with a small mouth is preferable, though the one shown in the figure gives satisfactory results. Provide *A* with a stopper or cork having two holes. Through these holes pass two glass tubes into the upper ends of which are sealed platinum wires. The upper end of each wire is attached to a narrow strip of platinum foil by passing the wire in and out through several holes pricked by a pin along the longer diameter of the strip. The lower ends of these wires are attached to the battery by double connectors. When the stopper is fitted with its tubes, press it into the neck of the bottle so firmly that there is no crack between the neck and stopper. This is easily accomplished, if the neck is small and the stopper is rubber. Corks are difficult to fit, unless both cork and neck are perfectly circular. Place the bottle in the ring of an iron stand and pour in warm water to heat the glass. Meanwhile melt some paraffine and when the bottle is warm enough not to be cracked by the hot paraffine, pour out the water and pour in the paraffine slowly until it about fills the neck of the bottle. As it cools, press it around the tubes and where the glass and paraffine meet. When cold, pour water into the bottle to detect any leak. If the apparatus is not tight, add more paraffine. It may be necessary to adjust the position of the platinum strips when the test tubes, *C* and *D*, are lowered into the bottle. The electrolysis is accomplished the same as in Exp. 38. *A* is nearly filled with water containing 10 per cent of sulphuric acid, *C* and *D* are filled with the same solution and clamped over the platinum strips as shown in Fig. 42. When one tube is full of gas, the current is stopped, the volume measured as in Exp. 38, and each gas is tested with a lighted taper or blazing stick of wood.

What is the ratio of the volumes? What are the gases?

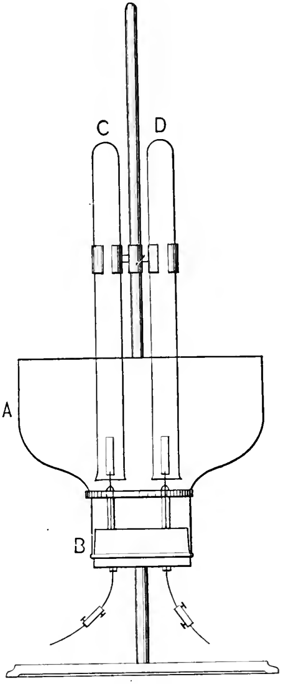

FIG. 42. — Apparatus for the electrolysis of water. — *Optional Method.*

It has been shown by many accurate trials of Experiment 38 (1) that only two gases are produced, viz., hydrogen and oxygen, (2) that the ratio of their volumes is 2 to 1, (3) that the sum of the weights of the products equals the weight of the water decomposed. It therefore follows that water is composed of hydrogen and oxygen combined in the proportion of 2 to 1 by volume. Additional experiments are necessary to prove the exact composition of water.

**Hydrogen and Water.** — The fact that hydrogen is a component of water, as shown in Exp. 38, may be verified by an experiment first performed by Lavoisier. He passed steam through a red hot gun barrel containing bits of iron. The solid product was, he said, "in the state of the black oxide precisely like that which had been [formed from iron when] burnt in oxygen." This experiment is historically interesting because it led to the name "hydrogen." Lavoisier says in his notes, "No name has appeared to us more suitable than that of hydrogen, that is to say, 'generative principle of water' — from *hudor* water, and *geinomai*, I produce."

### DECOMPOSITION OF WATER BY IRON.

**Experiment 40.** — *Method:* Pass steam over heated iron filings and collect the gaseous product over water.

*Apparatus:* An iron tube about 30 cm. long and 2 cm. in diameter, iron filings, steam generator, glass trap, pneumatic trough, three bottles (250 cc.), and stoppers and connectors.

*Process:* Fill the iron tube about half full of clean iron filings, hold it horizontal, and tap it gently on the table to provide a free channel for the gases. Support the tube in a furnace, or lay it on the rings of two iron stands. Provide each end with a one-hole rubber stopper. To one end attach a delivery tube which passes into a pneumatic trough. To the other end attach the steam generator. This is a 500 cc. flask half full of water and provided with a two-hole rubber stopper; through one hole passes a straight glass tube about 40 cm. long, open at both ends, and reaching to the bottom of the flask — to relieve any excessive back pressure; through the other hole passes a short glass tube bent at right angles and connected by rubber tubing with a glass trap,[1] which in turn is connected by rubber tubing with a

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[1] See Teachers' Supplement.

short glass tube projecting from the stopper of the iron tube. The trap collects any water driven over from the flask. The distance between the flask and iron tube must be as short as possible to prevent condensation of the steam before it reaches the iron tube. It is advisable to put in the flask several pieces of granulated zinc, pipe stem, or glass tubing, to insure steady boiling.

Fill the bottles with water and invert in the trough, but not over the hole in the shelf, since the first bubbles will be air. Heat the whole iron tube by means of one or more wing-top burners. Meanwhile heat the water in the steam generator, and by the time the steam has been formed, the iron tube will be sufficiently hot. If drafts cool the tube, protect it at the ends with asbestos boards. It is advisable, also, to keep the ends of the tube cool by hanging a piece of asbestos board from each end just inside the stoppers. Excessive heat may melt the stoppers. If they should melt and stop up the tube, the pressure gauge in the steam generator will indicate the difficulty. Reject the first portions of the gas. As soon as the hydrogen begins to be delivered, a few minutes will suffice to collect two bottles nearly full. Test the gas for hydrogen and for oxygen, applying as many different tests as the circumstances permit.

When the iron tube is cool, examine the contents. Compare it with the original iron and with magnetic oxide of iron.

Describe this experiment, stating exactly what it proves.

Other metals, zinc and magnesium, for example, will also decompose water, and hence furnish additional evidence of the fact that hydrogen is one component of water.

### CLASS-ROOM EXERCISE. XIII.

1. State exactly what Exp. 40 proves and suggests about the composition of water.

2. What does Exp. 23 prove and suggest about the composition of water?

3. If coal instead of iron had been used in Exp. 40, what industry would be illustrated? (Hint: look up "water gas.")

4. Summarize the evidence thus far presented about the composition of water.

5. Actual facts contributed by Davy to our knowledge of the composition of water.

**Oxygen and Water.**—The fact that oxygen is a constituent of water, shown in Exp. 38 and suggested in Exp. 23, may be verified by allowing chlorine and water to interact in the sunlight. Chlorine is a gas and an element, and later will be fully studied.

### DECOMPOSITION OF WATER BY CHLORINE.

**Experiment 41.**—Construct a chlorine generator and prepare chlorine as directed in Exp. 57. Pass the gas into a deep vessel of water until a sample of the liquid smells strongly of chlorine. The delivery tube should reach to the bottom of the vessel of water. Completely fill a flask with a slender neck with this saturated solution of chlorine, cork tightly, and stand in the sunlight. After several hours a small quantity of gas will collect at the top. Test the gas with a glowing match. Repeat the experiment, if the result is not satisfactory.

As the amount of gas is usually small, the apparatus shown in Fig. 43 gives more satisfactory results than the flask. It is a glass tube about 2 cm. in diameter and 1 m. long, closed at one end, either by sealing or by inserting a cork or solid stopper, and drawn out at the other end so that the smaller portion is about 4 cm. long and 1 cm. in diameter. The tube is filled with chlorine water, and the smaller end immersed in a vessel containing the same solution. When sufficient gas has been collected, cover the smaller end of the tube with the thumb, invert the tube, and test the gas, which will nearly fill the smaller portion of the tube.

FIG. 43.—Tube for the decomposition of water by chlorine.

Describe this experiment, stating what it shows about the composition of water. Sketch the apparatus used.

**Interaction of Sodium and Water.**—Additional facts concerning the composition of water may be obtained by a study of the interaction of sodium and water.

Sodium is an element. It is also a metal, though it is so soft it can be cut with a knife, and is much lighter in weight than those metals which are familiar. Nevertheless it has the luster characteristic of the metals, and chemically its behavior is markedly like the other metals. Sodium is kept beneath oil to protect it from the moisture of the atmosphere. *It should be used cautiously and strictly according to directions. Small fragments should not be left about or thrown into the refuse jar, but into a large vessel of water especially provided for the purpose.*

### INTERACTION OF SODIUM AND WATER.

*Remember the precautions to be observed in using sodium.*

**Experiment 42.**—*Supplies:* Sodium, pneumatic trough filled with water as usual, tea lead, test tube, forceps, litmus paper, iron stand and clamp.

(*a*) Scrape the brown coating from a piece of sodium, press it between filter paper to remove the excess of oil, cut off a piece not larger than a small pea, and drop it upon the water in the trough. Stand far enough away so that you can just see the action. Wait until you are sure the action has stopped, and then describe all you have seen.

(*b*) The action in (*a*) may be further studied as follows: Fill a test tube with water, invert it and clamp it in the trough so that the mouth is over the hole in the shelf of the trough. Wrap a small piece of sodium loosely in a piece of tea lead about 5 cm. square, make two or three small holes in the tea lead, and then thrust it under the shelf of the trough with the forceps. A gas will rise into the test tube. Proceed similarly with additional small pieces of sodium and dry tea lead until the test tube is nearly full of gas; then unclamp and remove, still keeping the tube inverted. Hold a lighted match, for an instant, at the mouth of the tube. Observe the result, watching especially the mouth of the tube. What is the gas? Why? Remembering that sodium is an element, where must the gas have come from? If there is any doubt about the nature of the gas, collect more, and subject it to those tests which will prove its nature.

(*c*) Put a piece of filter paper on the water in the trough, and before it sinks drop a small piece of sodium upon it. Stand back and observe the result. Wait for the slight explosion which usually occurs soon after the action stops. Describe all you have seen. What burned? What caused it to burn? To what is the vivid color probably due? (In answering these questions, utilize your knowledge (1) of the properties of the gases previously studied, and (2) of the usual accompaniment of chemical action, suggested here by the melting of the sodium.)

(*d*) Test the water in the trough with red litmus paper. Push the paper to the bottom or to the place where it is certain that chemical action between water and sodium has taken place. Test, until the red litmus paper has undergone a decided change in color. Describe this final result. With another piece of red litmus paper test a solution made by dissolving a small piece of sodium hydroxide in a test tube half full of water. Is the result similar? Dip the platinum test wire (see App. A, § 14) into this solution and hold it in the Bunsen flame. Describe the result. Is the color of this flame and that noticed in (*c*) the same?

Experiment 42 verifies the previous observation regarding the presence of hydrogen in water. It shows also that when water and sodium interact another substance is formed. This is sodium hydroxide. Its name suggests its components, for it is a compound of sodium, hydrogen, and oxygen. It belongs to a class of substances called alkalies, which will be considered later. These alkalies turn red litmus paper blue; and *sodium* hydroxide, like all sodium compounds, colors a Bunsen flame an intense yellow. The color is due to the volatilized sodium, and is the test not for hydroxides, but for sodium compounds. The sodium hydroxide is in solution in the trough, and by evaporating the water could be obtained, as a white solid, having all the properties of the sodium hydroxide actually used to make the solution in (*d*). Since sodium liberates hydrogen from water, and forms at the same time a compound also containing hydrogen, then the hydrogen in water must be divisible into two parts. If in Exp. 42 (*b*) the sodium had been weighed and its weight compared with the volume of hydrogen liberated, it would have been found that .1 gm. of sodium liberates 48.22 cc. of hydrogen. And if the sodium hydroxide thus formed had been dried and then heated with sodium, 48.22 cc. more of hydrogen would have been obtained. This shows that the hydrogen in water is divisible into two *equal* parts.

The facts revealed by the interaction of sodium and water, together with others which have been proved, permit us to express the reaction in its simplest form by the following equation:—

| $Na$ | + | $H_2O$ | = | $H$ | + | $NaOH$ | |
|---|---|---|---|---|---|---|---|
| Sodium | | Water | | Hydrogen | | Sodium Hydroxide | (7) |
| 23 | | 18 | | 1 | | 40 | |

This equation means that 23 gm. of sodium are needed to liberate 1 gm. of hydrogen from 18 gm. of water, and to form, at the same time, 40 gm. of sodium hydroxide.

**The Quantitative Composition of Water.**—The preceding experiments on the composition of water have been mainly qualitative. They have shown by analysis and synthesis that water is composed of hydrogen and oxygen, and that the ratio of their volumes is approximately 2 to 1. The most decisive evidence of the quantitative composition of water is obtained by the determination of (1) its exact volumetric composition, (2) its exact gravimetric composition, and (3) the density of steam. Volumetric means "by volume," and gravimetric means "by weight."

### VOLUMETRIC COMPOSITION OF WATER.

**Experiment 43.**—*Method:* Explode measured volumes of hydrogen and oxygen and measure the residual gas, using an excess of either gas.

*Apparatus:* The apparatus is shown in Fig. 44. *A* is a level tube about 90 cm. long, provided with a base, or fitted at one end with a one-hole stopper, to which is connected a piece of pressure tubing, *B* about 50 cm. long. When not in use, *A* is clamped in any convenient position. *F* is a eudiometer so arranged that the open end can be moved about in the reservoir, *E*. This reservoir is about 15 cm. deep and is made from a five-pint acid (or similar) bottle according to the directions given in App. A, § 6. It is supported by a ring or clamp, and is provided with a one-hole stopper, through which passes a short glass tube, *G*, the ends of which have a flange, made by melting each end in the Bunsen flame and pressing it quickly upon a hard surface. Before the flange is made, a rubber stopper, *C*, just large enough to fit the eudiometer, is pushed down firmly upon the other stopper. After the flange is made, the upper end of *G* is pulled down hard upon *C*, the tubing, *B*, is pulled over the lower end of *G* until it touches

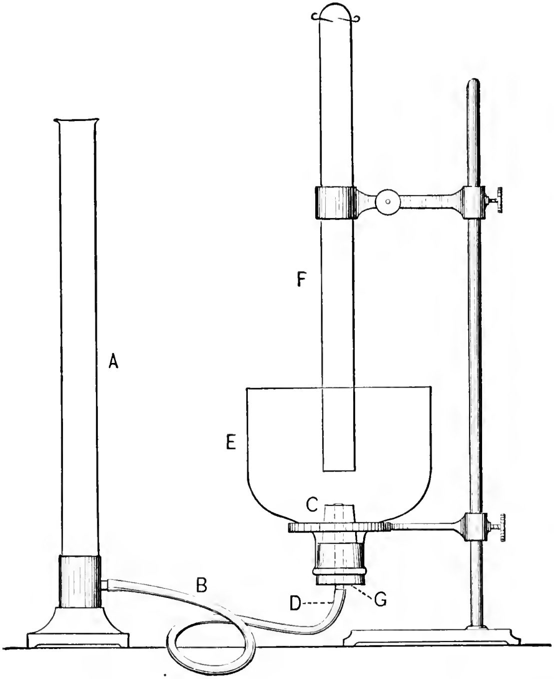

FIG. 44. — Apparatus for determining the volumetric composition of water.

the large stopper, and tied securely near the stopper and near the lower flange by shoemaker's waxed thread. A Hofmann screw is attached near the point *D*.

*Process:* Loosen the Hofmann screw, pour into *A* water that has been standing in the laboratory for at least twenty-four hours, until the reservoir, *E*, is two-thirds full; then tighten the screw and add more water until the level tube, *A*, is three-fourths full; clamp *A* at any convenient height to the iron stand which supports the apparatus. Pinch *B* to remove any imprisoned air. Fill the eudiometer, *F*, with some of the same water, invert in the reservoir, and clamp it so that the open end is just below the surface and near the wall of *E*.

Introduce into *F* about 10 cc. of oxygen free from air. The oxygen may be taken from a gas holder, or it may be prepared from the usual mixture in a test tube, since only a small volume is needed. Unclamp *F* and press it down upon the stopper, *C*; clamp *F* again, loosen the Hofmann screw, stand a thermometer in *E*, and allow the whole apparatus to remain undisturbed for at least ten minutes. Meanwhile, read the barometer, and write all necessary notes. Hold the top of the level tube near the eudiometer and raise or lower the level tube, until the water is at the same height in both *A* and *F*, then read the volume of oxygen in the eudiometer and the temperature of the water in *E*. Clamp *A* again in any convenient position. Record all readings as illustrated below.

Unclamp the eudiometer and lift it from *C*, and introduce into it from 25 to 30 cc. of hydrogen free from air. The hydrogen may be prepared from zinc and sulphuric acid in a test tube. Proceed as in the reading of the oxygen, observing the same precautions, especially the interval before reading the gas volume. This time, however, the eudiometer must be pressed down hard upon the stopper, *C*, taking care, of course, not to force the larger stopper out of place. Record the total volume and other readings as illustrated below.

Make sure that all stoppers and joints are tight, lower *A* to expand the gases in the eudiometer, clamp the Hofmann screw as tight as possible just below the glass tube, *G*, and clamp *A* as before. Explode the mixture of hydrogen and oxygen by passing an electric spark through the eudiometer. The spark may be obtained from a Ruhmkorff coil or plate machine. The explosion is indicated by a slight click, often accompanied by an instantaneous flash passing down the tube from the platinum wires to which the coil is connected. The ends of the platinum wires must be about 5 mm. apart in order to insure an explosion. After the explosion, loosen the screw and allow the water to seek its own level in the two tubes. Let the whole apparatus remain undisturbed for ten minutes or more, then read as before the gas volume, thermometer, and barometer.

A minute quantity of water is formed when the two gases unite, but so minute that its volume is not measurable. The residual gas can be proved by the usual tests to be hydrogen. The data obtained enable us to compute at once the volumetric composition of water.

*Calculation:* The recording of the data and calculation of the final result may be illustrated by an experiment actually performed by a student with this apparatus. Other determinations will vary somewhat from these figures, though the final result may be the same.

Data: —

I. Oxygen.

$V' = 10.7$ cc.
$t = 17°$ C.
$P' = 767$ mm.
$a = 14.42$ mm.

II. Hydrogen and oxygen.

$V' = 51.5$ cc.
$t = 17°$ C.
$P' = 767$ mm.
$a = 14.42$ mm.

III. Residue after the explosion.

$V' = 20.4$ cc.
$t = 17°$ C.
$P' = 767$ mm.
$a = 14.42$ mm.

These volumes, corrected (see App. B, formula 5) for temperature, pressure, and aqueous tension, become —

| | | |
|---|---|---|
| I. | Oxygen added | 9.97 cc. |
| II. | Hydrogen and oxygen added | 48.01 cc. |
| III. | Hydrogen left | 19.02 cc. |

Since 19.02 cc. of hydrogen remain, then the volume of hydrogen and oxygen which combined must be

$$48.01 - 19.02 = 28.99 \text{ cc.}$$

and volume of hydrogen which united with the oxygen must be

$$28.99 - 9.97 = 19.02 \text{ cc.}$$

If 19.02 cc. of hydrogen united with 9.97 cc. of oxygen to form water, then the ratio in which they combined was

19.02 : 9.97 or 1.9 : 1.

Accurate repetitions of this experiment have shown that the ratio in which the pure gases, hydrogen and oxygen, combine is 2 to 1, that is, the volumetric composition of water is two parts hydrogen and one part oxygen. The same result was obtained in Exp. 38, to be sure, but the result of Exp. 43, obtained by a more accurate method, establishes the fact beyond dispute.

### GRAVIMETRIC COMPOSITION OF WATER.

**Experiment 44.** — *Method:* Pass dry hydrogen over a weighed amount of copper oxide and collect the water formed in a weighed tube.

*Apparatus:* The construction and arrangement of the apparatus, except the generator, is shown in Fig. 45. The hydrogen may be generated as in Exp. 19, but it is difficult with this apparatus to obtain a slow evolution of gas as well as to be sure the generator will not be exhausted before the experiment is completed. If a gas holder or a reliable Kipp's apparatus is available, either one may be used. A simple, compact, and reliable generator is shown in Fig. 46. Two tubes, *A* and *B*, about 30 cm. long, 2 cm. in diameter, and open at both ends, are connected, as shown in the figure, by a piece of rubber tubing, *C*, about 40 cm. long. A Hofmann screw is attached at *E*. The generator is connected at *D* with the rest of the apparatus. The generator is operated as follows: Fill *B* two-thirds full of granulated zinc, separat-

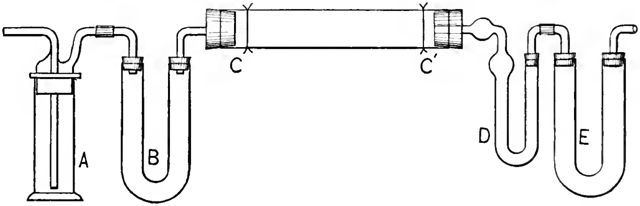

FIG. 45. — Apparatus for determining the gravimetric composition of water.

ing it from the bent tube at the bottom by a plug of cotton. Tighten the screw and fill *A* with dilute sulphuric acid. When the screw is loosened, acid will flow into *B*, and hydrogen will be evolved. It is advisable to clamp *B* into a permanent position, open the passage a little, and regulate the evolution of hydrogen by raising or lowering *A* by another clamp. A height will be found by trial which will produce the desired rate of evolution.

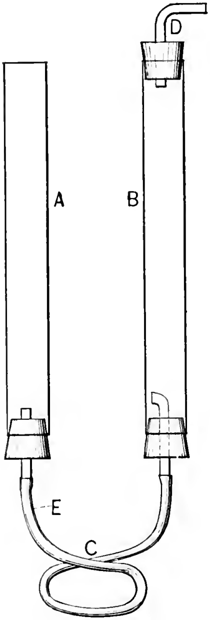

FIG. 46. — A generator for producing a steady current of hydrogen (and other gases to be subsequently studied).

In Fig. 45 *A* is a Drechsel drying bottle which should be half full of concentrated sulphuric acid, and *B* is a U-tube to be filled with calcium chloride — both serving to dry the hydrogen. *CC′* is a hard glass tube from 15 cm. to 20 cm. long, and from 1.5 cm. to 2 cm. in diameter. It is attached directly by rubber stoppers to *B* and *D*, the latter being a small (10 cm.) Marchand tube to be filled with fused calcium chloride. *E* is a U-tube to be filled with calcium chloride to prevent the moisture in the air from entering *D*. Only the tubes *CC′* and *D* are weighed before and after the experiment. At the points *C* and *C′* a copper wire (No. 20) is wound once or twice loosely around the combustion tube, and allowed to project at several points, like the spokes of a wheel, to prevent the tube from cracking and the stopper from melting at these points. This wire is *not* weighed with the tube. All joints must be tight. The apparatus may be supported by clamps at the necessary points, but not, of course, where they will interfere with the application of heat to *CC′*. Clean and dry the tubes *B* and *E*, and provide them with one-hole rubber stoppers, the holes of which are temporarily plugged with a short glass rod; fill each three-fourths full with lumps (not powder) of fused calcium chloride, push a plug of cotton down loosely upon the calcium chloride, wipe off any small particles of calcium chloride, and close the tubes with the plugged stoppers — removing the plugs, of course, when the tubes are in use. Clean and dry *D* and prepare it as follows: Push enough absorbent cotton down the straight limb and up the other limb to fill loosely the lower bulb, and then fill the tube with fused calcium chloride; nothing is put in the upper bulb; the straight limb is provided with a one-hole rubber stopper, which has previously been cleaned by boiling it successively in dilute sodium hydroxide, hydrochloric acid, and water. This stopper should have a short glass connector, bent at a right angle, and each end of the tube should be protected from the air by a short rubber tube plugged with a glass rod (see Fig. 47). The water which collects in the empty bulb during the experiment may be removed by a bit of rolled filter paper, and the tube thereby used several times without refilling. The stoppers fitted to $CC'$ should be cleaned, before use, by boiling as above described.

FIG. 47. — Glass plug.

*Process:* Clean and dry the combustion tube $CC'$ and put in it about 20 gm. of granulated copper oxide which has been heated in an iron or porcelain dish to remove moisture and organic matter. Introduce the copper oxide by the method illustrated in App. A, § 11, and hold it in place by two loose plugs of shredded asbestos which has been heated red-hot by the blast lamp. It may be heated in a porcelain dish, if the dish is heated and cooled gradually. Weigh the combustion tube and contents. Weigh the Marchand tube, $D$, *without the protectors.* Record the weights as shown below. Connect the apparatus as shown in Fig. 45, joining $CC'$ directly to $B$ and $D$, after the copper wires have been slipped over the ends of the combustion tube. Attach the generator, and pass a slow current of gas through the whole apparatus to drive out the air and incidentally to detect any leak. If there is no leak, then the gas will *not* bubble through $A$, when the open end of $E$ is covered for an instant with the finger. If there is a leak, *stop it before proceeding.* Regulate the flow of hydrogen so that about three bubbles a second pass through $A$. An ordinary Bunsen burner and a wing-top burner are needed, the latter for constant use, the former for an emergency. Heat the *whole* combustion tube slowly at first with the wing-top burner, gradually increasing the heat where the copper oxide is located. As the heat increases, moisture will collect in $CC'$. Heat the combustion tube from $C$ toward $C'$ so that the moisture will not collect at $C$, but be driven along beyond the copper oxide toward $C'$ and over into the bulb of $D$. The point $C$ must be kept hot, just as hot as possible without melting the stopper; if the moisture collects here, it is difficult to remove. The moisture often collects in the oxide, hence this must be kept hot enough from the outset to volatilize the water. If drops gather at $C'$, heat this point cautiously with a low Bunsen flame. The wing-top burner may be held at various angles to $CC'$, thereby allowing heat to be easily directed upon short spaces. When the copper oxide ceases to glow, lessen the heat somewhat and drive over into $D$ any moisture seen, or *thought to be*, in the combustion tube. Allow the hydrogen to run while the apparatus is cooling. When $CC'$ is cool enough to handle, disconnect the generator and draw air through the apparatus by applying the lips at the open end of $E$. Disconnect $CC'$, wipe out any organic matter from the inside at the ends, and then weigh. Disconnect $D$ and weigh *as before*. Record the weights as shown below.

*Calculation:* The form of recording the data and calculating the final result may be illustrated by figures taken directly from a student's note book. Other determinations of the gravimetric composition of water may, of course, give entirely different figures, though the final results may agree.

Data:—

I. Combustion tube.

| | Grams. |
|---|---|
| Weight of combustion tube and copper oxide before heating | 65.688 |
| Weight of combustion tube and copper oxide after heating | 64.783 |
| Weight of oxygen lost | 0.905 |

II. Marchand tube.

| | |
|---|---|
| Weight of Marchand tube after experiment | 76.741 |
| Weight of Marchand tube before experiment | 75.720 |
| Weight of water formed | 1.021 |

But 0.905 and 1.021 are in the same ratio as 8 and 9 (approximately).

Thus $$0.905 : 1.021 :: 1 : x$$

$$x = 1.128$$

But $$1 : 1.128 :: 8 : 9.024$$

That is, oxygen makes up $\frac{8}{9}$ of water, and the remaining $\frac{1}{9}$ is hydrogen.

*Final directions:* Protect from the air the whole drying apparatus by connecting it "tandem" and plugging the ends with the rod or rubber plug, as the case demands. Protect *D* with its rubber plugs. Pour the acid from the generator, wash the zinc and rubber tube free from acid, and recharge, or preserve empty, for future use.

Numerous chemists, whose work is reliable, have found this proportion ($\frac{8}{9}$) to be approximately correct.

**Density of Steam.** — The density of steam is the ratio of the weight of a given volume of steam to an equal volume of hydrogen under the same conditions. It could easily be calculated if we knew the weight of a liter of steam under given conditions of pressure and temperature, because the weight could then be compared directly with the weight of a liter of hydrogen. It has been found that .8063 gm. is the weight which a liter of steam would have, if it could exist as steam at the normal temperature and pressure. The determination is carried on, of course, at a temperature above 100° C., and the volume reduced to standard conditions, which are simply arbitrary points, selected irrespective of the physical states of matter. Dividing the weight of a liter of steam by the weight of a liter of hydrogen, we have —

$$\frac{.8063}{.0896} = 9$$

The significance of this fact will appear later, when the subject of molecular weights is studied. Just now all that can be said is that the molecular weight of a gas is twice its density. Hence 18 is the molecular weight of steam.

**Summary.** — The following facts have been revealed, partly by our experiments, and partly by suggestion, in the study of the composition of water: —

1. Water is a chemical compound of hydrogen and oxygen.

2. It is formed by burning hydrogen in the air, or by exploding a mixture of hydrogen and oxygen.

3. It can be decomposed by electricity into hydrogen and oxygen in the proportion of two volumes of the former to one volume of the latter.

4. Sodium liberates hydrogen from water, and forms at the same time a solid containing hydrogen equivalent to the hydrogen liberated. Iron also liberates hydrogen from water.

5. Chlorine liberates oxygen from water.

6. Two volumes of hydrogen and one volume of oxygen combine to form water, and the weight of the water formed equals the weight of the gases.

7. Water is formed by the union of two parts by weight of hydrogen with sixteen parts by weight of oxygen.

8. The molecular weight of water is 18.

The foregoing facts permit us to write the equation

$$\begin{array}{ccccc} H_2 & + & O & = & H_2O \\ \text{Hydrogen} & + & \text{Oxygen} & = & \text{Water} \\ 2 & + & 16 & = & 18 \end{array} \qquad (8)$$

as the simplest expression of the reaction for the formation of water from hydrogen and oxygen.

The significance of the density of oxygen (see page 50) is now apparent, at least in part, since 16 is the smallest number which can be used in the above equation. The meaning of the result of the determination of the gravimetric composition of water (see page 98) is likewise clearer. All facts point to the truth of equation (8).

### CLASS-ROOM EXERCISE. XIV.

1. Actual facts contributed to our knowledge of the composition of water by—
    - (*a*) Cavendish.
    - (*b*) Dumas.
    - (*c*) Gay-Lussac.

2 The formula of water.
    - (*a*) Why $H_2O$?
    - (*b*) Why not HO, as formerly?
    - (*c*) Why not $H_2O_2$?

3. Which of the following are oxidation and which reduction :—
    - (*a*) Burning of hydrogen.
    - (*b*) Formation of iron oxide from steam and iron.
    - (*c*) Formation of magnesium oxide by heating magnesium in air.

4. Symbol of Sodium, Iron, Copper, Hydrogen, Oxygen, Chlorine, Platinum.

5. Formula of water, steam, ice, sodium hydroxide, copper oxide, magnetic oxide of iron.

6. Define gravimetric, volumetric, eudiometer, ratio, density.

7. Sources of error in Exps 43 and 44, and how avoided.

8. Litmus.
    - (*a*) Source.
    - (*b*) Preparation of commercial form.
    - (*c*) Uses.

**Volumetric Composition of Steam.** — If water is decomposed, and the mixed gases are exploded when the surrounding temperature is above 100° C., then the product is gaseous water, that is, steam. If the experiment is performed quantitatively, it is found that three volumes of the mixed gases form two volumes of steam. The importance of this fact will appear later, when similar facts regarding the two gases, hydrochloric acid and ammonia, are revealed. The simple relation existing between the volumes of the components and resulting gas is expressed by the equation —

$$\begin{array}{ccccc} 2\,H_2 & + & O_2 & = & 2\,H_2O. \\ \text{2 Volumes of Hydrogen} & + & \text{1 Volume of Oxygen} & = & \text{2 Volumes of Steam} \\ 4 & + & 32 & = & 36 \end{array} \qquad (9)$$

This equation means that 4 gm. of hydrogen require 32 gm. of oxygen to form 36 gm. of steam.

**Reduction.** — The chemical change in the combustion tube in Exp. 44 is called **Reduction.** It usually means the removal of oxygen, and is, therefore, the opposite operation from oxidation. In Exp. 44, most, perhaps all, of the copper oxide was reduced to metallic copper, that is, oxygen was removed. Hydrogen is therefore a reducing agent, just as oxygen is an oxidizing agent. Reduction is an important chemical operation, especially industrially, since many ores are reduced from the oxide to the metal.

### PROBLEMS. XI.

1. What weight of hydrogen and of oxygen can be obtained from 125 gm. of water?

2. If 10 gm. of water are decomposed by electricity, what weight and what volume of hydrogen and of oxygen are formed?

3. If 216 gm. of water are decomposed by electricity, how many liters of oxygen are formed?

4. How many grams of water are formed by burning 5 l. of hydrogen?

5. 10 l. of hydrogen are to be burned in oxygen.
   (*a*) What volume of oxygen is necessary?
   (*b*) What weight of water will be formed?

6. What volume of oxygen is necessary to unite with 100 gm. of hydrogen?

7. A piece of sodium weighing 150 gm. is allowed to interact with water. Calculate —

(*a*) The weight and volume of hydrogen liberated.
(*b*) The weight of sodium hydroxide formed.

8. A gram of sodium amalgam (a mixture of sodium and mercury) liberated from water 200 cc. of hydrogen at 13° C. What per cent of sodium did the amalgam contain?

9. How many centigrams of hydrogen will be liberated when 23 cg. of sodium interact with water? How many centigrams, if the sodium had contained 8 per cent of impurities?

10. How many grams of hydrogen can be obtained by using 3 gm. of sodium? How many cubic centimeters?

11. Hydrogen is passed over 2.48 gm. of copper oxide, which at the end of the experiment weighed only 2.24 gm.; the water formed weighed 0.27 gm. In what ratio did the hydrogen and oxygen combine?

12. Calculate the ratio in which hydrogen and oxygen combine to form water from the following data: —

| | GRAMS. |
|---|---|
| Weight of copper oxide tube before the experiment . | 105.6 |
| Weight of copper oxide tube after the experiment . . | 101.6 |
| Weight of drying tube after the experiment . . . . | 84.8 |
| Weight of drying tube before the experiment . . . | 80.3 |

13. Calculate the ratio in which hydrogen and oxygen combine to form water from the following data, which were obtained by Dumas: —

| | GRAMS. |
|---|---|
| Weight of copper oxide tube before the experiment . | 334.598 |
| Weight of copper oxide tube after the experiment . | 314.236 |
| Weight of drying apparatus after the experiment . . | 449.263 |
| Weight of drying apparatus before the experiment . | 426.358 |

14. What weight of copper remains after 159 gm. of copper oxide have been reduced by hydrogen?

15. If 20 gm. of hydrogen are passed without loss through a tube containing hot copper oxide, what weight of water is formed?

16. Berzelius and Dulong, in 1820, obtained the following data in three determinations of the gravimetric composition of water: —

Loss of copper oxide tube in grams, respectively, 8.051, 10.832, and 8.246.

Weight of water formed in grams, respectively, 9.032, 12.197, 9.27. Calculate in each case the ratio of combination. What is the average result?

17. Dumas and Stas, in 1843, repeated the work of Berzelius and Dulong, and as an average of nineteen determinations found that 840.161 gm. of oxygen formed 945.439 gm. of water. Calculate the ratio of combination.

# CHAPTER VII.

## THE ATMOSPHERE.

THE atmosphere is the gaseous envelope surrounding the earth and extending into space. It is often called *the air*. The word *air* was used by the earlier chemists in the sense in which the word *gas* is now employed. The words *atmosphere* and *air* are often used interchangeably. The study of the atmosphere has occupied the time of many skilful chemists, and the progress of science has been hastened by their labors, especially by their investigations of the action of air on living things, its relation to combustion, the influence of its weight, and the proportion and interaction of its constituents. Boyle (1626–1691) began these investigations, and they are still in progress. A new impetus has recently been given to the work by the discovery of argon in the atmosphere by Rayleigh and Ramsay. This gaseous element, which had hitherto escaped detection, forms about $\frac{1}{120}$ part of the atmosphere.

**Composition of Air.** — It has already been shown that air contains oxygen and at least one other gas. It is now a favorable time to study this other constituent and the proportions in which these two gases exist in the air. These two objects may be simultaneously accomplished by a

### QUANTITATIVE EXAMINATION OF AIR.

*Phosphorus is used in this experiment. Before touching the phosphorus, read carefully the precautions stated below, and* **observe them.**

**Experiment 45.** — *Method:* Remove the oxygen from a measured volume of air by phosphorus, and measure the residual gas.

*Apparatus:* The arrangement of the apparatus is shown in Fig. 48. *A* is a tall jar, *B* is a 100 cc. graduated tube, *C* is a piece of worm-shaped phosphorus attached to a long piece (40–60 cm.) of copper wire, *D* is a clamp to support *B*.

*Process:* Fill *A* two-thirds full of water which has been exposed to the air in the laboratory for at least twenty-four hours. Fill the tube, *B*, with the same water, cover the end with the thumb or finger, hold it over the jar, and allow enough water to run out so that when the tube is lowered into the jar the 100 cc. mark is a short distance below the surface, if the water is at the same height within and without the tube. Clamp the tube in this position, stand a thermometer in the water in the jar, and allow the whole to remain undisturbed for about fifteen minutes. While the air is coming to the temperature of the water over which it is enclosed, prepare the phosphorus. Phosphorus is an element which oxidizes so quickly in the air that it often burns unexpectedly. To prevent its oxidation it is kept under water. *It is dangerous to handle, since the burns made by it are deep and painful.* Transfer a piece of phosphorus about 5 cm. long with the forceps from the bottle to a large porcelain mortar, battery jar, or metal dish nearly full of water. Wind the copper wire around it in the manner shown in Fig. 48. Quickly, but deliberately, transfer the phosphorus to the jar of water, *A*.

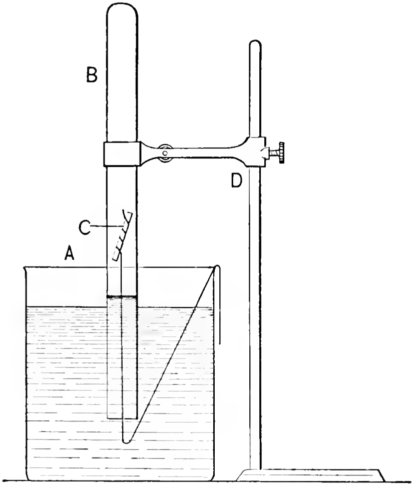

FIG. 48.—Apparatus for the quantitative examination of air.

Without touching the tube with the hand (see App. B, II., (3) and (4)), adjust it so that the water is at the same level within and without the tube. Read the volume of air in the tube, the thermometer, and barometer. Record as shown below.

Push the phosphorus up into the tube just above the surface of the water. White fumes indicate immediate action. In about ten minutes push the phosphorus half-way up the tube, attach the wire to the jar as shown in the figure, and allow the whole to remain undisturbed for an hour or more. Then pull the phosphorus down into the jar, *still keeping it under water*, and gently rub off with the fingers any deposit on the surface of the phosphorus. Insert the cleaned phosphorus into the tube as before, and if the white fumes appear, let the phosphorus remain a half-hour longer. If no fumes appear, then lower the phosphorus into the jar and let it remain there, taking care to keep the copper wire below the surface (to avoid accidental removal of the phosphorus from the jar). Lower the tube, *without touching it with the hand*, until the water is about the same level within and without, stand the thermometer in the water, and let the whole remain undisturbed for fifteen minutes. Then readjust the tube, if necessary, so that the inner and outer levels are the same, read the volume of gas, the thermometer, and barometer. Record as shown below. Unclamp the tube, grasp it near the end with the hand, close the open end with the thumb or finger, lift it from the water, invert, and test the residual gas several times with a glowing stick. Observe carefully its action with a flame, and record the observations. The name of this gas is **nitrogen**.

*Calculations :*

Data : —

I. Before the removal of oxygen.

Volume of air taken $= V' =$
Temperature $= t =$
Pressure $= P' =$
Aqueous tension $= a =$

Reduce this volume to the volume it would occupy under standard conditions by the formula

$$V_1 = \frac{V'(P' - a)}{760(1 + .00366 \times t)}.$$

Corrected volume $= V_1 =$

**II.** After the removal of oxygen.

Volume of gas remaining = $V'$ =
Temperature = $t$ =
Pressure = $P'$ =
Aqueous tension = $a$ =

Reduce to standard conditions as in I.

Corrected volume = $V_2$ =

From $V_1$ and $V_2$ find the percentage of oxygen and nitrogen in air. Insert the answer in the proper place in the

### TABLE OF THE COMPOSITION OF AIR.

| CONSTITUENT. | PER CENT BY VOLUME. | | |
|---|---|---|---|
| | Found. | Class Average. | Correct. |
| Oxygen . . . . . . . . | | | 20.96 |
| Nitrogen (and Argon) . . | | | 79.04 |

### LABORATORY EXERCISE. V.

1. Upon what property of phosphorus does the chemical action in Exp. 45 depend?

2. What became of the white fumes seen in Exp. 45?

3. Summarize the most obvious properties of nitrogen.

4. Compare the properties of nitrogen with those predicted in Laboratory Exercise II.

### CLASS-ROOM EXERCISE. XV.

1. Atmospheric nitrogen.
   - (*a*) Discovery and distribution.
   - (*b*) Physical functions in the atmosphere.
   - (*c*) Chemical relation to plants and animals.
   - (*d*) Its direct assimilation by plants.
   - (*e*) How prepared in large quantities.

2. Atmospheric oxygen.
   (*a*) Physical and chemical functions in the atmosphere.
3. Experimental.
   (*a*) Give several methods for showing that air contains oxygen.
   (*b*) Suppose air is passed over red-hot copper, what are the products? If the experiment were conducted quantitatively, what would it prove?
4. Historical.
   (*a*) Who first proved the gravimetric composition of the atmosphere? Describe the apparatus and state the method.
   (*b*) Bunsen's work on the volumetric composition of the air.
   (*c*) Lavoisier's work on the composition of the air.
   (*d*) Aristotle's conception of the word *air*.
5. Miscellaneous.
   (*a*) Sources of error in Exp. 45.
   (*b*) Precautions in handling phosphorus.

The atmosphere is not entirely composed of oxygen and nitrogen.

## OTHER CONSTITUENTS OF THE ATMOSPHERE.

**Experiment 46.** — (*a*) Prove by an experiment that water vapor is present in the atmosphere.

(*b*) Construct an apparatus like that shown in Fig. 49. Fill the test tube half full of lime water and draw air through the apparatus by means of an aspirator bottle or filter pump. (See App. A, § 17.) If no pump or aspirator is available, expose lime water in a bottle or beaker to the air. Allow the operation to continue, in either case, until a definite change is produced in the appearance of the lime water. In the aspirated lime water the change will appear throughout the liquid; in the exposed lime water look upon the surface for the change. If neither operation is satisfactory, use a concentrated solution of barium hydroxide in place of lime water.

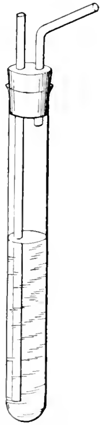

FIG. 49.—Apparatus for detecting the presence of carbon dioxide in the air.

The change is due to the formation of calcium carbonate. Lime water is the popular name of a solution of calcium hydroxide, and when it is exposed to carbon dioxide, the two interact. The chemical change is represented in its simplest form by the equation—

$$\underset{\substack{\text{Calcium}\\ \text{Hydroxide}}}{Ca(OH)_2} + \underset{\substack{\text{Carbon}\\ \text{Dioxide}}}{CO_2} = \underset{\substack{\text{Calcium}\\ \text{Carbonate}}}{CaCO_3} + \underset{\text{Water}}{H_2O} \qquad (10)$$

The formation of this easily detected calcium carbonate is the conclusive test for carbon dioxide. Usually a glass rod which has been dipped into lime water is held in the escaping gas, and if the lime water turns milky, this is a proof of the presence of carbon dioxide. Barium hydroxide acts the same as calcium hydroxide, and is sometimes used instead of the latter.

### CLASS-ROOM EXERCISE. XVI.

1. Atmospheric water.
    - (*a*) Source.
    - (*b*) Relation to plants and animals.
    - (*c*) What are clouds? What is dew, rain, frost, snow, hail?
    - (*d*) Amount present in air on a pleasant day.
    - (*e*) What is meant by the "dew point" of air?
2. Atmospheric carbon dioxide.
    - (*a*) Source.
    - (*b*) Relation to plants and animals.
    - (*c*) Carbon dioxide is heavier than air. Why, then, is the former gas not always at or near the earth's surface?
    - (*d*) Proportion in the air?
    - (*e*) How determined quantitatively?
3. Miscellaneous.
    - (*a*) Name the gases occasionally found in the atmosphere; source of each.
    - (*b*) Name those rarely found.

(*c*) "The air is a mixture." Give in your own language three reasons for this statement.

(*d*) If the nitrogen were entirely removed from a vessel of air, would the physical properties of the oxygen be changed?

(*e*) If air is a mixture, why do we speak of its "composition"?

(*f*) Chemical effect of lightning on the atmosphere.

### WEIGHT OF A LITER OF AIR.

**Experiment 47.** — *Method:* Find the weight and volume of air exhausted by a pump from a bottle.

*Apparatus:* The apparatus is shown in Fig. 50. *A* is a large bottle (2 liters or more) provided with a one-hole rubber stopper, through which passes a glass tube, *B*, bent as shown in the figure. *C* is a piece of pressure tubing (40–50 cm. long). A Hofmann screw is attached at the point *D*. It is advisable to make a flange on the outer end of the glass tube by melting it in the Bunsen flame and pressing it quickly upon a hard surface. The stopper must fit air tight, and a *little* vaseline should be rubbed on that part of the stopper which enters the neck of the bottle. If a filter pump is used to exhaust the bottle, a straight calcium chloride tube, filled as usual, should be placed between the pump and the bottle.

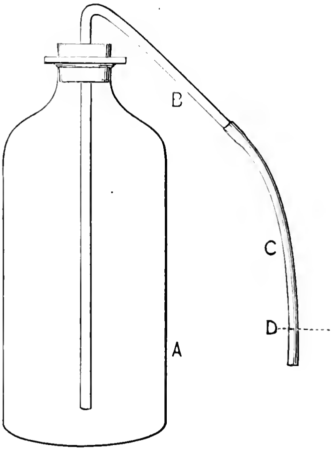

FIG. 50. — Apparatus for determining the weight of a liter of air.

*Process:* Slip *C* over the end of *B*, and unless it fits exceedingly tight, tie it just above the flange with waxed thread. Attach the screw loosely at *D*, push in the stopper as far as possible, wipe the bottle with a piece of cheese cloth or soft paper, and weigh the whole apparatus on the trip scales. Stand the bottle in the center of the left-hand scale pan and arrange the weights symmetrically around the center of the other. Weigh to a decigram, and leave the weights on the scales ready for the second weighing, so that the same large weights may be used each time. This precaution will eliminate many of the errors due to weighing. Call this weight $a$, and record as shown below.

Slip the end of the rubber tube over the exhaust nozzle of an air pump, and exhaust the air by rapid, short strokes of the handle. Tighten the Hofmann screw so that it will be air tight. The screw should not be more than five centimeters from the end. Detach from the pump, and weigh again — on the same scales. If less than 1 gm. of air has been removed, exhaust again. Call this weight $b$, and record as shown below.

Lower the rubber tube into a jar of water, open the clamp cautiously *under water*, and allow the water to run in slowly. The sudden rush of water into the bottle is easily controlled by pinching $C$ near the end between the thumb and forefinger. The jar can be kept full by allowing water to flow continuously through a tube reaching to the bottom of the jar. Care must be taken to keep the jar free from air bubbles by adapting the inflow to the outflow. When no more water runs into the bottle, raise or lower the bottle until the water is at the same level in both jar and bottle, pinch the tube and remove it from the jar. Tighten the screw, loosen the stopper, wipe the whole apparatus dry, and weigh again, observing the previous precautions. Record this weight as $c$, as shown below.

*Calculation:*

Data: —

| | | Grams. |
|---|---|---|
| $a$ | Weight of bottle before exhaustion = | |
| $b$ | Weight of bottle after exhaustion = | |
| $c$ | Weight of bottle and water = | |

Therefore

$a - b$ = Weight of air removed = gm.

$c - b$ = Volume of air removed = cc.

Therefore

$$c - b : a - b :: 1000 : x$$

$$x = \quad \text{gm.}$$

Summary of results: —

WEIGHT OF A LITER OF AIR.

| FOUND. | CLASS AVERAGE. | WEIGHT UNDER STANDARD CONDITIONS. |
|---|---|---|
| 1. | | |
| 2. | | |
| 3. | | 1.293 |
| Average, | | |

### LABORATORY EXERCISE. VI.

1. Draw a general conclusion from the result of Exp. 47.

2. Tabulate the weights of a liter of hydrogen, oxygen, and air, stating the conditions under which the weights were obtained.

**Weight of a Liter of Nitrogen.** — A liter of nitrogen under standard conditions weighs 1.256 gm.

### CLASS-ROOM EXERCISE. XVII.

1. Atmospheric pressure.
    (*a*) Meaning.
    (*b*) How measured.
    (*c*) What is "an atmosphere"?
    (*d*) Why does it vary?
    (*e*) Why is it less on a stormy than on a pleasant day?
2. Liquid air.
    (*a*) Early experiments.
    (*b*) Dewar's work.
    (*c*) Tripler's process.
    (*d*) Properties.
    (*e*) Industrial uses.
3. Historical.
    (*a*) Discovery and properties of argon.
    (*b*) Discovery and properties of helium.
    (*c*) Discovery and properties of "recently announced constituents of the air"

## PROBLEMS. XII.

1. A liter of dry air is passed over heated copper. The copper increases in weight 0.297 gm. What per cent of oxygen by weight is contained in this sample of air?

2. A mixture of 25 cc. of air and 50 cc. of hydrogen is exploded, and the residue measures 60.3 cc. What per cent of oxygen was contained in this sample of air?

3. Dumas, in determining the composition of air by passing air over heated copper and measuring the residual nitrogen, obtained the following data: –

| | Grams. |
|---|---|
| Weight of tube and copper before experiment | 120.00 |
| Weight of tube and copper after experiment | 121.15 |
| Weight of globe exhausted | 852.00 |
| Weight of globe and nitrogen | 855.85 |

Calculate the per cent of each constituent by weight and volume.

4. Bunsen analyzed the air at Marburg in 1846 by exploding a measured volume with a measured volume of hydrogen. Calculate the per cent of each constituent from the following analyses: —

| | | |
|---|---|---|
| I. | Air used | 428 93 cc. |
| | Air and hydrogen used | 749.77 cc. |
| | Residue | 480.09 cc. |
| II. | Air used | 448.00 cc. |
| | Air and hydrogen | 743.01 cc. |
| | Residue | 461.72 cc. |

5. A student placed a piece of phosphorus in a volume of air which, when corrected, was 87.9 cc. The residue when corrected was 70.58 cc. Calculate the per cent of each constituent. Is the result high or low? How much?

6. If the composition of normal air by weight is 76.995 per cent of nitrogen and 23.005 per cent of oxygen, what is the per cent of each constituent by volume?

7. Dumas and Boussingault in 1841 found in a sample of air 12.373 gm. of nitrogen and 3.68 gm. of oxygen. What per cent of each constituent did they find?

8. What weight of hydrogen is necessary to unite with the oxygen in 100 gm. of air to form water?

9. Assuming that air contains exactly 23 per cent of oxygen by weight, how many grams of carbon are necessary to unite with the oxygen in 100 gm. of air to form carbon dioxide? (Hint: see equation (6).)

10. How many kilograms of air are necessary to yield 100 kg. of oxygen?

11. 20 cc. of air are mixed with 30 cc. of hydrogen and the mixture is exploded. What volumes of what gases remain?

12. The simplest equation for the reaction between phosphorus and oxygen is —

$$2\,P + 5\,O = P_2O_5$$

How much phosphorus is necessary to remove the oxygen from a liter of air?

13. Calculate in the following cases the weight of a liter of air ·

| | | |
|---|---|---|
| (*a*) | Weight of air exhausted . . . . . . . . . . | 3.4 gm. |
| | Volume of air exhausted . . . . . . . . . . | 2571.6 cc. |
| (*b*) | Weight of air exhausted . . . . . . . . . . | 1.3 gm. |
| | Volume of air exhausted . . . . . . . . . . | 990.6 cc. |

If the results are high or low, state how much.

14. Calculate the weight of the following volumes of air: 5 l., 8.5 l., 10 l.

# CHAPTER VIII.

## ACIDS, BASES, AND SALTS.

EXTENSIVE study of the properties of chemical compounds has led to a general, though incomplete, classification of many of them into three grand divisions, long known as Acids, Bases, and Salts. Each division has its distinctive properties, though the divisions are closely related and occasionally overlap. The old limits of the classification and its distinguishing terms have not been modified to meet the new compounds. The nomenclature of these compounds is antiquated. "Even to-day each of these words (acid, base, and salt) is encumbered with a residue of ancient usage from which it is almost impossible to set it free." The present chapter is, therefore, a preliminary study, not from choice, but from necessity.

### GENERAL PROPERTIES OF ACIDS.

**Experiment 48.** *Supplies:* Test tubes, dilute sulphuric, nitric, and hydrochloric acids, acetic acid, glass rod, litmus paper (both colors), zinc, calcium carbonate (marble), lime water.

Fill a test tube one-third full of one of the following dilute acids: sulphuric, hydrochloric, and nitric; fill another test tube one-third full of concentrated acetic acid. Label the tubes in some distinguishing manner.

(*a*) Dip a *clean* glass rod into each acid and cautiously taste it. Describe the taste by a single word.

(*b*) Dip a clean glass rod into each acid and put a drop on both kinds of litmus paper. The striking change is characteristic of acids; draw a general conclusion from it.

(*c*) Slip a small piece of zinc into each test tube successively. If no chemical action results, warm gently. Test the most obvious product by holding a lighted match at the mouth of each tube. If no decisive action results, provide the test tube with a stopper and simple delivery tube, and collect any product in a test tube over water. This latter method will probably be unnecessary except with the acetic acid. Draw a general conclusion from the result in the case of all except the nitric acid.

(*d*) Wash the test tubes, add fresh acid, and slip a piece of calcium carbonate (marble) into each tube successively. Proceed as in (*c*), only in this case the gas must be tested not only with a burning match, but as follows: Dip a clean glass rod into lime water, and hold it in the escaping gas. What is the gas? (Hint: recall a similar experiment under air). Draw a general conclusion from these results.

The results in Exp. 48 are characteristic of all acids, and substances acting thus are said to be *acid*, or to have *acid properties*, or to have an *acid reaction*. The action in (*b*) is true only for litmus, but it is a striking, simple test, and should be remembered. In (*c*) nitric acid seems exceptional. It is not, however, and its action with metals will be explained under Nitric Acid (see page 197). Any metal could be used in place of zinc. The result in (*d*) is true of all carbonates, though the action, as in the case of metals, varies in vigor.

An acid may be a solid, liquid, or gas, or an aqueous solution of each. It may be dilute or concentrated, and the general properties vary somewhat with the strength of the acid. Each acid has special properties, which will be considered under the acids themselves.

**Nomenclature of Acids.** — Oxygen is a constituent of most acids, and the names of these acids correspond to the proportion of oxygen which they contain. The best known acid of an element usually has the suffix *-ic*, *e.g.* sulphuric, nitric, phosphoric, etc. If the element forms another acid containing less oxygen, this acid has the suffix *-ous*, *e.g.* sulphurous, nitrous, chlorous, etc. Some elements form an acid containing less oxygen than the *-ous* acid; these acids retain the suffix *-ous*, and have, also, the prefix *hypo-*, *e.g.* hyposulphurous, hyponitrous, hypochlorous, etc. If an element forms an acid containing more oxygen than the *-ic* acid, such acids retain the suffix *-ic*, and have, also, the prefix *per-*, *e.g.* persulphuric, perchloric, periodic, etc.

Acids which do not contain oxygen have the prefix *hydro* and the suffix *-ic*, *e.g.* hydrochloric, hydrobromic, etc. This last class is small.

The nomenclature of acids is well illustrated by the series of chlorine acids:—

### ACIDS OF THE ELEMENT CHLORINE.

| NAME. | FORMULA. |
|---|---|
| Hydrochloric | $HCl$ |
| Hypochlorous | $HClO$ |
| Chlorous | $HClO_2$ |
| Chloric | $HClO_3$ |
| Perchloric | $HClO_4$ |

The series is incomplete in the case of many elements, but the names are formed according to the above principles.

### GENERAL PROPERTIES OF BASES.

**Experiment 49.**—*Supplies:* Test tubes, litmus paper (both kinds), glass rod, short glass tube, solid sodium hydroxide and potassium hydroxide; ammonium hydroxide and calcium hydroxide.

Make dilute solutions of sodium hydroxide and potassium hydroxide by adding a very small piece to a test tube half full of water, provide, also, an equal quantity of dilute ammonium hydroxide and of calcium hydroxide.

(*a*) Rub a little of each liquid between the fingers, and describe the feeling. *Cautiously* taste each liquid by touching to the tip of the tongue a rod moistened in each, and describe the result.

(*b*) Test each liquid with both kinds of litmus paper. Describe the most striking result. Compare with the action of acids on litmus paper.

(*c*) Blow through a glass tube into each liquid, except the ammonium hydroxide, for several minutes. Describe the change, if any. Then add to each test tube in succession a few drops of dilute hydrochloric acid and test the gas evolved for carbon dioxide by holding in each tube a glass rod with a drop of lime water on the end. The amount of gas may be small, hence the rod must be held near the surface of the liquid as soon as possible after the acid has been added. Explain the chemical change.

The results in Exp. 49 are characteristic of bases, and substances acting thus are said to be *basic*, to have *basic properties*, or to have a *basic reaction*. The word *alkaline* is often used instead of basic. The result in (*a*) is typical, though it varies in intensity between wide limits from mere unpleasantness to violent caustic action. The results in (*b*) are generally true, though there are important exceptions. This is the customary test for a base. The result in (*c*) is not usually significant, though it serves to emphasize the relation which acids and bases bear to carbonates.

**Nomenclature of Bases.** — The term *base* is exceedingly broad in meaning and loose in application. Most bases are *hydroxides*, since they consist of hydrogen and oxygen combined with a metal. *Hydrate* is sometimes used as a synonym of hydroxide, both words suggesting the relation of these compounds to water. Hydroxides are usually called *alkalies*, and their chemical action is called *alkaline*. There is no general rule covering the nomenclature of bases. Hydroxides are distinguished from each other by placing the name of the metal before the word hydroxide, *e.g.* sodium hydroxide, potassium hydroxide, calcium hydroxide, barium hydroxide. The common hydroxides have long been known by several different names. Thus sodium hydroxide is called sodium hydrate, or caustic soda; ammonium hydroxide is called ammonia, or ammonia water.

### A GENERAL PROPERTY OF SALTS.

**Experiment 50.**—*Supplies:* Test tubes, litmus paper (both colors), glass rod, sodium chloride, potassium nitrate, potassium sulphate, and barium chloride.

Make, as in the preceding experiment, dilute solutions of sodium chloride (common salt), potassium nitrate, potassium sulphate, and barium chloride. Test each solution with litmus paper and describe the result. Compare with the action of acids and of bases on litmus paper.

This result is true only of salts which are called normal salts (see below). It is also characteristic of bodies which are, chemically speaking, not salts, such as water. Substances, whether salts or not, which act toward litmus like common salt, are said to be *neutral*, or to have a *neutral reaction.* A salty or saline taste is characteristic of most salts, and this property doubtless led to their classification with sodium chloride as a type, since this compound has been known for ages. Some compounds which are salts as far as their chemical relations are concerned do not have a neutral reaction. Thus sodium carbonate, which is the sodium salt of carbonic acid, is one of the most marked alkalies, being, in fact, known in commerce simply as "alkali."

**Nomenclature of Salts.**—Salts containing only two elements follow the rule for binary compounds, and hence end in *-ide.* This suffix is added to an abbreviated form of the name of the non-metal, *e.g.* chloride, bromide, fluoride, etc. Notice that the prefix *hydro-* is omitted, and that the name of the metal precedes, *e.g.* sodium chloride, potassium bromide, calcium fluoride, etc. It will soon be shown experimentally that salts may be regarded as derived from acids by replacing the hydrogen of the acid by a metal. Only the ending of the acid is changed to indicate this relation. Thus —

*-ic* becomes *-ate.*
*-ous* becomes *-ite.*

Hence —

Sulphuric acid forms *sulphates.*
Sulphurous acid forms *sulphites.*
Nitric acid forms *nitrates.*
Nitrous acid forms *nitrites.*
Chloric acid forms *chlorates.*
Hypochlorous acid forms *hypochlorites.*
Permanganic acid forms *permanganates.*

The name of the replacing metal is retained, *e.g.* potassium sulphate, sodium nitrite, calcium hypochlorite, potassium permanganate. Notice that the prefixes, *hypo-* and *per-*, are not changed.

Salts in which all the hydrogen is replaced by a metal are called *normal salts.* Sometimes only a part of the hydrogen is replaced, and the salt still has acid properties; such salts are called *acid salts, e.g.* acid sodium sulphate is the salt derived from sulphuric acid by replacing half of the hydrogen by sodium; its formula is $HNaSO_4$; the formula of normal sodium sulphate is $Na_2SO_4$. On the other hand, if a salt is so made that it still retains a portion of the base, it is called a *basic* salt, *e.g.* basic nitrate of bismuth may be regarded as the salt derived from bismuth hydroxide, in which only a portion of the base is replaced by the essential part of nitric acid; its formula is $Bi(OH)_2NO_3$. Basic lead nitrate, $Pb(OH)NO_3$, is formed by boiling solutions of lead nitrate and lead oxide. Basic salts bear the same relation to hydroxides as the acid salts to acids. They are prepared in numerous ways, and are often complex compounds. Thus basic lead carbonate (white lead) has the formula $2PbCO_3 . Pb(OH)_2$, and a basic copper carbonate (malachite) has the formula $CuCO_3 . Cu(OH)_2$.

**Testing the Nature of Substances.** — The presence or absence of an acid or base may be found by the litmus test, but such a test reveals only the properties of the substance examined. Other tests must be made to determine whether or not a particular substance is an acid, base, or salt.

## THE NATURE OF COMMON SUBSTANCES.

**Experiment 51.** — Determine by the litmus test the nature of lemon juice, vinegar, soap, sweet and sour milk, washing soda, borax, wood ashes, faucet water, baking soda, sugar, cream of tartar, the juice of any ripe fruit and any green fruit.

Make a solution of each of the solids before testing. Tabulate the results as follows: —

NATURE OF COMMON SUBSTANCES.

| ACID. | ALKALINE. | NEUTRAL. |
|---|---|---|
| | | |

### LABORATORY EXERCISE. VII.

1. Predict the reaction of "acid phosphate," "soda water," sour bread, basic lead acetate, sour wine, tart preserves, pickles.

2. Determine experimentally and then explain the reaction of (*a*) copper sulphate, (*b*) ammonium chloride, (*c*) alum, (*d*) potassium carbonate. (If necessary, consult a text-book for the explanation.)

**The Nature of Salts.** — The relation of salts to acids and bases is found experimentally by a study of —

### THE INTERACTION OF ACIDS AND BASES.

**Experiment 52.** — Dissolve a small piece of sodium hydroxide in an evaporating dish half full of water. Slowly add dilute hydrochloric acid, until a drop taken from the dish by means of a glass rod reddens blue litmus paper. Then evaporate to dryness by heating over a piece of wire gauze. Since the residue mechanically holds traces of the excess of hydrochloric acid added, it is necessary to remove this acid before applying any test. Heat the dish until all the yellow color disappears, then moisten the residue carefully with a few drops of warm water and heat again to remove the last traces of acid. This precaution is essential to the success of the experiment. Test a portion of the residue with litmus paper to find whether it has acid, basic, or neutral properties. Taste a little. Test (*a*) a solution of the residue for a chloride, and (*b*) a portion of the solid residue for sodium. Draw a definite conclusion from the total evidence.

**Neutralization.** — The operation in Exp. 52 is called *neutralization.* Other acids and bases produce similar results. It shows qualitatively the relation existing between acids and bases, and the relation of salts to each of the other two classes. Further information on the subject of neutralization may be obtained by determining the quantitative relation between acids and bases when representatives of the two classes interact.

## A QUANTITATIVE EXAMINATION OF THE INTERACTION OF ACIDS AND BASES.

**Experiment 53.** — *Method:* Neutralize a measured quantity of an acid with a base.

*Apparatus:* Two burettes and a support, two beakers of the same size, two glass rods (blunt form).

*Solutions:* (*a*) Add 8 cc. of dilute hydrochloric acid (1 vol. of acid to 3 vols. of water) to 200 cc. of filtered water. Keep the solution in a clean, stoppered flask or bottle which is labelled Hydrochloric Acid.

(*b*) Make a solution of dilute sulphuric acid (1 vol. of acid to 5 vols. of water) as in (*a*). Label it Sulphuric Acid.

(*c*) Dissolve about 2 gm. of sodium hydroxide in 200 cc. of filtered water. Preserve as described above, and label it Sodium Hydroxide.

(*d*) Make a solution of potassium hydroxide as in (*c*). Label it Potassium Hydroxide.

(*e*) Dissolve 1 gm. of solid phenolphthalein in 100 cc. of alcohol (50 per cent — see Table X., App. C). Keep in a glass stoppered bottle. This solution is to be used instead of litmus. It colors alkaline solutions magenta, and undergoes no appreciable change in acid solutions. This solution or any other used similarly is called an *indicator*, because it indicates the nature of a solution, whether acid or alkaline.

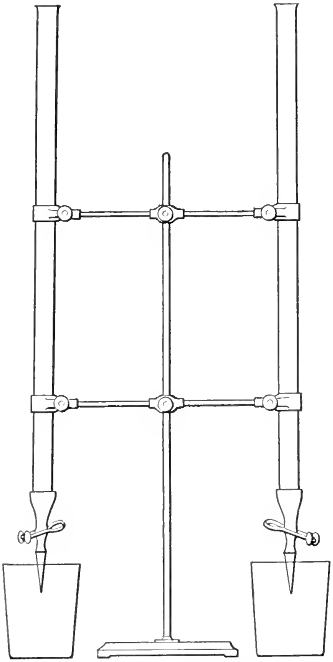

FIG. 51. — Burettes arranged for use.

*Process:* Wash two burettes with distilled or filtered water, and dry the inside with a stick having a piece of soft paper attached to one end. Arrange them as shown in Fig. 51. Label one clamp "acid" and the other "alkali." Fill the burette marked "acid" with the hydrochlo acid already prepared, taking care that the liquid is free from air bubbles at the lower end, and finally that the surface of the acid is level with the zero mark. Remember that the *lowest* point of the meniscus is the correct position. The burette holds 50 cc. from the 0 mark to the 50 mark, but it should be filled to the glass tip, and drawn off from 0 to 50 — but no lower. Any number of cubic centimeters may be obtained by opening the pinchcock and allowing the liquid to flow out of the glass tip until the level has fallen to the desired point. Fill the burette marked "alkali" with the sodium hydroxide solution in the same way. When the operation is over, temporarily or finally, the burettes should be washed and left standing full of clean water.

See that each burette reads 0, then place a beaker on a sheet of white paper under the "acid" burette, open the pinchcock, and allow 15 cc. of acid to run into the beaker. Add two or three drops of phenolphthalein solution. Then allow the sodium hydroxide solution to run slowly, drop by drop, into the acid. Stir constantly with the glass rod. As the action proceeds, the magenta color will be more permanent; add the alkali very slowly, as this point is approached. The aim is to add just enough alkali to neutralize the acid. It is customary to pass *slightly* beyond this point, just far enough to be sure the color is permanent. It requires skill and judgment to stop the action at the right time. Record as shown below the exact number of cubic centimeters of sodium hydroxide necessary to neutralize the 15 cc. of hydrochloric acid. Draw into another beaker 15 cc. of acid and find the exact amount of alkali necessary to neutralize it. Do the same with 10 cc. of acid, refilling the "alkali" burette, if necessary. The tint should be the same in each case. Tabulate the results thus: —

I. Hydrochloric acid and sodium hydroxide.

15 cc. of acid neutralize      cc. of alkali.
15 cc. of acid neutralize      cc. of alkali.
10 cc. of acid neutralize      cc. of alkali.
Ratio: [1] Acid, 3–3–2. — Alkali, – – .

If the work has been performed with reasonable accuracy, the results in I. will show: —

(1) Definite quantities of acid neutralize definite quantities of alkali.

(2) Neutralizing power is independent of the amount of solution used.

---

[1] Same as 15–15–10.

Wash the burette used for the alkali, and, after drying it, fill with the potassium hydroxide solution. Refill the "acid" burette with hydrochloric acid, and proceed with these solutions as before. Tabulate the results thus: —

II. Hydrochloric acid and potassium hydroxide.

15 cc. of acid neutralize      cc. of alkali.
15 cc. of acid neutralize      cc. of alkali.
10 cc. of acid neutralize      cc. of alkali.
Ratio: Acid, 3–3–2. — Alkali, – – .

These results likewise reveal the facts stated in I. (1) and (2). They also give (3) the ratio of the neutralizing power of the two alkalies, viz., — to —. ((3) is found by comparing the quantity of each alkali necessary to neutralize a definite quantity — 15 cc. for example — of acid.)

Wash and dry the "acid" burette, and fill with the sulphuric acid. Refill the "alkali" burette with the potassium hydroxide solution. Proceed as before, using successively 15, 15, and 10 cc. of acid and the necessary amount of alkali. Tabulate the results thus: —

III. Sulphuric acid and potassium hydroxide.

15 cc. of acid neutralize      cc. of alkali.
15 cc. of acid neutralize      cc. of alkali.
10 cc. of acid neutralize      cc. of alkali.
Ratio: Acid, 3–3–2. — Alkali, – – .

Refill the "acid" burette, but replace the potassium hydroxide by sodium hydroxide. Proceed as above, and tabulate the results thus: —

IV. Sulphuric acid and sodium hydroxide.

15 cc. of acid neutralize      cc. of alkali.
15 cc. of acid neutralize      cc. of alkali.
10 cc. of acid neutralize      cc. of alkali.
Ratio: Acid, 3–3–2. — Alkali, – – .

If the whole experiment has been performed with reasonable accuracy, it will also be found that (4) the ratio of the neutralizing power of the two alkalies is the same for both acids. An actual example will make this point clear. A student found that the ratio of the alkalies in the case of one acid was 11.7 : 19.5, and in the case of the other acid was 12.9 : 21.6. Now, by calculation the second ratio is 12.9 : 21.5, the slight difference being readily explained by the fact that the experimental errors are combined in this one quantity, 21.6.

Tabulate the results so that it will be seen that the ratio of the two alkalies is the same for both acids. Sketch the apparatus.

**Discussion of Experiments 52 and 53.** — These experiments show that when a base and an acid interact a salt is formed. Accurate repetitions of these and similar experiments permit the chemical changes involved in the above experiments on neutralization to be written thus: —

| | | | | | | | |
|---|---|---|---|---|---|---|---|
| $HCl$ | + | $NaOH$ | = | $NaCl$ | + | $H_2O$ | |
| Hydrochloric Acid | | Sodium Hydroxide | | Sodium Chloride | | Water | (11) |
| 36.5 | + | 40 | = | 58.5 | + | 18 | |
| $HCl$ | + | $KOH$ | = | $KCl$ | + | $H_2O$ | |
| Hydrochloric Acid | | Potassium Hydroxide | | Potassium Chloride | | Water | (12) |
| 36.5 | + | 56 | = | 74.5 | + | 18 | |
| $H_2SO_4$ | + | $2\ KOH$ | = | $K_2SO_4$ | + | $2\ H_2O$ | |
| Sulphuric Acid | | Potassium Hydroxide | | Potassium Sulphate | | Water | (13) |
| 98 | + | 112 | = | 174 | + | 36 | |
| $H_2SO_4$ | + | $2\ NaOH$ | = | $Na_2SO_4$ | + | $2\ H_2O$ | |
| Sulphuric Acid | | Sodium Hydroxide | | Sodium Sulphate | | Water | (14) |
| 98 | + | 80 | = | 142 | + | 36 | |

These equations show, besides the usual facts, that in neutralization the hydrogen of the acid and the metal of the base exchange places, forming thereby a salt and water.

**Definitions.** — The experiments in this chapter show that, in addition to certain general properties, acids, bases, and salts are chemically related to a marked degree. No adequate definition can be given of each. The following, however, are generally accepted. An **Acid** is a substance containing hydrogen which is easily replaced by a metal.[1] The replacing metal may be obtained from a compound of the metal, *e.g.* the oxide, carbonate, or hydroxide. Acids are, as a rule, sour, and turn blue litmus red. A **monobasic** acid contains in a molecule only one atom of replaceable hydrogen, *e.g.* nitric acid, $HNO_3$. **Dibasic**, **Tribasic**, etc., acids contain, respectively, two, three, etc., replaceable atoms, *e.g.* sulphuric acid, $H_2SO_4$, and phosphoric acid, $H_3PO_4$. The basicity of acids explains the existence of normal and acid salts. A **base** is a substance containing a metal combined with oxygen and hydrogen, the metal of which easily replaces the hydrogen of acids when the base interacts with an acid. Bases, as a rule, turn red litmus paper blue. A **monacid** base is one in which the metal can be replaced by one atom of hydrogen, *e.g.* sodium hydroxide, $NaOH$. **Diacid** and **Triacid** bases contain a metal replaceable by respectively two and three atoms of hydrogen, *e.g.* $Ca(OH)_2$ and $Al(OH)_3$. A **Salt** is the main product of the interaction of an acid and a base. Salts may be formed in several ways, but they may always be theoretically viewed as if they were formed directly from a fundamental acid or base.

A **Metal** can now be regarded not merely as a hard, lustrous substance, but as a base-forming element. A **Non-metal** likewise may be defined as an acid-forming element. All non-metallic elements, except hydrogen,

---

[1] Water is an exception.

and a few others, form oxides which unite with water and thereby produce an acid; thus:—

$$\begin{array}{ccccc} SO_3 & + & H_2O & = & H_2SO_4 \\ \text{Sulphur} & & \text{Water} & & \text{Sulphuric} \\ \text{Trioxide} & & & & \text{Acid} \end{array}$$

These oxides are often called **Anhydrides.** Similarly, many metallic elements form oxides which unite with water and thereby form hydroxides; thus:—

$$\begin{array}{ccccc} CaO & + & H_2O & = & Ca(OH)_2 \\ \text{Calcium} & & \text{Water} & & \text{Calcium} \\ \text{Oxide} & & & & \text{Hydroxide} \end{array}$$

Such oxides are called **Basic Oxides**. There are some oxides which are neither acidic nor basic. **Hydroxyl** is a name applied to the group OH, which may be regarded as a factor in the molecules of many compounds. It does not exist as elements and compounds exist, but in many reactions this group acts as a unit. It is often called a radical, owing to its fundamental character,—*-yl* being the specific suffix of many radicals.

### CLASS-ROOM EXERCISE. XVIII.

1. Summarize the properties of acids, bases, and salts.
    (*a*) Physical properties.
    (*b*) Chemical properties.
2. Neutralization.
    (*a*) Definition.
    (*b*) Illustration.
    (*c*) Exactly what does Exp. 53 prove?
    (*d*) Application of the principle in quantitative analysis.
    (*e*) Sources of error in Exp. 53, and how avoided.

3. Explain and illustrate the expressions: —
    - (*a*) Acidity of bases.
    - (*b*) Basicity of acids.
    - (*c*) "Every oxygen acid has a corresponding anhydride."
    - (*d*) "Most normal salts are neutral."
    - (*e*) "Sulphuric acid is the starting-point of sulphates."
    - (*f*) "Metals form basic compounds."
    - (*g*) "Hydroxyl is a radical."
4. Historical.
    - (*a*) Historical development of the terms *acid*, *alkali*, and *salt*.
    - (*b*) Lavoisier's contribution to the development of the subject of acids.
    - (*c*) Davy's work on the theory of Lavoisier.
5. Miscellaneous.
    - (1) Explain the terms *volatile alkali*, *caustic alkali*, and *fixed alkali*. What are the alkali metals?
    - (2) Why are there no acid nitrates?
    - (3) Are there any exceptions to the statement, "all substances from which a metal liberates hydrogen are acids"?
    - (4) Literal meaning of acid (adj.), per-, hypo-, alkali.
    - (5) What is the etymological relation of hydrogen to the terms *hydrate*, *hydroxide*, *hydro-* (in *hydrochloric*), *anhydride*?
    - (6) The operation in Exp. 53 is called *titration*. What is meant by "titration to color"?
6. Nomenclature.
    - (*a*) Name the —
        - Sodium salt of hydrochloric acid.
        - Potassium salt of hydrochloric acid.
        - Lead salt of hydrochloric acid.
        - Calcium salt of hydrochloric acid.
        - Barium salt of hydrochloric acid.
        - Zinc salt of hydrochloric acid.
        - Silver salt of hydrochloric acid.
    - (*b*) Name the same salts of nitric acid and of sulphuric acid.
    - (*c*) Name the same salts of hypochlorous acid, of chloric acid, and of nitrous acid
    - (*d*) Name the base corresponding to sodium, potassium, ammonium, calcium, barium, zinc, lead, copper.

(*e*) Name the —

Sodium salt of sulphurous acid.
Potassium salt of manganic acid.
Lead salt of nitric acid.
Potassium salt of perchloric acid.
Calcium salt of hydrofluoric acid.
Sodium salt of nitrous acid.
Calcium salt of hypophosphorous acid.
Sodium salt of carbonic acid.
Potassium salt of tartaric acid.
Lead salt of chromic acid.

(*f*) Significance of the terms *hydroxide*, *hydrate*, *anhydride*, and *hydroxyl*.

(*g*) Select the hydroxyl groups from —

(1) $NaOH$
(2) $KOH$
(3) $Ca(OH)_2$
(4) $Ba(OH)_2$
(5) $Al(OH)_3$
(6) $C_2H_5(OH)$

# CHAPTER IX.

## ATOMS, MOLECULES, AND RELATED SUBJECTS.

**Equivalents.** — The equivalent of zinc was found in Exp. 22, and the result, together with similar results, which may be easily obtained, furnish fundamental data for the subjects discussed in this chapter.

### EQUIVALENT OF MAGNESIUM.

**Experiment 54.** — Determine the equivalent of magnesium by the method of Exp. 22. Use the same apparatus. The reaction proceeds so easily that only about 25 cubic centimeters of warm dilute hydrochloric acid (and no platinum wire) are needed. Use from 0.06 to 0.08 gm. of clean magnesium ribbon, which should be slipped into a short glass tube so that it will sink. Collect the gas in a 100 cc. tube. Proceed in other respects exactly as directed in Exp. 22. Record and calculate as follows: —

I. Data: —

Weight of magnesium taken = gm.

Observed volume of hydrogen = cc.

Observed temperature = °.

Observed pressure = mm.

Tension of aqueous vapor = mm.

II. Calculation: —

Formula for reduction to standard conditions —

$$V = \frac{V'(P' - a)}{760(1 + .00366t)}$$

Corrected volume of hydrogen = cc.
Weight of hydrogen = gm.
Since gm. of magnesium is equivalent to gm. of hydrogen,
then gm. of magnesium are equivalent to 1 gm. of hydrogen.
Equivalent of magnesium =

## EQUIVALENT OF ALUMINIUM.

**Experiment 55.** — Determine the equivalent of aluminium by the method used for zinc and magnesium. Use from 0.06 to 0.07 gm. of aluminium for a 100 cc. tube. Use warm concentrated hydrochloric acid instead of sulphuric acid; no platinum wire is needed. The action may be vigorous and must be watched to prevent the metal from being forced over into the dish. Otherwise, proceed as in the other determinations. Record and calculate as in Exp. 54.

The equivalent or equivalent weight of an element was defined in Chapter III. as that weight of an element chemically equivalent to one part by weight of hydrogen. More specifically it is the number of grams of an element which combines with or replaces one gram of hydrogen. Thus in Exp. 22, 32.5 gm. (approximately) of zinc were found to be equivalent to 1 gm. of hydrogen. In Exps. 54 and 55, 1 gm. of hydrogen was found to be equivalent respectively to 12 gm. of magnesium and 9 gm. of aluminium (approximately), and these numbers, 12 and 9, are the equivalents of these two elements. Some metals do not liberate hydrogen easily from acids, and their equivalents are indirectly found. Sodium, for example, liberates hydrogen from water, and careful experiments show that 23 gm. of sodium are equivalent to 1 gm. of hydrogen. Oxygen is withdrawn from water by passing steam over powdered iron (see Exp. 40), water itself is formed by passing hydrogen over heated copper oxide (see Exp. 44), and both of these experiments, if performed accurately, show that 8 gm. (approximately) of oxygen are equivalent to 1 gm. of hydrogen. Exp. 15 also shows that 8 gm. of oxygen always combine with 12 gm. of magnesium, a number which has been found by an independent experiment (Exp. 54) to be the equivalent of magnesium.

Many metals precipitate another when one is suspended in the solution of the other, and if the metals are weighed the equivalent of one is easily found, thus: —

| | Grams. | |
|---|---|---|
| 1 gm. of zinc precipitates | .97 | of copper. |
| 1 gm. of zinc precipitates | 3.33 | of silver. |
| 1 gm. of magnesium precipitates | 2.64 | of copper. |
| 1 gm. of magnesium precipitates | 9.00 | of silver. |
| 1 gm. of copper precipitates | 3.43 | of silver. |

Now if we recalculate these values and express the results in terms of the equivalents already known, the table reveals other equivalents, thus: —

| Grams. | | Grams. | |
|---|---|---|---|
| 32.5 | of zinc are equivalent to | 31.7 | of copper. |
| 32.5 | of zinc are equivalent to | 108.2 | of silver. |
| 12. | of magnesium are equivalent to | 31.7 | of copper. |
| 12. | of magnesium are equivalent to | 108. | of silver. |
| 31.7 | of copper are equivalent to | 108. | of silver. |

Actual analysis of chlorides reveals the following facts: —

| | |
|---|---|
| 1 gm. of hydrogen combines with | 35.5 gm. of chlorine. |
| 23 gm. of sodium combine with | 35.5 gm. of chlorine. |
| 16 gm. of sulphur combine with | 35.5 gm. of chlorine. |
| 39 gm. of potassium combine with | 35.5 gm. of chlorine. |
| 108 gm. of silver combine with | 35.5 gm. of chlorine. |

Analysis of bromides gives the following results:—

1 gm. of hydrogen combines with 80 gm. of bromine.
23 gm. of sodium combine with 80 gm. of bromine.
39 gm. of potassium combine with 80 gm. of bromine.
108 gm. of silver combine with 80 gm. of bromine.

Analysis of some oxides shows the following results in addition to those given above:—

23 gm. of sodium combine with 8 gm. of oxygen.
39 gm. of potassium combine with 8 gm. of oxygen.
108 gm. of silver combine with 8 gm. of oxygen.

Some sulphides show similar results:—

1 gm. of hydrogen combines with 16 gm. of sulphur.
23 gm. of sodium combine with 16 gm. of sulphur.
39 gm. of potassium combine with 16 gm. of sulphur.
108 gm. of silver combine with 16 gm. of sulphur.

A summary of the above data may be tabulated thus:—

| ELEMENT. | EQUIVALENT. |
|---|---|
| Hydrogen | 1 (by definition) |
| Oxygen | 8 |
| Magnesium | 12 |
| Aluminium | 9 |
| Zinc | 32.5 |
| Copper | 31.7 |
| Silver | 108 |
| Sodium | 23 |
| Potassium | 39 |
| Chlorine | 35.5 |
| Bromine | 80 |
| Sulphur | 16 |

A comparison of some equivalent weights and the approximate atomic weights of the same elements accepted by chemists reveals a remarkable coincidence.

COMPARATIVE TABLE OF EQUIVALENTS AND ATOMIC WEIGHTS.

| ELEMENT. | EQUIVALENT. | APPROXIMATE ATOMIC WEIGHT. |
|---|---|---|
| Hydrogen | 1 | 1 |
| Oxygen | 8 | 16 |
| Chlorine | 35.5 | 35.5 |
| Bromine | 80 | 80 |
| Sulphur | 16 | 32 |
| Zinc | 32.5 | 65 |
| Copper | 31.7 | 63.5 |
| Magnesium | 12 | 24 |
| Sodium | 23 | 23 |
| Potassium | 39 | 39 |
| Silver | 108 | 108 |
| Aluminium | 9 | 27 |

In many cases they are the same, but in some cases the atomic weights are a simple multiple of the equivalent. The first weights called atomic weights were, in many instances, identical with equivalents, and it has taken much labor and thought to select the multiple which is the proper atomic weight of the element in question.

Atomic weights cannot be interpreted without a preliminary consideration of the atomic theory.

**The Atomic Theory.** — Up to the time of Dalton (1766–1844) there were two views regarding the constitution of matter. The first of these is known as the dynamic. This view advocated infinite subdivision; it claimed that there is no limit to the division of matter. A piece of iron, for example, according to this theory, could be divided forever. The other view is known as the atomic. This view advocated finite divisibility. It said that if subdivision goes on, particles will finally be reached which cannot be further subdivided These particles were called atoms by the Greeks, and this latter theory is substantially the atomic theory of the Greeks.

Both these views are old; they date from the time of Lucretius (95–52 B.C.). They are opposed to each other, they are entirely speculative, they are metaphysical, unscientific, and finally have no relation to chemistry as we know it.

Dalton shared the general views of his contemporaries regarding the constitution of matter. He believed in the solid particles of Newton, *i.e.* in the corpuscular theory. He also believed in the existence of individual gases, particularly those which were discovered by Priestley, Cavendish, Black, and others. He believed in the law of definite proportions by weight, and in the law of multiple proportions (see page 201), discovered by Dalton himself, both of which emphasized the general idea of combination by units.

Dalton's meteorological observations led him to believe that aqueous vapor exists separately from the other constituents of the air, also that gases themselves are made up of distinct particles. His work on the solubility of gases in water led him to undertake to determine "the relative weights of ultimate particles of bodies."

Now the above work was largely physical. His views on all matter seem to be the result of shrewd generalizations from his general and vaguely supported ideas of matter, mainly in its physical relations. These ideas were expressed at several times and more particularly in a course of lectures delivered in 1809–1810 at the Royal Institution in London. Regarding atoms, he said in substance in this lecture: "Matter is composed of particles of definite size and weight. These particles are called atoms. All atoms of the same kind of matter are alike. Atoms of different kinds of matter are different. All atoms combine by wholes to form compounds; their relative weight only can be determined." This is the atomic theory of Dalton.

One of the essential properties of matter is weight, and Dalton's first efforts were to determine some of the weights of atoms. To do this he formulated rules and actually found several weights, which he called atomic weights. This work, mingled with crude speculation on the shapes of atoms, seems to have been done at different times and probably preceded to some extent Dalton's final views and work on atoms, though the public expression of his views undoubtedly gave an impetus toward the gravimetric conception of atoms. Dalton's contemporaries did not receive his theory favorably. It is true, no facts, then known, conflicted with the theory, but it is equally true that no set of facts actually supported it. The work of determining atomic weights spread, however, and was given a strong impetus by the Swedish chemist, Berzelius. It has never ceased to attract skilful, persistent workers.

Analysis of chemical compounds reveals the proportions of the components by weight. Thus, one part by weight of hydrogen combines with 35.5 parts by weight of chlorine, and with 80 of bromine. These parts by weight of chlorine and bromine combine in each instance with 39 parts by weight of potassium. The early experimenters soon found that some elements enter into combination with other elements in several proportions. Thus, 8 parts by weight of oxygen combine with one part by weight of hydrogen, but in the larger number of instances 16 parts by weight of oxygen combine with various parts by weight of other elements, *e.g.* zinc, sulphur, and magnesium. Furthermore, sulphur combines with hydrogen in the proportion of 16 parts to 1, but in the vast majority of cases the number of parts by weight of sulphur which combines with other elements is found to be 32. In a word, multiples of combining weights were found, and the problem then reduced itself to selecting the particular multiple to be called the atomic weight. That number which represents the combining proportion of an element in the greatest number of cases, its favorite number, so to speak, was selected as the atomic weight of that element. The number, then, which we call the atomic weight represents not an absolute weight, but a relative weight, *i.e.* how many times heavier the atom in question is than the atom of hydrogen. The absolute weight of an atom of hydrogen is not known. Its weight is arbitrarily chosen as one.

The atomic theory is viewed at the present time as a satisfactory explanation of the laws of combination of matter revealed by experiment. It is generally adopted as a working hypothesis, not as a fact.

An atom according to the present views is that particle of matter which has not yet been subdivided into smaller particles. It combines with itself or with other atoms and thereby forms molecules.

**Atomic Weights.** — The atomic weight of an element, therefore, is that number which represents how many times heavier the atom of an element is than the atom of hydrogen (see "Various Figures for Atomic Weights," page 142).

Hence, the numbers hitherto called the equivalents of the elements are also, in many cases, the atomic weights of the respective elements. Thus,

23 is the atomic weight of sodium.
108 is the atomic weight of silver.
80 is the atomic weight of bromine.
127 is the atomic weight of iodine.
35.5 is the atomic weight of chlorine.

The determination of exact atomic weights is a difficult operation, and no single method furnishes sufficient data for a final selection of the numbers. If atoms combined in only one proportion, atom for atom, analysis would reveal the proportions, and if one were the standard, or if its weight in terms of the standard were known, the relative weight of the other could be easily calculated. For example, Stas, a Belgian chemist who made masterly determinations of atomic weights, found that 121.4993 gm. of silver chloride were formed by burning 91.462 gm. of silver in chlorine. It was known to him that an atom of silver unites with one atom of chlorine to form silver chloride, and that the atomic weight of chlorine is 35.453. Hence, he calculated the atomic weight of silver thus:—

$$\begin{array}{r} 121.4993 \\ \underline{91.4620} \\ 30.0373 \end{array} \text{ weight of chlorine used.}$$

$$91.462 : 30.0373 :: x : 35.453$$

$$x = 107.95, \text{ atomic weight of silver.}$$

Since many atoms combine in several proportions, the number selected as the atomic weight is based on considerations such as vapor density, agreement with the periodic law, specific heat, etc. Only one of these approximate or check methods, however, will be considered.

**Law of Specific Heats.** — Dulong and Petit in 1819 announced the Law of Specific Heats, viz.: —

*The product of the atomic weight and the specific heat of the solid elements is a constant quantity.*

This law is not perfectly accurate, since in many cases this product is only approximately 6.4 — the number selected by Dulong and Petit as the constant — as appears from the

TABLE OF SPECIFIC HEATS.[1]

| ELEMENT. | SPECIFIC HEAT. | ATOMIC WEIGHT. | PRODUCT. |
|---|---|---|---|
| Aluminium | 0.214 | 27.0 | 5.8 |
| Calcium | 0.170 | 39.91 | 6.8 |
| Copper | 0.095 | 63.18 | 6.0 |
| Iron | 0.114 | 55.88 | 6.4 |
| Lead | 0.031 | 206.4 | 6.4 |
| Magnesium | 0.250 | 24.3 | 6.1 |
| Potassium | 0.166 | 39.03 | 6.5 |
| Silver | 0.057 | 107.66 | 6.1 |
| Sodium | 0.293 | 23.00 | 6.7 |
| Sulphur | 0.178 | 31.98 | 5.7 |
| Tin | 0.055 | 118.8 | 6.5 |
| Zinc | 0.094 | 65.1 | 6.1 |

An example will make this law more intelligible. The specific heat of silver is .057; if 6.4 is divided by this num-

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[1] From Meyer's "Outlines of Theoretical Chemistry," p. 22.

ber, the quotient is 112.2. This is not the accepted atomic weight of silver, but it checks the other determinations, since 112.2 points to 108 and not to any of its multiples. On the other hand the specific heat of mercury is .032; if 6.4 is divided by this number, the quotient, 200, is the atomic weight of mercury, and this number is confirmed by other methods.

**Various Figures for Atomic Weights.** — Obviously the figures for the atomic weight of the same element will vary with the standard. Hence different atomic weight tables will be found, some on one standard, some on another. Hydrogen was, until recently, the generally adopted standard. Since many atomic weights are now directly determined with reference to oxygen, the hydrogen standard is being replaced by the oxygen standard, *i.e.* 16 and not 1 is the standard atomic weight. But as the atomic weight of hydrogen becomes only 1.008, if oxygen is taken as exactly 16, the variation in the usual table of atomic weights is very slight, especially when whole numbers are used. Appendix C contains a table of atomic weights on both standards. The approximate atomic weight (see App. C, Table I.) of the most common elements should be learned, not slavishly, but by constant association.

**Molecular Weights.** — Since atoms combine to form molecules, a molecular weight is the sum of the weights of the atoms in a molecule. If the molecule is the simple combination of two atoms, then the molecular weight is the sum of the respective atomic weights. Thus experimental evidence points conclusively to the fact that sodium chloride contains one atom of sodium and one atom of chlorine, hence its formula is $NaCl$ and its molecular weight is 58.5 (23 + 35.5). Since molecules may theoretically contain any number of atoms, the molecular weight and formula are rigidly connected. It is customary to assume the simplest formula until evidence reveals an error. Thus HO was formerly regarded as the formula of water; but when the density of steam showed that the molecular weight must be 18, this fact, together with other data of a similar value, led to a change from HO to the present formula, $H_2O$.

**Avogadro's Law.** — Various generalizations have contributed to the solution of the problem of molecular weight determination. The oldest of these is Avogadro's Law, which is —

*There is an equal number of molecules in equal volumes of all gases at the same temperature and pressure.*

This law was announced by Avogadro, an Italian physicist, in 1811, but was not recognized by chemists and physicists as a fundamental law of science until about 1868. It has been repeatedly verified experimentally and deduced mathematically (from the kinetic theory of gases). This law means that the weights of equal volumes of gases are in the same proportion as the molecular weights of the gases themselves. Thus the weight of a liter of oxygen was found by Exp. 17 to be approximately 1.43 gm. A liter of hydrogen under the same conditions weighs .0896 gm. But .0896 and 1.43 are in the same ratio as 2 to 32, the accepted *molecular* weights of hydrogen and oxygen, since it is believed that these gases, like most elementary gases, have two atoms in a molecule. The density of all compound gases has been calculated, and the relations of some appear in the following table: —

| Gas. | Weight of a Liter in Grams under Standard Conditions. | Weight of a Molecule compared with the Weight of One Molecule of Hydrogen ($H_2 = 2$). | Weight of a Molecule compared with the Weight of One Atom of Hydrogen ($H = 1$). |
|---|---|---|---|
| Carbon dioxide | 1.97 | 22 | 44 |
| Hydrochloric acid | 1.61 | 18.25 | 36.5 |
| Chlorine | 3.18 | 35.5 | 71 |
| Ammonia | 0.77 | 8.5 | 17.0 |
| Nitrogen | 1.256 | 14 | 28 |
| Steam | .806 | 9.0 | 18.0 |

The numbers in the last column are the molecular weights of the respective gases. They are equal to the sum of the weights of the combination of atoms which represents a molecule of the different gases. The molecular weight of a gas is, therefore, twice its specific gravity referred to hydrogen.

**Calculation of Atomic and Molecular Weights from Experimental Data.** — Many of the experiments in this book furnish data for the calculation of approximate atomic and molecular weights.

The atomic weight of magnesium may be calculated by utilizing the results of Exps. 15 and 54. A student found that .37 gm. of magnesium combined with .25 gm. of oxygen in forming magnesium oxide. Accepting 16 as the atomic weight of oxygen, we may write the proportion —

$$\text{Weight of oxygen} : \text{Weight of magnesium} :: \text{Atomic weight of oxygen} : \text{Atomic weight of magnesium}$$

$$.25 : .37 :: 16 : x$$

$$x = 23.7$$

The same student verified this result by finding the equivalent of magnesium to be 11.95. The atomic weight of magnesium is twice the equivalent, hence the atomic weight becomes 23.9 (*i.e.* $2 \times 11.95$), a result agreeing fairly well with 23.7 previously found. The approximate atomic weight of magnesium is 24.

Again, from the results of the first experiment the molecular weight of magnesium oxide may be found. Thus .25 gm. of oxygen formed .62 gm. (*i.e.* $.25 + .37$) of magnesium oxide. Hence the proportion:—

| Weight of oxygen | : | Weight of magnesium oxide | :: | Atomic weight of oxygen | : | Molecular weight of magnesium oxide |
|---|---|---|---|---|---|---|
| .25 | : | .62 | :: | 16 | : | $x$ |
| | | | | | | $x = 39.7$ |

The molecular weight of magnesium oxide is 40 (*i.e.* $24 + 16$), which is in fair agreement with the above experimental determination. It is generally accepted that a molecule of magnesium oxide contains one atom each of oxygen and magnesium.

The atomic weight of zinc is similarly found. A student obtained these results:—

1.05 gm. of zinc united with .26 gm. of oxygen to form 1.31 gm. of zinc oxide. Using the same form of proportion as in the case of magnesium, we may write

$$.26 : 1.05 :: 16 : x$$

$$x = 64.61$$

An average of the determinations of a large class working in the same laboratory on the equivalent of zinc was 32.45. Since the atomic weight of zinc is twice the equivalent, the second result gives 64.9 as the atomic weight of zinc. The approximate atomic weight of zinc is 65.

A student in performing Exp. 16 obtained these results:

| | Grams. |
|---|---|
| Weight of potassium chlorate used | 1.500 |
| Weight of oxygen lost . . . . . | .585 |
| Weight of potassium chloride left | .915 |

Now chemists believe that a molecule of potassium chlorate contains three atoms of oxygen, hence we may write the proportion: —

| Weight of three atoms of oxygen | : | Molecular weight of potassium chlorate | : : | Weight of oxygen | : | Weight of potassium chlorate |
|---|---|---|---|---|---|---|
| or 48 | : | $x$ | : : | .585 | : | 1.5 |

$x = 123.07$

The molecular weight of potassium chlorate, found by adding the approximate atomic weights, is 122.5. Again, the molecular weight of potassium chloride may be found by the proportion: —

| Weight of three atoms of oxygen | : | Molecular weight of potassium chloride | : : | Weight of oxygen | : | Weight of potassium chloride |
|---|---|---|---|---|---|---|
| or 48 | : | $x$ | : : | .585 | : | .915 |

$x = 75.07$

The molecular weight of potassium chloride, found by adding its atomic weights, is 74.5, it being assumed that the molecule contains one atom each of potassium and chlorine.

The molecular weight of water may be calculated from the result of Exp. 44. A student found that .905 gm. of oxygen formed 1.021 gm. of water. Hence: —

| Weight of oxygen | : | Weight of water | : : | Atomic weight of oxygen | : | Molecular weight of water |
|---|---|---|---|---|---|---|
| or .905 | : | 1.021 | : : | 16 | : | $x$ |

$x = 18.05$

This result agrees closely with 18, the molecular weight of water. These results are only approximate, but they serve to illustrate some of the methods employed in the determination of atomic and molecular weights.

**Valence.**—Investigation shows that each element has power to hold in combination a definite number of atoms of other elements, and that the power varies with the elements which interact. Thus chlorine, oxygen, nitrogen, and carbon, under certain conditions, combine respectively with 1, 2, 3, and 4 atoms of hydrogen. This capacity is called **valence**. It is measured by the number of hydrogen atoms with which one atom of the element combines to form a molecule. Hydrogen and all elements combining with one atom of hydrogen, or its equivalent, are called univalent elements, *e.g.* sodium and potassium; those combining with two atoms are called bivalent; with three, trivalent; with four, tetravalent; with five, pentavalent, or quinquivalent. Elements of the same valence replace each other atom for atom. If the valence is different, such a number of atoms or groups of atoms must be taken that the valence of the interchanging parts will be equal. Thus $Cu(NO_3)_2$ is the formula of copper nitrate, because in this compound copper has the valence two and is equivalent to two $NO_3$ groups which are univalent, since they combine directly with one atom of hydrogen. Similarly, $Na_2O$ is the formula of sodium oxide, but CaO is the formula of calcium oxide, the valence of oxygen and calcium being two, while the valence of sodium is one. The valence of an element is the same in the same compound, but often varies in different compounds, according to circumstances, such as temperature and character of the other element. Valence must not be assumed, but learned by experience, since the whole theory is not a sharply defined conception.

**Percentage Composition.** — The formula of a compound is not only an abbreviation of its name, but also an expression of its composition. Thus $H_2SO_4$ means that in every molecule of sulphuric acid 2 parts are hydrogen, 32 are sulphur, and 64 are oxygen, if the parts are taken by weight. As there are 98 parts $(2 + 32 + 64)$ in the molecule, then —

$$\frac{2}{98} \text{ are hydrogen.}$$

$$\frac{32}{98} \text{ are sulphur.}$$

$$\frac{64}{98} \text{ are oxygen.}$$

Should the composition in per cent be desired, it can be calculated thus: —

$$2 : 98 :: x : 100$$
$$x = 2.04\,\% \text{ of hydrogen.}$$
$$32 : 98 :: x : 100$$
$$x = 32.65\,\% \text{ of sulphur.}$$
$$64 : 98 :: x : 100$$
$$x = 65.31\,\% \text{ of oxygen.}$$
$$\text{Total, } 100.00\,\%$$

By this method it is possible to calculate the percentage composition of any compound, if its formula is known.

**Formula.** — If the percentage of each element in a compound is divided by the atomic weight of the element, the quotients will be in the same ratio as the number of atoms of each element in the compound. Thus the percentage composition of sulphuric acid is—

$$H = 2.04.$$
$$S = 32.65.$$
$$O = 65.31.$$

Now if each percentage is divided by the propei atomic weight, we may write (approximately):—

$$\frac{2.04}{1} = 2 = \text{number of hydrogen atoms.}$$

$$\frac{32.65}{32} = 1 = \text{number of sulphur atoms.}$$

$$\frac{65.31}{16} = 4 = \text{number of oxygen atoms.}$$

Hence the simplest formula of sulphuric acid is $H_2SO_4$. The simplest formula calculated from the percentage composition is not always the molecular formula, that is, it does not always express the composition and number of atoms which are in a molecule of the compound when in a gaseous state. Since every formula is designed to be a molecular formula, it is usually necessary to determine it from independent experimental data. The molecular weight of a compound is twice its vapor density, hence if the density is known or can be determined, then the molecular formula can easily be calculated from the simplest formula. For example, the simplest formula of water based on gravimetric analysis is HO, but since its vapor density is 9, its molecular weight is 18, a number which requires the formula $H_2O$. Again, the formula $H_2O$ is supported by numerous facts independently discovered, such as the atomic weight of oxygen, the divisibility of hydrogen into two equal parts, and the ratio of the volumes of the components. Another illustration will make the subject still clearer. The simplest formula of a compound of carbon and hydrogen was found to be $CH_2$. Its vapor density was found to be 81.4. Therefore its molecular weight must be 162.8, or nearly twelve times that corresponding to $CH_2$, hence the molecular formula is $C_{12}H_{24}$. Molecular formulas of other compounds may be similarly found.

Until recently the vapor density was the only reliable check on molecular weight determinations, but, as will presently be seen, newer and more comprehensive methods are now applicable.

**Elevation of Boiling Point and Depression of Freezing Point.**—It has long been known that a solution of common salt boils at a higher temperature and freezes at a lower temperature than pure water. (See Exp. 31.) The elevation of the boiling point and depression of the freezing point are controlled by laws, which are easily determined by methods applicable to many compounds, and molecular weights are now studied by these methods, especially by the depression of the freezing point. Investigation of molecular weights determined by the latter method shows that the depression of the freezing point of a solution is proportional to the weight of the substance dissolved in a given volume, and that the same number of molecules of any dissolved substance produces the same fall in temperature of the freezing point of the solution; *i.e.* for solutions of different substances but of equal concentration, the depression of the freezing point is inversely proportional to the molecular weights. Thus if $X$ is the depression found experimentally to be produced by a one per cent solution of sugar, and $Y$ the depression produced by a one per cent solution of urea, the following proportion may be written :—

$$Y : X :: \text{mol. wt. of sugar} : \text{mol. wt. of urea.}$$

The molecular weight of sugar is known to be 342, and the molecular weight of urea calculated from the above proportion is 60 — a number agreeing closely with the molecular weight found by other methods. This method is used to determine whether a molecular weight is a given number or its multiple, a question often arising in the case of different organic compounds having the same percentage composition. The method is applicable, however, only in those cases in which (1) there is no chemical action between the solvent and the solute, and (2) the solute does not decompose.

**Ions and Ionization.** — The molecular weights of many substances which had been determined by the older methods agree with the values obtained by the newer methods. There is a marked disagreement, however, in the case of solutions of strong acids, strong bases, and salts, or in general, in the case of *electrolytes*, *i.e.* all substances the aqueous solutions of which conduct electricity. In some instances the molecular weight is only half of the molecular weight found by other methods, *i.e.* the solution acts as if twice as many molecules were in solution. Thus the molecular weight of sodium chloride was found by two experiments to be 29.4 and 30.2, instead of 58.5. An extensive experimental study of solutions, supplemented by keen conclusions, derived from facts of the most varied kind, has led to the general adoption of the theory — known as the Theory of Electrolytic Dissociation — that —

*When acids, bases, and salts dissolve in water their molecules partly split up into ions.*

Ions are atoms, or groups of atoms, charged with electricity; they are carriers of electricity, which they give up when they reach the electrodes of the battery, or source of electric current. There are always two kinds in every solution, — negative ions or anions, and positive ions or cations. Hydrogen, metals, and metallic radicals (like ammonium) form positive ions; salt radicals, like $SO_4$, and the hydroxyl group OH, form negative ions. Potassium chloride in solution is largely split up into the ions, $\overset{+}{K}$ and $\overset{-}{Cl}$; potassium nitrate becomes largely $\overset{+}{K}$ and $\overset{-}{NO_3}$.

This dissociation, or ionization, increases as the dilution increases, and in the case of strong acids, bases, and their salts the dissociation is nearly complete even in moderately concentrated solutions. The percentage of dissociation of some familiar substances (normal solutions at 18°) is shown in the following —

TABLE OF IONIZATION.

| SUBSTANCE. | PER CENT OF IONIZATION. |
|---|---|
| Hydrochloric acid . . . . . . . . . . . . . | 78 |
| Nitric acid . . . . . . . . . . . . . . . | 82 |
| Sulphuric acid . . . . . . . . . . . . . . | 51 |
| Potassium chloride . . . . . . . . . . . . | 75 |
| Potassium nitrate . . . . . . . . . . . . . | 64 |
| Potassium sulphate . . . . . . . . . . . . | 53 |
| Potassium hydroxide . . . . . . . . . . . . | 77 |
| Sodium hydroxide . . . . . . . . . . . . | 73 |

**Applications of the Theory of Electrolytic Dissociation.** — Many reactions, already studied, are more intelligible when interpreted in the light of this theory. Thus, when solutions of silver nitrate and sodium chloride are mixed, insoluble silver chloride is formed. The two solutions before mixing contain the ions of sodium, chlorine, silver, and the group $NO_3$; when mixed, the ions of silver and chlorine unite to form silver chloride, a compound *insoluble* in water, and hence not decomposed or ionized; the solution still contains ions of sodium and of $NO_3$, together with the ions of the salt added in excess. If solutions of silver nitrate and potassium chlorate are mixed, no silver chloride is formed, because no chlorine ions are available. Potassium chlorate dissociates into the ions $\overset{+}{K}$ and $\overset{-}{ClO_3}$. Ordinary tests are tests for ions. This theory has been widely applied by the German chemist, Ostwald, to the explanation of many fundamental facts of analytical chemistry.

### CLASS-ROOM EXERCISE. XIX.

1. Explain and illustrate the following: —
   - (*a*) Atom, molecule, mass.
   - (*b*) Equivalent, atomic weight, molecular weight.
   - (*c*) Specific heat.
   - (*d*) Ion, anion, cation, ionization, electrolytic dissociation, depression of freezing point, elevation of boiling point.
2. Historical.
   - (*a*) Essential facts in the life of Dalton.
   - (*b*) The contributions of Berzelius to the subject of atomic weights.
   - (*c*) The work of Stas on atomic weights and its value at the present time.
   - (*d*) Dulong and Petit and their joint labors.
   - (*e*) Avogadro, Ampère, and Cannizzaro, and their contributions to Avogadro's Law.

3. The modern theory of solution.
   - (*a*) Historical development.
   - (*b*) Brief but exact statement of the theory.
   - (*c*) The hydrate theory.
   - (*d*) Application to simple reactions.

4. Valence.
   - (*a*) Give the valence of—

Sodium in sodium chloride.
Calcium in calcium chloride.
Oxygen in water.
Oxygen in calcium oxide.
Nitrogen in ammonia gas ($NH_3$).
Nitrogen in nitric acid ($HNO_3$).
Sulphur in sulphur dioxide ($SO_2$).
Sulphur in sulphur trioxide ($SO_3$).
Sulphur in sulphuric acid.
Carbon in carbon dioxide.
Carbon in carbon tetrachloride ($CCl_4$).
Zinc in zinc sulphate.
Magnesium in magnesium sulphate.

## PROBLEMS. XIII.

1. If 0.34 gm. of sodium yields 165 cc of hydrogen (corrected volume), what is the equivalent of sodium ?

2. If 0.48 gm. of magnesium yields 443 cc. of hydrogen (corrected volume), what is the equivalent of magnesium ?

3. A piece of zinc weighing 1.95 gm. evolved 670 cc. of hydrogen from sulphuric acid. Calculate the equivalent of zinc.

4. If 0.48 gm. of zinc replaces 1.6 gm. of silver, calculate the equivalent of silver.

5. A gram of aluminium yields 1271 cc. of hydrogen at 10° C. by interaction with hydrochloric acid. What is the equivalent and the atomic weight of aluminium ?

6. When 0.5 gm. of aluminium was warmed with potassium hydroxide, 660 cc. of hydrogen at 13° C. were liberated. What is the atomic weight of aluminium ?

7. If 0.3697 gm. of aluminium liberates 0.04106 gm. of hydrogen from sodium hydroxide, what is the atomic weight of aluminium ?

8. A chemist found that 8.2144 gm. of alum, $(NH_4)_2SO_4Al_2(SO_4)_3 . 24 H_2O$, dried at 21°–25 C., yielded 0.9258 gm. of aluminium oxide, $Al_2O_3$. What is the atomic weight of aluminium ? (Assume O = 15.961, S = 31.966, N = 14.01.)

9. The same chemist found that 6.9617 gm. of aluminium bromide, $AlBr_3$, required 8.4429 gm. of silver to precipitate all the bromine. What is the atomic weight of aluminium ? (Assume Ag = 107.649, and Br = 79.754.)

10. If 0.2 gm. of magnesium liberates 195.5 cc. of hydrogen at 13° C., what is the atomic weight of magnesium ?

11. A compound of silver and chlorine weighing 121.1993 gm. was produced by heating 91.462 gm. of silver in a current of chlorine gas. What is the atomic weight of silver ? (Assume Cl = 35.5, and that one atom of silver unites with one atom of chlorine.)

12. A molecule of calcium oxide contains one atom of calcium and one atom of oxygen. The percentage composition is 71.43 per cent of calcium and 28.57 per cent of oxygen (O = 16). What is the atomic weight of calcium ?

13. Calculate the percentage composition of—

(*a*) Water.
(*b*) Magnetic oxide of iron, $Fe_3O_4$.
(*c*) Crystallized sodium carbonate, $Na_2CO_3 . 10 H_2O$.
(*d*) $Ca_3(PO_4)_2 . Ca_2(PO_4)F$.
(*e*) Alum, $Al_2K_2(SO_4)_4 . 24 H_2O$.
(*f*) Zinc blende, ZnS.
(*g*) Calamine, zinc carbonate, $ZnCO_3$.
(*h*) Feldspar, $K_2Al_2 Si_6O_{16}$.
(*i*) $Cu_3(PO_4)_2 . 3 H_2O$.
(*j*) $Ca_3(PO_4)_2 . CaCO_3$.

14. Calculate the percentage composition of—

(*a*) Potassium chlorate.
(*b*) Sugar, $C_{12}H_{22}O_{11}$.
(*c*) Calcium sulphate, $CaSO_4$.
(*d*) Zinc sulphate.
(*e*) Magnesium oxide.
(*f*) Copper oxide, CuO.

15. Calculate the simplest formula of the compounds which have the indicated percentage composition, and give the name of the compound: —

(*a*) Hydrogen = 11.11
Oxygen = 88.89

(*b*) Sodium = 32.39
Sulphur = 22.54
Oxygen = 45.07

(*c*) Carbon = 27.27
Oxygen = 72.72

16. Calculate the simplest formula of the compounds which have the indicated percentage composition: —

(*a*) Nitrogen = 82.353
Hydrogen = 17.647

(*b*) Oxygen = 30
Iron = 70

(*c*) Hydrogen = 1.0
Carbon = 11.99
Potassium = 39.06
Oxygen = 47.95

17. Calculate the molecular weight of the following compounds by finding the sum of the atomic weights: —

(*a*) Copper sulphate.
(*b*) Barium chloride.
(*c*) Manganese dioxide.
(*d*) Calcium oxide.
(*e*) Sodium hydroxide.
(*f*) Potassium hydroxide.
(*g*) Sodium carbonate.
(*h*) Potassium nitrate.

18. A liter of sulphurous oxide gas ($SO_2$) weighs 2.8672 gm. What is its molecular weight?

19. If 1500 cc. of carbon monoxide gas (CO) weigh 1.8816 gm., what is its molecular weight?

20. A volume of gas measures 5.40675 l. at 15° C. and 740 mm. and weighs 7.616 gm. What is its molecular weight?

21. Calculate the molecular formula of the compounds having the indicated percentage composition and vapor density: —

(*a*) Carbon = 73.8
Hydrogen = 8.7
Nitrogen = 17.1
Vapor density = 80.2

(*b*) Carbon = 92.3
Hydrogen = 7.7
Vapor density = 38.8

(*c*) Carbon = 39.9
Hydrogen = 6.7
Oxygen = 53.4.
Vapor density = 30.5

# CHAPTER X.

## CHLORINE AND HYDROCHLORIC ACID.

**Chlorine.** — It has been shown in previous experiments that hydrochloric acid contains hydrogen, but as hydrochloric acid is not an element, it must contain something else besides hydrogen. Chemists have shown that hydrochloric acid is a compound of two elements, and confirmatory evidence of this fact will be obtained later. Obviously, then, to obtain this other constituent of hydrochloric acid, we must treat hydrochloric acid with something which will liberate this constituent and at the same time combine with the hydrogen. Several experiments have shown that oxygen combines readily with hydrogen, hence if hydrochloric acid is heated with some substance which yields oxygen easily, this other element in the acid will be liberated, provided it does not combine easily with oxygen. It has also been shown that some oxides when heated yield oxygen, and, as a matter of fact, manganese dioxide was used by early chemists to prepare oxygen. Furthermore, Scheele, the discoverer of this other constituent of hydrochloric acid, actually prepared it by heating hydrochloric acid with ores (mainly oxides) of manganese. He did not recognize the new product as an element and it was not until about 1810 that Davy proved its elementary nature and gave it the name **Chlorine**, by which it is now known. Chlorine should not be inhaled, as it irritates the sensitive lining of the throat and nostrils.

### THE PREPARATION OF CHLORINE FROM HYDROCHLORIC ACID AND MANGANESE DIOXIDE.

*Perform this Experiment in the Hood.*

**Experiment 56.** — Put a dozen small pieces of manganese dioxide in a test tube, cover with concentrated hydrochloric acid, and heat gently. Describe the action and enumerate the characteristic properties of the most apparent product.

### PREPARATION AND PROPERTIES OF CHLORINE.

*Perform this Experiment in the Hood.*

**Experiment 57.** — *Apparatus:* Construct an apparatus like that shown in Fig. 52. The flask *A* holds 500 cc. and stands on a sand bath; the connecting tube *a* is about 7 cm. long and the ends of *E* and *B* are about 5 cm. apart; *C* is a wooden block with a hole in the center large enough to admit *B*, which reaches to the bottom of the bottle *D*. To remove *D*, hold *C* in one hand and with the other pull *D* aside, bending at the same time the whole delivery tube at the point *a*.

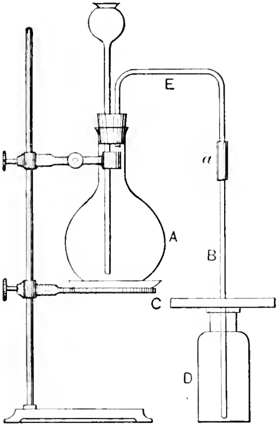

FIG. 52. — Apparatus for the preparation of chlorine gas.

*Supplies:* When the apparatus is arranged, provide 30 gm. of granulated manganese dioxide, six bottles, a little powdered antimony, a little powdered iron, bundle of fine brass wire, a taper, a strip of calico, paper with writing in black and in red ink, litmus paper, short glass tube drawn out to a fine point, deflagrating spoon, turpentine.

*Caution — Turpentine is inflammable.*

*Process:* Put the manganese dioxide in the flask, insert the stopper, and slowly pour through the safety tube enough concentrated hydrochloric acid to cover the manganese dioxide. Heat gently and finally stand the burner under the sand bath so that the flame just touches the iron pan. Chlorine is rapidly evolved and passes into *D*, which should be removed when full (as seen by the color) and covered with a glass plate. Collect six bottles and proceed at once with the following experiments: —

(*a*) Sprinkle a pinch or two of powdered antimony into a bottle of chlorine. Heat some powdered iron in a deflagrating spoon and pour it into another bottle. Into a third bottle thrust a loose bundle of very fine brass wire which has been slightly heated. Describe the action, giving especially in each case the evidence of chemical action and of new products.

(*b*) Into a bottle of dry chlorine put a piece of calico, litmus paper (both colors), and paper containing writing in black and in red ink. Allow the whole to remain undisturbed for a few minutes and then describe the change, if any. Add several drops of water, and describe the change. Draw a general conclusion from the whole experiment.

(*c*) Hold a burning taper in a bottle of chlorine long enough to observe the result. Draw a conclusion and verify it thus: Attach the pointed glass tube to a rubber tube connected with the gas cock, light the gas and lower a *small* jet into a bottle of chlorine. Does the result verify the above conclusion? The same result may be shown as follows: Fold a strip of filter paper (about 10 cm. wide) into the shape shown in Fig. 53; cautiously heat[1] about 10 cc. of turpentine in a large test tube; saturate the paper with the hot turpentine and drop it into a bottle of chlorine. Describe the result. When the action is over, examine the paper, and draw a conclusion regarding the action between hot turpentine and chlorine.

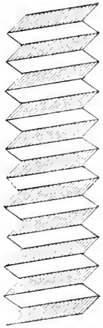

FIG. 53.—Fluted paper for Exp. 57 (*c*).

Wax (in the taper), illuminating gas, and turpentine are mainly compounds of hydrogen and carbon. Explain the result in (*c*).

### BLEACHING BY CHLORINE.

**Experiment 58.** — Put a little bleaching powder into a test tube and add enough water to make a thin paste. Add a few drops of dilute

---

[1] Hold the test tube with the holder. Remember that turpentine ignites easily. If the turpentine catches fire, press a damp towel over it.

sulphuric acid to the paste, and then dip a strip of bright colored calico into the mixture. Remove the calico and wash it with water. Describe the change.

If time permits, repeat with a piece of unbleached cotton cloth and describe the change.

**A liter of chlorine** under standard conditions weighs 3.18 grams.

CLASS-ROOM EXERCISE. XX.

1. Bleaching and bleaching powder.
    (*a*) What is bleaching powder? How is it made? For what kind of bleaching is it commonly used?
    (*b*) The phenomenon in Exp. 57 (*b*) is due to what general chemical change? Will chlorine bleach a lead pencil mark or newspaper ink? Why? Predict the result when red hot charcoal is thrust into a bottle of chlorine.
    (*c*) Why was the turpentine heated in Exp. 57 (*c*)?
    (*d*) What is the per cent of "available chlorine" in bleaching powder?
2. Additional study of chlorine.
    (*a*) Summarize the physical properties of chlorine. What is its most striking property? Is it heavier or lighter than air? How do you know? The method in Exp. 57 is called "collecting a gas by *downward* displacement." Why so called?
    (*b*) What class of chemical compounds is formed in Exp. 57 (*a*)? Name others in this class already used or studied. Could chlorine be collected over mercury? Why?
    (*c*) What does Exp. 57, as a whole, show regarding the chemical activity of chlorine? In what respects does its activity resemble and differ from the activity of oxygen? Is chlorine ever found uncombined? Why? Does it form many compounds? Are they in general stable or unstable?
    (*d*) For what chemical element has chlorine a strong attraction? What experiment shows this?
    (*e*) How may chlorine be quickly distinguished from all other gases previously studied?
    (*f*) The industrial preparation of chlorine. (*g*) Uses.

3. The equation for the interaction of hydrochloric acid and manganese dioxide is —

$$\begin{array}{ccccccccc} 4\,HCl & + & MnO_2 & = & Cl_2 & + & MnCl_2 & + & 2\,H_2O \\ 146 & + & 87 & = & 71 & + & 126 & + & 36. \end{array}$$

Name all the products. How much chlorine can be made from 247 gm. of manganese dioxide?

4. A bottle of chlorine water was exposed to the sunlight until all the chlorine disappeared.

- (*a*) What two products were formed?
- (*b*) Write the equation for the reaction.
- (*c*) What weight of chlorine is necessary to form 20 gm. of the gaseous product?
- (*d*) What volume of chlorine is necessary to form 20 gm. of the other product?

5. If dry steam and chlorine are passed through a red-hot porcelain tube, two gases are formed.

- (*a*) Name the gases.
- (*b*) Write the equation for the reaction.
- (*c*) How may the gases be separated?

6. Historical.

- (*a*) The discovery of the elementary nature of chlorine.
- (*b*) Essential facts in the life of Davy?
- (*c*) A brief account of Faraday's contributions to chemistry.

7. Liquid chlorine.

- (*a*) Faraday's work.
- (*b*) Present method of liquefaction.
- (*c*) Transportation.
- (*d*) Uses.

8. Nascent state.

- (*a*) Definition.
- (*b*) Illustration.
- (*c*) Theoretical value of the conception.
- (*d*) Chlorine is sometimes prepared by heating a mixture of sodium chloride, sulphuric acid, and manganese dioxide. Interpret the reaction from the standpoint of the nascent state. Write the equation for the reaction.

9. Miscellaneous.
   - (*a*) Atomic weight of manganese.
   - (*b*) Molecular weight of manganese dioxide, and of manganese chloride.
   - (*c*) Symbol of antimony, carbon, and manganese.

**Hydrochloric Acid.** — It was shown in Exp. 19 that hydrochloric acid contains hydrogen, and in Exp. 56 that it contains chlorine. These results should be verified by a synthesis of the acid. It will first be necessary, however, to know the characteristic properties of hydrochloric acid, so that the product of the synthesis may be readily recognized. It should be remembered that hydrochloric acid, as we have used it, is a water solution of the gas, and that the same term, hydrochloric acid, includes both the gas and its solution.

### CHARACTERISTIC PROPERTIES OF HYDROCHLORIC ACID GAS.

**Experiment 59.** — Heat a little concentrated hydrochloric acid in a test tube, and observe the white fumes and pungent odor. The gas is so soluble in water that it draws the moisture from the air, forming thereby dense white fumes. Try the action of the gas on moistened litmus paper and on a burning match. Moisten a glass rod with ammonium hydroxide, hold it near the mouth of the test tube, and note the striking result. This is a confirmatory test for hydrochloric acid gas, and should always be used to verify its presence. Summarize the characteristic properties of hydrochloric acid gas.

### SYNTHESIS OF HYDROCHLORIC ACID GAS.

*Remember that a mixture of hydrogen and air explodes if brought near a flame.*

**Experiment 60.** — Construct a hydrogen generator (see Exp. 19) and attach about 15 cm. of rubber tubing to the delivery tube. Before generating the hydrogen, collect a bottle of chlorine from a generator constructed like Fig. 54. If heat is necessary, clamp *A* to a stand and lengthen *B* (see Fig. 82). Generate a rather rapid current of hydrogen and let it run about three minutes; attach the platinum tip (see App. A, § 13) to the rubber tube, and let the gas run two minutes longer; then light the hydrogen and lower the flame well into the bottle of chlorine. Let the action continue long enough to observe the color and size of the flame, and the color and quantity of the product, then extinguish the flame by pinching the rubber tube. Cover the bottle with a glass plate and, after the action in the hydrogen generator has been stopped or the generator removed to a safe place, test the contents of the bottle as the gas was tested in Exp. 59. Do the results agree in every particular? If not, repeat until the disagreement is removed or explained.

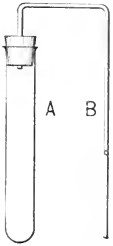

Fig. 54. — Simple generator for chlorine gas.

Summarize the evidence of the qualitative composition of hydrochloric acid gas.

It was shown in Exp. 57 (*b*) that chlorine unites easily with metals, forming chlorides, just as oxygen forms oxides. There are many compounds of chlorine and metals, the most common being sodium chloride, which would have been formed in Exp. 57 (*b*) if sodium, instead of iron and antimony, had been burned in chlorine. It is expensive and inconvenient to prepare large quantities of hydrochloric acid by a synthesis, so the chlorides are allowed to interact with a compound which yields hydrogen easily. Sulphuric acid and sodium chloride are usually used.

## PREPARATION AND PROPERTIES OF HYDROCHLORIC ACID.

**Experiment 61.** — Construct and arrange an apparatus like Fig. 52 to collect the gas. The apparatus shown in Fig. 55 is used to prepare a solution of the gas, and is the same as shown in Fig. 52, except that *B* passes into the absorption apparatus, *FG*. Instead of *F*, the apparatus *H*, shown in Fig. 56, may be used. The flask *A* stands on a sand bath, and the gas is collected in *D* by downward displacement, exactly as in the collection of chlorine.

I. To 15 cc. of cold water add slowly and with constant stirring 40 gm. of concentrated sulphuric acid. While the mixture is cooling put 40 gm. of sodium chloride in the flask *A*, arrange the apparatus to collect the gas, and then pour the *cold* acid through the safety tube upon the salt. Heat gently with a low flame. Collect four bottles, as in the preparation of chlorine, cover with glass plates, and set aside until needed. Disconnect at *a*, and replace *D* by *FG* (or *HG*). The flask *F* (or *H*) contains about 50 cc. of water, and the tube reaches just to the surface of the water, but in *G* the tube dips into the water. Pass the gas into the absorption apparatus, until the mixture in *A* seems exhausted, or for a reasonable time.

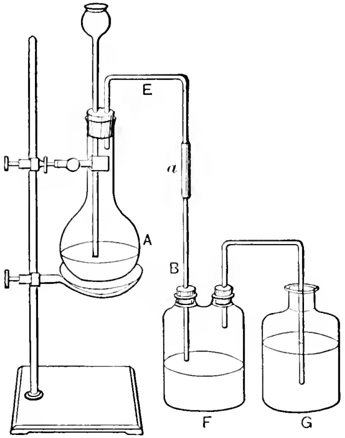

FIG. 55. — Apparatus for preparing hydrochloric acid.

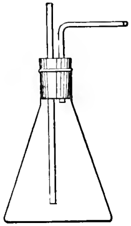

FIG. 56. — Apparatus *H*, which may replace *F* in Fig. 55.

II. Meanwhile proceed to study the hydrochloric acid gas already collected.

(*a*) Invert a bottle, and hold its mouth beneath the water in a pneumatic trough. Describe the result. What property of the gas does this result illustrate? How would a rise of temperature affect the result? Compare this property with that of oxygen, hydrogen, nitrogen, and air under similar conditions. Apply the litmus test to the liquid in the bottle, or to the liquid beneath the mouth of the bottle.

(*b*) Insert a burning stick into another bottle of the gas. Describe the result. How do oxygen, hydrogen, nitrogen, chlorine, carbon dioxide, and air act under similar conditions?

(*c*) Hold a piece of wet filter paper near the mouth of another bottle of the gas. Describe the result. What is the cause ?

(*d*) Drop in the remaining bottle of hydrochloric acid gas a piece of filter paper saturated with ammonium hydroxide. The dense white fumes are ammonium chloride, and are formed by direct combination of ammonia gas and hydrochloric acid gas. The equation is —

$$\begin{array}{ccccc} HCl & + & NH_3 & = & NH_4Cl \\ & & \text{Ammonia gas} & & \text{Ammonium} \\ & & & & \text{Chloride} \end{array}$$

Determine the solubility of the ammonium chloride in water.

(*e*) Enumerate other properties of hydrochloric acid gas which you have observed.

III. Disconnect *FG* and continue to heat the flask (in the hood) until no more gas is evolved. Save the contents of the flask for Exp. 62. Meanwhile study the general properties of the solution of hydrochloric acid gas in *F* according to Exp. 48. Record the results.

IV. Every acid has properties peculiar to itself. To a solution of silver nitrate add a few drops of nitric acid and then a little of the hydrochloric acid made in *F*. A white, curdy precipitate, called *silver chloride*, is formed. Expose a little of this precipitate to the sunlight. Determine its solubility by adding separate small portions to cold water, hot water, and warm ammonium hydroxide. Tabulate the properties of silver chloride. This whole experiment (IV.), solubility included, constitutes the *test* for hydrochloric acid. The same test answers for chlorides. Try it. (Compare Exp. 29 (*b*).)

It has been shown that when zinc and sulphuric acid interact (1) a new compound is formed, and (2) a definite weight relation exists between the zinc and hydrogen. It is possible, though not so convenient, to prove that similar relations exist between the factors and products when sodium chloride and sulphuric acid interact. Hydrochloric acid gas is not so easily collected as hydrogen ; furthermore, the chemical action between sodium chloride and sulphuric acid is not always complete. However, extensive work by chemists has shown that only one substance besides hydrochloric acid is formed, if the chemical action is complete. It is only necessary, therefore, in a qualitative test, to confine our examination to this substance in a study of

### THE INTERACTION OF SODIUM CHLORIDE AND SULPHURIC ACID.

**Experiment 62.** — I. Add enough warm water to the flask *A* to loosen the solid, and pour a portion of the contents into a casserole or evaporating dish. Evaporate to complete dryness slowly over a piece of wire gauze and then heat strongly (in the hood) until all fumes of sulphuric and hydrochloric acids have been driven off.

II. (*a*) Dissolve a little of the residue from I. in distilled water. Test part of the solution for a chloride and the remainder for a sulphate. Draw a general conclusion from the result.

(*b*) Test another portion thus: Moisten the platinum test wire (see App. A, § 14) with water, dip the loop into the solid, and hold it in the Bunsen flame. Note the intense color. Clean the wire with hot water or sandpaper, and repeat with any compound known to contain sodium, or which you have made from sodium. (Compare Exp. 42 (*c*).) What does this result prove qualitatively about the specific nature of the substance under examination? Give the name of the substance. State the chemical changes which led to its formation.

### COMPOSITION OF HYDROCHLORIC ACID GAS.

*This experiment, including the preparation of the sodium amalgam, should be done in the hood.*

**Experiment 63.** — *Method:* The composition of hydrochloric acid gas may be found by decomposing a known volume of the gas with sodium amalgam. Hydrogen is the only volatile product, and is measured directly in the tube used in the experiment. The mercury in the amalgam takes no part in the reaction.

*Supplies:* The apparatus shown in Fig. 57; a tightly fitting cork for the tube *E*, tall jar, crucible (Hessian), 100 gm. of mercury, 5 gm.

of sodium, stiff iron wire, iron pan, glass-stoppered bottle. In Fig. 57, *A* is a 500 cc. flask, *B* is a small dropping funnel, *C* is a U-tube to be filled with fragments of glass drenched with concentrated sulphuric acid, *D* is a delivery tube reaching to the bottom of the tube *E*. The tube *E* is graduated and its total capacity should be known.

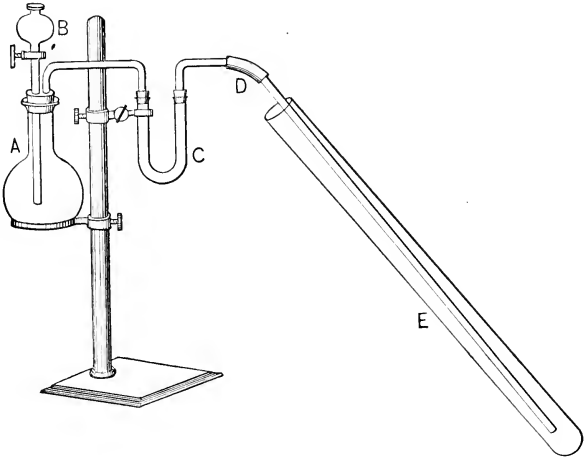

FIG. 57.—Apparatus for determining the composition of hydrochloric acid gas.

I. *Preparation of sodium amalgam:* Put 100 gm. of mercury in a rather large Hessian or iron crucible, slip the crucible into a ring, clamped to an iron stand, so that about two-thirds of the crucible is below the ring. The crucible should have an iron, porcelain, or sand cover. Cut about 5 gm. of sodium into pieces the size of a bean and slide all the sodium at once from a piece of paper into the mercury. Cover the crucible immediately and stand a lighted burner under it. Hold the rubber tubing between the thumb and forefinger and stand aside from the apparatus. As soon as the heat is sufficient, the reaction takes place suddenly, often accompanied by a flash of light. The instant a flash or noise occurs, pinch the tube to extinguish the burner flame, and stand back until the fumes of mercury disappear. Extinguish the gas, remove the cover, and stir the amalgam vigorously with a stiff iron wire. Before the amalgam cools, pour it into a shallow iron pan.

Allow it to solidify, then break it quickly into small pieces and put it into a glass-stoppered bottle. Preserve carefully, as sodium amalgam is decomposed by moisture.

II. *Process:* (*a*) Arrange the apparatus so that the hydrochloric acid gas will pass from the generator through the drying apparatus to the *bottom* of the graduated tube, as shown in Fig. 57. Generate the gas by allowing concentrated sulphuric acid to drop slowly from the funnel into concentrated hydrochloric acid.

(*b*) Fill the tube *E* completely with dry gas free from air. Allow the generator to run moderately fast, and occasionally test the escaping gas at the open end of the tube with moist blue litmus paper. As the operation proceeds, gradually draw the graduated tube away from the delivery tube, keeping it as nearly vertical as possible. Do not touch the tube with the hands, but use a test tube holder. Have ready about 10 gm. of powdered sodium amalgam, wrapped in a small piece of thin paper, and when the tube is full of gas, withdraw it from the delivery tube, thrust the sodium amalgam into the tube and instantly insert the stopper. Invert the tube repeatedly to allow all the gas to come in contact with the sodium amalgam. As the action proceeds, a white deposit coats the inside of the tube and the sodium amalgam slowly liquefies. When the action seems complete, uncork the tube under water. If the preceding part has been correctly performed, the water will rush into the tube. Clamp the tube so that the water is at the same level within and without, but not so that the tube is over the stream of hydrogen arising from the excess of amalgam at the bottom of the jar. After it has stood five minutes, adjust the levels (if necessary) without touching the tube with the hand, and read the volume. What is the relation between the original and the final volume?

(*c*) Unclamp the tube, place the thumb over the open end, lift it from the water, invert, remove the thumb, and instantly hold a lighted match at the mouth. What gas was in the tube?

(*d*) Summarize the evidence of the composition of hydrochloric acid gas. Sketch the apparatus.

**Discussion of Experiment 63.** — The result of Exp. 63 shows approximately that a simple relation exists between the volumes of hydrogen and chlorine derived from hydrochloric acid gas. It has been proved by many experimenters that these volumes are equal. This result has been confirmed by decomposing hydrochloric acid by an electric current, just as water was decomposed; furthermore it has been shown that when equal volumes of hydrogen and chlorine are allowed to combine, hydrochloric acid is formed without a residue of either component. When the gases so combine, the equation for the reaction is:—

$$H_2 + Cl_2 = 2\,HCl.$$

This simple relation, together with the equally simple relation revealed by the volumetric composition of steam, will be subsequently explained.

**The Formula of Hydrochloric Acid Gas is HCl.**— No derivative of hydrochloric acid is known which contains less hydrogen or less chlorine in a molecule. Hence HCl is the simplest expression of the facts. Again, the molecular weight of hydrochloric acid gas is 36.5, since its density is 18.25.

### CLASS-ROOM EXERCISE. XXI.

1. Theoretical.

The equation for the complete interaction of sodium chloride and sulphuric acid is—

$$\begin{array}{ccccccc} 2\,NaCl & + & H_2SO_4 & = & Na_2SO_4 & + & 2\,HCl \\ & & & & \text{Sodium} & & \\ & & & & \text{Sulphate} & & \\ 117 & + & 98 & = & 142 & + & 73 \end{array} \qquad (15)$$

(*a*) Why is this equation correct ?
(*b*) Why does it contain 2 NaCl instead of NaCl ?
(*c*) How would you prove this equation ?

2. Further study of hydrochloric acid.

(*a*) How made industrially ? (*b*) Its usual impurities.
(*c*) Cause of its yellow color.
(*d*) Commercial name, and why so called?

(*e*) Explain the statement, "The coefficient of absorption of hydrochloric acid gas at 0° C. is 503."
(*f*) What happens when a solution of hydrochloric acid is boiled?
(*g*) The liquefaction of hydrochloric acid gas.

3. Historical.
(*a*) The history of hydrochloric acid up to Priestley's time.
(*b*) Priestley's work. (*c*) Davy's work.
(*d*) The various names of chlorine and hydrochloric acid, and their significance.

4. Miscellaneous.
(*a*) Name and describe any compounds of chlorine not already considered.

## PROBLEMS. XIV.

1. How many grams of chlorine can be obtained from the complete action of hydrochloric acid on one kilogram of manganese dioxide? How many liters of chlorine?

2. How much chlorine (by weight) will 3.6 l. of water absorb, provided it absorbs twice its volume?

3. A manufacturer needs 10 l. of commercial hydrochloric acid, sp. gr. 1.174. How much sodium chloride and sulphuric acid are necessary?

4. What volume of chlorine would be obtained at 13° C. and 740 mm. by heating 10 gm. of manganese dioxide with an excess of hydrochloric acid?

5. How much manganese dioxide will yield a liter of chlorine?

6. A liter of commercial hydrochloric acid, sp. gr. 1.16, is required. How much sodium chloride, sulphuric acid, and water (in the receiver) must be used?

7. How many liters of hydrogen gas can be obtained from 1 l. of hydrochloric acid gas?

8. Assume (1) 100 parts of potassium chlorate yield 60.85 parts of potassium chloride, (2) 100 parts of potassium chloride yield 192.4 parts of silver chloride, (3) the formulas $KClO_3$, $KCl$, and $AgCl$, (4) $O = 16$, and (5) 100 parts of silver combine with 32.87 parts of chlorine. From these assumptions, find the atomic weights of chlorine, potassium, and silver.

9. A chemist found that 91.462 gm. of silver, when heated in a current of chlorine, yielded 121.4993 gm. of silver chloride. What is the atomic weight of chlorine? (Assume Ag = 108.09.)

10. If a volume of dry chlorine measures 22.32 l. at 0° C. and 760 mm., what will it measure at 17° C. and 730 mm.?

11. The simplest form of the equation for the interaction of chlorine and water is —

$$H_2O + Cl_2 = 2\ HCl + O.$$

- (*a*) How many liters of hydrochloric acid gas can be made from 875 gm. of water?
- (*b*) How many liters of oxygen will be produced?
- (*c*) How many liters of chlorine will be required?

12. If 20 l. of hydrogen are mixed with 10 l. of chlorine, (*a*) how many liters of hydrochloric acid gas are produced, and (*b*) which gas and how much remains in excess?

13. Hydrogen and chlorine were mixed and exploded, and the product weighed 73 gm.

- (*a*) What weight of each gas was used?
- (*b*) What volume of each gas was used?

14. What volume of hydrogen is necessary to unite with one gram of chlorine?

15. Calculate the percentage composition of—

- (*a*) Hydrochloric acid gas.
- (*b*) Sodium chloride.
- (*c*) Sodium sulphate.
- (*d*) Silver chloride.
- (*e*) Potassium chloride.
- (*f*) Potassium chlorate.
- (*g*) Acid sodium sulphate, $HNaSO_4$.

16. How much sodium chloride is necessary to make a kilogram of hydrochloric acid gas?

17. How many pounds of sodium chloride are necessary to make 150 lb. of hydrochloric acid solution containing 31.5 per cent of available acid?

18. How many grams of manganese dioxide are necessary to make 100 gm. of chlorine from hydrochloric acid?

19. How many grams of manganese dioxide are necessary to make 40 l. of chlorine at 37° C.?

# CHAPTER XI.

## COMPOUNDS OF NITROGEN.

**Ammonia.** — The word *ammonia* is loosely used to include both the gas and its solution, just as hydrochloric acid means either the gas or its solution. The solution of ammonia gas is, more strictly, ammonium hydroxide, as will appear from subsequent experiments.

Ammonia is best prepared from its compounds.

### PREPARATION OF AMMONIA.

**Experiment 64.** — (*a*) Mix and grind together in a mortar a little ammonium chloride and lime (calcium oxide). What product is revealed by the odor?

(*b*) Dissolve a little ammonium chloride in water, add a little sodium hydroxide solution, and warm gently. What is the most apparent product?

### PREPARATION AND PROPERTIES OF AMMONIA GAS AND AMMONIUM HYDROXIDE.

**Experiment 65.** — *Method:* Prepare the gas and the solution substantially as in the case of hydrochloric acid.

*Apparatus:* Prepare the gas in a modified form of the apparatus shown in Fig. 52. The gas is collected by *upward* displacement in the apparatus shown in Fig. 58. It is an inverted form of the "*B-C-D* part" of Fig. 52; *B* is connected with *E* by *a*, as before. The solution — ammonium hydroxide — is prepared in the apparatus shown in Fig. 55. This experiment must be performed without interruption, and all parts of the apparatus must be constructed and ready for immediate use before the experiment proper begins.

*Supplies:* 30 gm. of ammonium chloride, 30 gm. of quicklime which has been exposed to the air long enough to make it crumble easily, 5 bottles, 4 glass plates, litmus paper, stick or splint of wood, pneumatic trough or dish filled (as usual) with cold water, filter paper.

*Process:* I. Mix thoroughly the ammonium chloride and lime. Slip the mixture into the flask *A*, insert the stopper and its tubes, and add a little water through the safety tube to moisten the mass. Connect the collection apparatus, and see that all joints are tight. Heat *A* gently with a low flame. Collect four bottles of gas (in dry bottles), removing each when full, and covering with a glass plate. It can easily be told when the bottle is full by holding a piece of moist red litmus paper near the mouth. *Do not smell at the mouth of the bottle.*

Disconnect at *a*, attach the absorption apparatus *FG*, and allow the ammonia gas to be absorbed by the water in both *F* and *G*, while the collected gas is being studied, *remembering to shake F occasionally.*

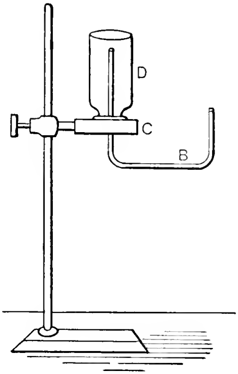

FIG. 58.—Apparatus for collecting ammonia gas.

II. The general properties of ammonia gas are found by the following experiments: —

(*a*) Try the action of a bottle of the gas on moist litmus paper; try the action on a burning stick or taper (having a small flame); notice, also, the color, and *very cautiously* the odor by wafting the gas gently toward the nose. Is it heavier or lighter than air, chlorine, and hydrochloric acid gas? Has it acid or alkaline properties? Will it burn or support combustion?

(*b*) Hold a bottle with its mouth just below the surface of the water in a pneumatic trough or other vessel. Watch the result. Describe it. What property of the gas does it reveal? Is it a marked property? Test the liquid in (or just below) the bottle with litmus paper (both colors). What has become of the gas?

(*c*) Pour a few drops of concentrated hydrochloric acid into an empty, warm, dry bottle. Cover it with a glass plate, invert it, and stand it upon a covered bottle of ammonia gas. Remove both plates at once, and hold the bottles together by grasping them firmly about their necks.

Describe the action. Give all the evidence of chemical action. If heat was evolved, state where and why. Allow the white product to settle, then scrape out as much as possible upon a paper, transfer it to a test tube, and warm gently with a little sodium hydroxide solution. What is formed? What was the white product? (If the last question cannot be answered from the evidence here presented, consult Exp. 61 (*d*) and Exp. 64 (*b*).)

III. Disconnect the absorption apparatus, and remove the generator to a hood. Study the properties of the ammonium hydroxide in *F* as follows: —

(*a*) Determine the general properties by repeating Exp. 49 (*a*) and (*b*).

(*b*) Boil a little of the liquid gently in a test tube. What gas is evolved?

(*c*) Try the effect of ammonium hydroxide on a grease spot. What practical application is made of this property?

The equation for the interaction of ammonium chloride and lime — calcium oxide — is —

$$\begin{array}{ccccccccc} 2\,NH_4Cl & + & CaO & = & 2\,NH_3 & + & CaCl_2 + H_2O & & \\ \text{Ammonium Chloride} & + & \text{Calcium Oxide} & = & \text{Ammonia Gas} & + & \text{Calcium Chloride} & & (16) \\ 107 & + & 56 & = & 34 & + & 111 \quad + \quad 18 & & \end{array}$$

**Composition of Ammonia Gas.** — Ammonia gas is a compound of two elements. A qualitative and a quantitative examination of the gas will reveal its components and their relations.

### QUALITATIVE EXAMINATION OF THE COMPOSITION OF AMMONIA GAS.

**Experiment 66.** — *Method:* Pass dry ammonia gas over heated magnesium powder and test the two products.

*Apparatus:* The apparatus is shown in Fig. 59. A is a 500 cc. flask, provided with a safety gauge tube and a delivery tube connected directly with a U-tube *B*. The left-hand limb of *B* is to be empty, the bend loosely plugged with cotton, and the other limb filled with lumps of quicklime. *CC′* is a glass tube about 25 cm. long and 2 cm. in diameter; it is loosely filled with soda lime, except at the ends, where the soda lime is replaced by loose plugs of cotton. Before the final adjustment, this tube should be tapped along its entire length to insure an unobstructed channel for the ammonia gas. These tubes, *B* and *CC′*, serve to dry the ammonia gas — a precaution absolutely essential to the success of the experiment. If the ammonia gas is evolved slowly, these tubes are sufficient; but if haste is necessary, then a calcium chloride drying jar, filled with lime, should be placed between *B* and *CC′*. The tube *CC′* is attached to *B* and *DD′*, as shown in the figure. *DD′* is a combustion tube from 12 to 15 cm. long. Two copper wires at *D* and *D′*, arranged as usual, serve to prevent accidental fracture of the tube or melting of the stoppers. A delivery tube, partly shown as *E*, passes from the combustion tube into a glass dish, arranged, as usual, to collect a gas over water. The apparatus is supported by iron stands provided with clamps which may be attached at convenient points, such as *C′* and *D′*. Various modifications may be made to economize space, but such changes should not shorten the drying tubes. The flask *A* stands on a sand bath or asbestos board, and the ammonia gas is evolved by heating concentrated ammonium hydroxide over a low flame, which is controlled by a Hofmann screw attached to the gas tube (lengthened, if necessary) near the operator. The combustion tube is heated with a wing-top burner. Two or three empty bottles will also be needed.

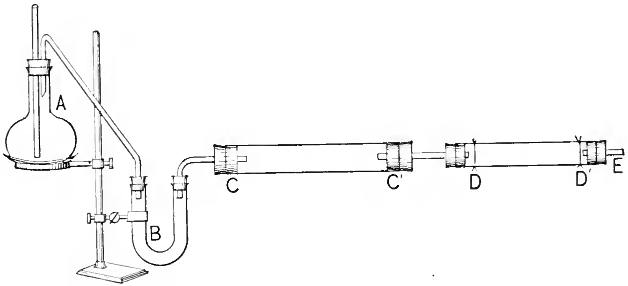

FIG. 59. — Apparatus for determining the qualitative composition of ammonia gas.

*Process:* Fill *A* half full of concentrated ammonium hydroxide and connect securely with the tubes *B* and *CC'*. Put enough magnesium powder in *DD'* to make a thin layer from 5 to 7 cm. long. (See App. A, § 11.) Connect *DD'* with the drying and delivery tubes. Fill the glass dish with enough dilute sulphuric acid to submerge the end of the (lengthened) delivery tube *E*. Fill the bottles (preferably of 125 cc. capacity) with cold water, and stand them conveniently near the dish, so that they may be instantly inverted, if necessary. Heat the flask *A* gently, and begin to heat the combustion tube. Search for leaks, which may be detected by the odor, as soon as any ammonia gas is known to have been evolved. Watch three points — (1) the safety gauge tube, which will reveal any excessive back pressure; (2) the magnesium, which will change in color as the experiment proceeds; and (3) the end of the delivery tube, where the gas which is to be collected rises, and where the escape of the ammonia gas indicates, approximately, the rate of evolution of the latter gas. When it is judged that the apparatus is free from air, heat the combustion tube throughout its entire length, and, finally, heat the magnesium strongly. A slow current of ammonia gas should be kept passing all the time. As the temperature rises, the magnesium changes to a greenish yellow color, and a gas is evolved which does not dissolve in water. When this color appears, slip a bottle over the end of the delivery tube and collect the gas. Continue to heat the magnesium until the evolution of this new gas slackens. Collect three bottles of the gas evolved and set them aside temporarily. Cool the combustion tube slowly, and, when cool enough to handle, disconnect the combustion tube from the delivery tube and the drying apparatus, and quickly pour the contents of the combustion tube into a small dry bottle provided with a glass stopper. Preserve for subsequent examination.

Test each bottle of gas with a lighted match. What is the gas? What is its source? What, then, is one component of ammonia gas?

Test the greenish powder formed in the combustion tube thus: Pour a little into a test tube half full of water. What gas is evolved? Repeat, if the observation is not specific. Record this observation, and reserve it for subsequent explanation.

This experiment reveals one component of ammonia gas, and the product of the interaction of the greenish yellow powder and water suggests the other.

The other component of ammonia gas may be found as "the residual gas" by

## THE QUANTITATIVE EXAMINATION OF THE COMPOSITION OF AMMONIA GAS.

**Experiment 67.** — *Method:* Decompose ammonia gas in a measured volume of chlorine. The method depends mainly upon the fact, already shown, that chlorine will withdraw hydrogen from compounds of that element.

*Apparatus:* In addition to the special apparatus shown in Fig. 60, there is needed a chlorine generator, arranged as shown in Fig. 52, except that the "*B-C-D* part" is replaced by a tube shaped like *E* in Fig. 30, and long enough to reach to the bottom of a pneumatic trough; a tall jar filled with water at the temperature of the room; a little (about 25–30 cc.) concentrated ammonium hydroxide, and dilute sulphuric acid. The apparatus shown in Fig. 60 consists of the graduated tube *A*, of known capacity; it is provided with a gas-tight, one-hole rubber stopper, carrying the small dropping funnel *B*, which is never removed from the stopper. The latter may be made by sealing a piece of tubing (5 cm. long and 2 cm. in diameter) to a stop-cock tube near the stop-cock; the other end is then cut off at any desired point. It is desirable, though not necessary, that the tube *A* be graduated. The dropping funnel shown in Fig. 32 may be used instead of *B*, but it is not so satisfactory.

FIG. 60. — Apparatus for determining the quantitative composition of ammonia gas.

*Process:* Generate the chlorine in the hood as directed in Exp. 57. Meanwhile ascertain the *total* capacity of the tube *A* (see App. A, § 22). Then fill it with water, at the temperature of the room, invert, and clamp it in the pneumatic trough so that the open end is over the hole in the shelf. By this time the chlorine generator should be free from air; this condition, however, can be told only by experience, but since it is absolutely essential to an accurate result, the generator must, in case of doubt, be allowed to run at least fifteen minutes after the time the chlorine bubbles from the end of the delivery tube. When the chlorine is believed to be free from air, slip the end of the delivery tube under the shelf, and pass a rapid stream of gas up into the tube; the tube will soon be filled, since the gas will not dissolve in the water to any appreciable extent. When the tube is full, let it stand a few minutes until the water runs down; be sure the tube is completely full before removing it from the trough.

Meanwhile, fill the dropping funnel half full of concentrated ammonium hydroxide, open the stop-cock, and allow the liquid to fill the lower tube. Close the stop-cock and stand the funnel in a beaker or bottle in another hood, or in some convenient place not far from the chlorine apparatus. Unclamp the tube, cover the open end with the thumb or second finger, invert, and quickly insert the stopper with its funnel. Push in the stopper gas-tight. Fill the funnel with concentrated ammonium hydroxide, cautiously open the stop-cock, and allow a drop or two of the ammonium hydroxide to enter the tube. Considerable heat, and sometimes a faint flash of light, together with dense, white fumes, indicate chemical action. Add ammonium hydroxide, drop by drop, until the reaction seems complete. Take care not to let any gases out of or any air into the tube when the stop-cock is opened. Now add dilute sulphuric acid through the funnel until at least one-third of the tube is full. Fill the funnel with water, cover with a small piece of filter paper, and invert the whole apparatus in a tall jar of water. Remove the paper, open the stop-cock, allow the water to run in, and let the whole apparatus remain undisturbed for at least fifteen minutes. Then, without touching the upper part of the tube, raise or lower it until the water is at the same height within and without; clamp the tube in this position and read the volume of "the residual gas." What proportion of the original volume is it?

Close the stop-cock, lift the tube from the water, remove the stopper, lower a lighted match attached to a wire into the tube, and observe the result. Test the gas also with lime water, as previously directed. What is the gas? Why? Why not any of the other gases studied?

Record the two definite results of this experiment.

Sketch the apparatus used for the quantitative examination of the composition of ammonia gas.

**Discussion of Experiments 66 and 67.** — When ammonium hydroxide, which has been shown to be identical with a solution of ammonia gas, is added to chlorine, the chlorine withdraws the hydrogen from the ammonia gas, forming hydrochloric acid gas and "the residual gas." But the hydrochloric acid gas instantly combines with the ammonia gas to form ammonium chloride, which is seen at first as the dense white fumes, but which finally dissolves in the water added. These changes continue until all the chlorine is removed. The excess of ammonia present, both as gas and in solution, is neutralized with sulphuric acid, and the solid product — ammonium sulphate — also dissolves in the water. Hence at the end of the experiment only "the residual gas" is present *as a gas*, and the volume and nature may be found as directed. It was shown in Exp. 63 that chlorine combines with its own volume of hydrogen. Therefore there must have been three times as much hydrogen as nitrogen in the ammonia gas, since there was three times as much chlorine at the beginning as there was nitrogen at the end of the experiment.

Since it has been shown that ammonia is a compound of hydrogen and nitrogen, the yellowish green compound formed in Exp. 66 must be a compound of magnesium and nitrogen. It is, in fact, magnesium nitride, $Mg_3N_2$. When added to water the interaction yields ammonia gas as one product, thus confirming the evidence of the composition of ammonia gas furnished by Exp. 67.

**Equations for the Chemical Action in Experiments 66 and 67.** — The following equations represent in the simplest form the chemical action in these experiments. In Exp. 66 —

$$2\,NH_3 \quad + \quad 3\,Mg \quad = \quad \underset{\substack{\text{Magnesium} \\ \text{Nitride}}}{Mg_3N_2} \quad + \quad 3\,H_2$$

$$Mg_3N_2 \quad + \quad 3\,H_2O \quad = \quad 2\,NH_3 \quad + \quad 3\,MgO$$

In Exp. 67 —

$$\mathrm{NH_3} + 3\,\mathrm{Cl} = 3\,\mathrm{HCl} + \mathrm{N}$$

$$\mathrm{HCl} + \mathrm{NH_3} = \mathrm{NH_4Cl}$$

$$\mathrm{H_2SO_4} + 2\,\mathrm{NH_3} = \underset{\text{Ammonium Sulphate}}{\mathrm{(NH_4)_2SO_4}}$$

**The Volumetric Composition of Ammonia Gas.** — Additional evidence of the composition of ammonia gas is furnished by a determination of

### THE VOLUMETRIC COMPOSITION OF AMMONIA GAS.

**Experiment 68.** — *Method:* Pass electric sparks through a measured volume of dry ammonia gas and measure the final volume.

*Apparatus:* Eudiometer, mercury, mercury trough, ammonia gas generator and drying apparatus, induction coil, battery, meter stick, barometer, and thermometer.

*Process:* Fill the eudiometer with clean, dry mercury, invert it in a trough of mercury and clamp it in a perpendicular position. Generate ammonia gas and dry it as in Exp. 66. Pass the current through the drying apparatus long enough to expel all the air, and then introduce about 20 cc. of the dry ammonia gas into the eudiometer. Arrange a thermometer so that the bulb dips into the mercury, and allow the whole apparatus to stand undisturbed for at least fifteen minutes. Then, without touching that part of the tube which contains the gas, read and record the following: —

Volume of the gas $= V' =$ cc.
Height of mercury in the tube $= H =$ mm.
Thermometer $= t =$ °.
Barometer $= P' =$ mm.

Connect the eudiometer with the induction coil and pass sparks through the gas for about half an hour. Meanwhile reduce the observed gas volume to standard conditions by the formula —

$$V = \frac{V'(P' - H)}{760\,(1 + (.00366 \times t))}.$$

The factor $H$ must be subtracted from the barometer reading because the gas is supporting only that column of mercury which is the difference between the barometer column and the column extending from the surface of the mercury in the trough to the upper surface of the mercury in the tube — found by actual measurement with a meter stick.

Read the volume again roughly, noting also the barometer and thermometer. Continue to pass sparks through the gas. Read again soon, and if there is any marked change in volume, continue to pass the sparks. The sparking must continue until the volume is constant. When there is no further change in volume, disconnect the wires, allow the whole apparatus to stand undisturbed for ten or fifteen minutes, then read accurately as before. Reduce the observed volume to standard conditions. What is the relation between the original and final volumes?

**Discussion of Experiment 68.** — The result of this experiment simply shows that when ammonia gas is subjected to the action of electric sparks the volume is doubled. Further consideration is necessary to determine the nature of the chemical change, which may be best explained by an illustration. A volume of ammonia gas measured 100 cc., and after sparking, the residual volume was 200 cc. Then 90 cc. of dry oxygen were added, the mixture was exploded, and 65 cc. of gas remained. It is known, of course, from previous experiments, that ammonia gas yields by decomposition the two gases, hydrogen and nitrogen. Hence the total volume of hydrogen and oxygen which disappeared as a result of the explosion must have been 225 cc. (*i.e.* 290–65). But two-thirds of this volume must have been hydrogen, and the remainder oxygen, since these gases unite in the proportion of two to one by volume to form water. Therefore, 150 cc. ($\frac{2}{3}$ of 225 cc.) of hydrogen must have come from the decomposed ammonia gas; and since only 200 cc. of mixed gases were formed by the sparking, the remaining 50 cc. must be the nitrogen. These two gases — hydrogen and nitrogen —

therefore, are combined in ammonia gas in the proportion of three parts to one by volume.

The facts revealed by Exp. 68 are most simply expressed by the equation: —

$$\begin{array}{ccccc} N_2 & + & 3\,H_2 & = & 2\,NH_3 \\ \text{One volume of nitrogen} & + & \text{Three volumes of hydrogen} & = & \text{Two volumes of ammonia gas} \end{array} \qquad (17)$$

**A liter of ammonia gas** under standard conditions weighs .77 gm.

**Formula of Ammonia Gas.** — The gravimetric composition of ammonia gas reveals the fact that fourteen parts of nitrogen combine with three parts of hydrogen. The vapor density of ammonia gas is 8.5 (approximately on the hydrogen standard), and hence its molecular weight is 17. The simplest formula which meets all these requirements is $NH_3$.

**Law of Gay-Lussac.** — A comparison of the volumetric composition of water vapor (steam), hydrochloric acid gas, and ammonia gas (see pages 101 and 169) reveals the fact that in each case the volume relations can be expressed by whole numbers. Thus —

**Two** volumes of hydrogen unite with **one** volume of oxygen } to form **two** volumes of water vapor.

**One** volume of hydrogen unites with **one** volume of chlorine } to form **two** volumes of hydrochloric acid gas.

**Three** volumes of hydrogen unite with **one** volume of nitrogen } to form **two** volumes of ammonia gas.

The simple relation between the volumes of the components, and between the components and their compounds in the three cases examined, is just as simple in all cases.

The law was first pointed out by Gay-Lussac, who stated the relation substantially as follows: —

*Gases combine in volumes which bear a simple relation to each other and to the product.*

The law may be expanded thus — "When two or more elementary gases unite to form a gaseous compound, the volume of a fixed quantity of the compound stands to the volumes of the combining elements in a ratio which can be expressed by whole numbers."

**Ammonium Compounds.** — It has been shown in several experiments (Exp. 64 (*b*) and Exp. 65 (*c*)) that ammonium chloride is formed by direct combination of ammonia gas and hydrochloric acid gas. This compound may be formed in the same way as any other salt.

### INTERACTION OF AMMONIUM HYDROXIDE AND HYDROCHLORIC ACID.

**Experiment 69.** — Put about 50 cc. of ammonium hydroxide in an evaporating dish or casserole, and add dilute hydrochloric acid, drop by drop, until the solution is just neutral or faintly alkaline. A convenient dropping apparatus is shown in Fig. 61. Evaporate to dryness on a water bath or very slowly over a piece of wire gauze. Test the residue as follows: —

(*a*) Is it an acid, base, or salt?

(*b*) Warm a little with sodium hydroxide solution. What is formed? Draw a conclusion as to the nature of the residue.

(*c*) Support the dish on a piece of wire gauze and warm gently until a decided change occurs. Describe the result. What compound do the fumes suggest?

(*d*) Verify the observations and conclusions by repeating (*b*) and (*c*) with ammonium chloride from the laboratory bottle.

FIG. 61. — Convenient dropping apparatus — pointed tube open at both ends.

(*e*) Draw a conclusion from these observations regarding the main product of the interaction of ammonium hydroxide and hydrochloric acid.

The compound formed in Exp. 69 was ammonium chloride, and it belongs to a series of compounds known as the **Ammonium compounds.** If sulphuric acid or nitric acid had been used in Exp. 69, then ammonium sulphate or ammonium nitrate would have been formed. The ammonium series of compounds is strictly analogous to the sodium and other metallic series, *i.e.* ammonium and sodium enter into similar chemical reactions. Thus, these two series of compounds are well known:—

| | |
|---|---|
| $NaCl$ | $NH_4Cl$ |
| $NaOH$ | $NH_4OH$ |
| $NaNO_3$ | $NH_4NO_3$ |
| $Na_2SO_4$ | $(NH_4)_2SO_4$ |
| $Na_3PO_4$ | $(NH_4)_3PO_4$ |

Ammonium, however, is not an element, like sodium, but a compound; and not a compound, strictly speaking, but a group of atoms — one atom of nitrogen and four atoms of hydrogen — so related to each other and to other atoms, that they act as one atom of a metal. Such a group of atoms, which is incapable of independent existence, is called a **radical.** The chemical relations of radicals are exactly the same as elements. Ammonium has never been isolated, or if it has, it is so unstable that it immediately decomposes into ammonia and hydrogen. It is believed that ammonium hydroxide is a solution of an extremely unstable compound. The other compounds of ammonium mentioned above are well defined.

The symbol of ammonium is $NH_4$, and its valence is one. The formulas of its compounds have already been given.

### CLASS-ROOM EXERCISE. XXII.

1. Further study of ammonia gas.
    - (*a*) Source of atmospheric ammonia gas.
    - (*b*) Specific gravity.
    - (*c*) Its relation to the subject of combustion.
    - (*d*) Solubility in water.
    - (*e*) Quantitative test for ammonia gas.
2. Liquid ammonia.
    - (*a*) Faraday's work.
    - (*b*) Present method.
    - (*c*) Properties.
    - (*d*) Use in manufacture of ice.
    - (*e*) Formula. Why?
3. Historical.
    - (*a*) Early knowledge of ammonia gas and its compounds.
    - (*b*) Discovery of the gas by Priestley.
4. Ammonium hydroxide.
    - (*a*) Exact meaning of the name as usually used.
    - (*b*) Source of commercial substance.
    - (*c*) Specific gravity and strength of commercial substance.
    - (*d*) What is the difference between liquid ammonia and ammonium hydroxide?
    - (*e*) Uses for domestic purposes, in medicine, and in agriculture.
    - (*f*) What is anhydrous ammonia?
5. Names.
    - (*a*) Why is ammonia gas sometimes called *alkaline air*, *volatile alkali?*
    - (*b*) Why should ammonium chloride be called *sal-ammoniac*, muriate of ammonia, *sal-ammoniacum*, chloride of ammonia?
    - (*c*) What different meanings may the term *ammonia* have?
    - (*d*) Ammonium carbonate is sometimes called *sal-volatile*. Why?
    - (*e*) What is *liquor ammoniae?*
    - (*f*) What is ammoniacal liquor?
    - (*g*) What is gas liquor?
    - (*h*) Is there any difference between lime, quicklime, and calcium oxide? If so, what is it?

(*i*) What is *aqua ammonia?* Ammonia water? Spirits of hartshorn?

6. Ammonium salts.

(*a*) Give the formula, method of preparation, properties, and uses of ammonium chloride, ammonium nitrate, and ammonium sulphate.

7. Composition of ammonia gas.

(*a*) Summarize the experimental evidence of the qualitative composition of ammonia gas.

(*b*) Summarize the evidence of the volumetric composition of ammonia gas.

(*c*) Why is $NH_3$ the simplest formula of ammonia gas?

(*d*) When ammonia gas is passed over red-hot copper oxide, what are the products? What would the experiment prove in regard to the composition of ammonia gas, if the copper oxide and the products were weighed?

8. Miscellaneous.

(*a*) What is the valence of nitrogen in ammonia gas? In ammonium? In ammonium hydroxide?

(*b*) Calculate the atomic weight of nitrogen from data already given.

(*c*) Why is the formation of white fumes from the combination of ammonia gas and another gas not necessarily evidence of ammonium chloride?

9. Gay-Lussac's Law.

(*a*) Exact statement.

(*b*) Illustrations.

(*c*) Theoretical significance.

(*d*) Essential facts in the life of Gay-Lussac.

(*e*) Humboldt's contribution to the law.

## PROBLEMS. XV.

1. How many grams of ammonia gas can be obtained from 2140 gm. of ammonium chloride by heating with lime? (See Equation 16.)

2. What volume of ammonia gas can be obtained by heating 31.47 gm. of ammonium chloride with lime?

3. How many grams of ammonium chloride must be heated with lime to make 80 l. of ammonia gas?

4. Water absorbs about 50 per cent of its weight of ammonia gas at the ordinary temperature and pressure. How much ammonium chloride and calcium oxide are needed to produce 10 kg. of ammonium hydroxide?

5. If 22.4 l. of hydrogen at 0° C. and 760 mm. weigh 2 gm., what will an equal volume of ammonia gas weigh at 100° C. and 380 mm.?

6. If 250 cc. of chlorine are measured at 10° C. and 750 mm. and then allowed to interact with ammonia gas, what volume of nitrogen will be liberated at the same temperature and pressure?

7. How many cubic centimeters of the component gases can be obtained by the decomposition of a liter of ammonia gas?

8. A volume of dry ammonia gas equal to 29 cc. is decomposed by electric sparks, and the resulting mixture is exploded after 30 cc. of oxygen have been added. What gases remain, and what are their volumes?

9. 15 cc. of ammonia gas are decomposed by electricity, and 40 cc. of oxygen are added; the mixture is then exploded. What are the gases and their volumes before and after the explosion?

10. What weight of ammonia gas and of chlorine is necessary to produce a liter of nitrogen?

11. A kilogram of ammonium chloride is heated with lime. What volume of ammonia gas at 12° C. is produced?

12. 180 cc. of dry ammonia gas are decomposed by electric sparks.
   (*a*) What is the volume of each product?
   (*b*) If 130 cc. of oxygen are added and the mixture exploded, what gases and what volumes remain?

13. 100 cc. of dry ammonia gas are decomposed by electric sparks; then 100 cc. of oxygen are added and the mixture exploded.
   (*a*) What gases and what volumes remain after the decomposition of the ammonia gas?
   (*b*) After the explosion?

14. What is the weight of 30 l. of dry ammonia gas at the normal temperature and pressure?

15. What volume of nitrogen and of hydrogen can be obtained by passing electric sparks through one liter of dry ammonia gas, and what volume of oxygen will be necessary to unite with all the hydrogen thus obtained to form water?

16. Calculate the percentage composition of —

(*a*) Ammonium chloride, $NH_4Cl$.
(*b*) Ammonium hydroxide, $NH_4OH$.
(*c*) Ammonium sulphate, $(NH_4)_2SO_4$.
(*d*) Ammonium nitrate, $NH_4NO_3$.
(*e*) Magnesium nitride, $Mg_3N_2$.

17. Calculate the formula of a substance having the percentage composition: —

(*a*) Nitrogen = 82.35
Hydrogen = 17.64
(*b*) Nitrogen = 26.17
Hydrogen = 7.48
Chlorine = 66.35

18. How many pints of a solution of ammonium hydroxide (sp. gr. 0.88) weigh 4 lb.?

**Nitric Acid.** — Another useful compound of nitrogen is nitric acid.

### PREPARATION OF NITRIC ACID.

*Precaution — Do not get concentrated sulphuric or nitric acids on the skin or clothing.*

**Experiment 70.** — Prepare the acid in the apparatus shown in Fig. 62. *A* is a tubulated retort supported by a clamp and resting on a sand bath. The neck of the retort passes into the receiver *B*, which catches the nitric acid as it distils from the retort. Condensation of the acid fumes is facilitated by wrapping wet filter paper around the neck of the retort where it enters the receiver, and upon the flask itself, if it becomes hot. The neck is clamped just tightly enough to hold the retort in position. Turn *B* occasionally to expose a fresh surface to the fumes; it is advisable, though not always necessary, to place a block of wood against the bottom of the receiver to keep it in the desired position.

Slip about 35 gm. of sodium nitrate from a paper into the retort (see Fig. 101, App. A, § 12), and pour upon it, through a funnel, standing in the tubulure, about 25 cc. of concentrated sulphuric acid. Adjust the apparatus as shown in Fig. 62. Heat gently, and nitric acid will pass into the receiver Distil at as low a temperature as possible, as long as any nitric acid runs down the neck of the retort. Pour the nitric acid into a small bottle for use in the next experiment. Allow the retort to cool, add just enough warm water to loosen the solid mass, pour the contents into any convenient vessel, and preserve for Exp. 73.

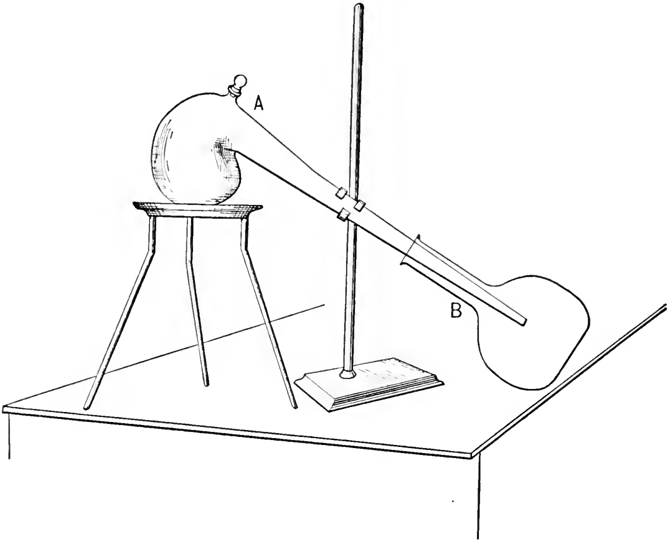

FIG. 62. — Apparatus for preparing nitric acid.

The general properties of nitric acid have been shown in a preceding experiment (see Exp. 48). They should be recalled or the work repeated.

### A SPECIAL PROPERTY OF NITRIC ACID.

**Experiment 71.** — Add twice its volume of water to the nitric acid made in Exp. 70, and proceed as follows: —

(*a*) Boil a piece of a quill toothpick in a portion of this diluted nitric acid. How is the quill changed at first? What is the effect of continued heating? Pour off the acid, and wash the quill with water. Is the color permanent?

(*b*) Add a dozen or more drops of nitric acid to a dilute solution of indigo. Describe the change. Will ammonium hydroxide restore the original color ? Is the change temporary or permanent ? What, in all probability, is the general character of the change — combination or decomposition ?

Draw a general conclusion from the whole experiment regarding the action of nitric acid on organic matter, which is typified by the quill and indigo.

### LABORATORY EXERCISE. VIII.

1. Examine a bottle of pure nitric acid and of the commercial acid, and record —

(*a*) Color of each acid.

(*b*) Presence or absence of a gas above the liquid, and color of gas, if any.

2. Examine a bottle of nitric acid which has been exposed to the sunlight. What does this observation show about the stability of nitric acid ?

3. Remove the stopper from a bottle of strong nitric acid. Is it very hygroscopic ? What other acid acts similarly ?

4. Smell of the acid *cautiously*, and record the result.

### A TEST FOR NITRIC ACID.

**Experiment 72.** — To a test tube one-fourth full of water add five or six drops of concentrated nitric acid ; add also an equal volume of concentrated sulphuric acid. Shake until the acids are well mixed, then cool by holding the test tube in running water. Make a cold, dilute solution of fresh ferrous sulphate and pour this solution carefully down the side of test tube upon the nitric acid mixture. Where the two solutions meet a brown or black layer will appear, consisting of a compound formed by the interaction of the nitric acid and the ferrous sulphate. It is an unstable compound and will often decompose, if the test tube is shaken. Record the observation.

This test is also used for a nitrate. Try it with a solution of sodium nitrate. Record the result.

**The Interaction of Sodium Nitrate and Sulphuric Acid.** — It has already been shown that nitric acid is one product of the interaction of sodium nitrate and sulphuric acid, and since chemists have shown that *if the action is complete* only one other substance is formed, it remains to determine its nature.

### THE INTERACTION OF SODIUM NITRATE AND SULPHURIC ACID.

**Experiment 73.** — Pour the solid residue obtained in Exp. 70 into a casserole or an evaporating dish, and evaporate to dryness over a piece of wire gauze in the hood. As the mass approaches pasty consistency lessen the heat to avoid spattering. When the mass is dry, heat strongly as long as white, choking fumes are evolved. This last operation is done to remove all traces of sulphuric acid, and to complete the chemical change. Allow the dish to cool gradually, and when cool, dissolve some of the white solid in distilled water and test separate portions for a sulphate and nitrate. Which is it? Test another portion for sodium, as previously directed. What is the name of the white substance?

Draw a general conclusion regarding the chemical action which occurs in the preparation of nitric acid by the interaction of sulphuric acid and sodium nitrate.

There are two equations expressing the chemical changes revealed by the preceding study of the preparation of nitric acid. If the temperature is low, then the reaction is represented thus: —

$$\underset{\substack{\text{Sodium}\\\text{Nitrate}}}{NaNO_3} + H_2SO_4 = HNO_3 + \underset{\substack{\text{Acid Sodium}\\\text{Sulphate}}}{HNaSO_4} \qquad (18)$$

$$85 + 98 = 63 + 120$$

But if the temperature is high, then the reaction is represented thus: —

$$2\,NaNO_3 + H_2SO_4 = 2\,HNO_3 + Na_2SO_4 \qquad (19)$$

$$170 + 98 = 126 + 142$$

**Composition of Nitric Acid.** — Numerous facts pointing to the qualitative composition of nitric acid are found by simple experiments.

### SYNTHESIS OF NITRIC ACID.

**Experiment 74.** — *Apparatus:* The apparatus consists of a small bottle provided with a two-hole rubber stopper; a platinum wire passes through each hole and is held firmly in place by a short piece of glass rod; each wire terminates in a narrow strip of platinum foil. There will be needed, also, a Ruhmkorff coil and a battery, or a static machine, and the usual connections for the battery.

*Process:* Put in the bottle a few drops of neutral water and a piece of blue litmus paper. Adjust the platinum terminals so that they are near the bottom of the bottle and from 3 mm. to 6 mm. apart. Connect the platinum wires with the coil, and pass sparks through the apparatus for about twenty minutes, or until there is definite evidence of a chemical change revealed by the litmus paper.

What is the nature of the change? Independent evidence has shown that the compound formed under these conditions is nitric acid. Since only air and water are present, what elements combine to form nitric acid?

### ANALYSIS OF NITRIC ACID.

**Experiment 75.** — (*a*) Generate hydrogen from zinc and dilute sulphuric acid in a large test tube or small flask, and when the evolution is brisk drop dilute nitric acid slowly into the vessel. The nitric acid may be dropped in from a burette or pipette, or from a glass tube (15 cm. long) shaped like Fig. 61. A brown gas may appear at intervals. If so, it indicates a secondary and undesirable action. The difficulty may be prevented by (1) regulating the evolution of hydrogen so that it is continuous but not too rapid, and (2) by adding the nitric acid slowly. The essential point of the experiment is to allow the nitric acid to interact with the hydrogen, but not to be decomposed by interaction with the zinc. Add about 10 cc. of nitric acid, then pour off the liquid from any excess of zinc, and evaporate the liquid to dryness; heat at first over the free flame in a casserole or in a dish which stands on a piece of wire gauze, and finally heat over a low flame which is kept constantly moving to prevent spattering.

Put a portion of the residue in a test tube and warm gently with sodium hydroxide solution. A thick, white precipitate will be formed, hence the tube must be shaken vigorously and the heat applied with care. A gas will be liberated. Test it by determining the odor and by holding a piece of moist red litmus paper *in the escaping vapor* — not against the side of the test tube which may have been covered with the sodium hydroxide. What is the gas? What is its source? What evidence does it furnish regarding the elements contained in nitric acid?

(*b*) *Perform this experiment in the hood.*

Boil a few grams of coarsely powdered ferrous sulphide in a test tube with a little concentrated nitric acid. Allow the action to proceed a few minutes, and after the tube is cool, dilute with two or three times its volume of distilled water. Filter, and test the filtrate for a sulphate.

Remembering the essential difference between a sulphide and a sulphate, explain the general chemical change which has taken place. What evidence of the qualitative composition of nitric acid is contributed by this experiment?

### CLASS-ROOM EXERCISE. XXIII.

1. Historical.
    (*a*) The alchemists' knowledge of nitric acid.
    (*b*) Lavoisier's contribution to our knowledge of the composition of nitric acid.
    (*c*) Cavendish's determination of the exact composition of nitric acid.
    (*d*) Derivation and significance of the term *aqua fortis.*

2. Commercial nitric acid.
    (*a*) Preparation.
    (*b*) Specific gravity.
    (*c*) Impurities.
    (*d*) Uses.

3. Composition of nitric acid.
    (*a*) Summarize the experimental evidence of the composition of nitric acid.
    (*b*) If nitric acid gas is passed over heated copper, two of the three products are water and nitrogen. Of what elements must the nitric acid have been composed?

(*c*) If nitric acid is allowed to trickle through a red-hot porcelain tube, a gas is formed which vigorously supports combustion. What gas (probably) is it, and what does its formation show about the composition of nitric acid? About the stability of nitric acid?

(*d*) If a mixture of dry ammonia gas and air is passed over heated platinum sponge (*i.e.* over a mass of porous platinum), nitric acid is formed. What does this fact prove about the composition of nitric acid? (The platinum takes no chemical part in the reaction.)

(*e*) Salts of nitric acid are often found in the water which falls at the beginning of a thunder-storm. Why?

4. Miscellaneous.

(*a*) Why are there no acid nitrates?

(*b*) What is the valence of nitrogen in nitric acid?

(*c*) Explain and illustrate the statement "nitric acid is a powerful oxidizing agent."

## PROBLEMS. XVI.

1. (*a*) How much sulphuric acid is needed for the complete transformation of 606 gm. of potassium nitrate into nitric acid?

(*b*) How much sulphuric acid, if the reaction takes place at a low temperature?

2. Five hundred grams of potassium nitrate are transformed into nitric acid by heating with sulphuric acid at a low temperature. Calculate —

(*a*) The amount of nitric acid produced.

(*b*) The amount of sulphuric acid required.

(*c*) The amount of acid potassium sulphate formed.

3. What is the least quantity of sulphuric acid which can be used to decompose 500 gm. of potassium nitrate in the preparation of nitric acid?

4. How many pounds of nitric acid can be obtained by distilling 400 lb. of sodium nitrate with sulphuric acid?

5. How many grams of nitric acid can be obtained by heating a kilogram of sodium nitrate with sulphuric acid at a low temperature?

6. How much (*a*) sodium nitrate and (*b*) potassium nitrate are necessary to produce the maximum weight of nitric acid by heating each with 140 kg. of 97 per cent sulphuric acid?

7. Calculate the percentage composition of—

   (*a*) Nitric acid.
   (*b*) Potassium nitrate.
   (*c*) Sodium nitrate.
   (*d*) Acid sodium sulphate.
   (*e*) Acid potassium sulphate.

**The Interaction of Nitric Acid and Metals.** — The interaction of nitric acid and metals is complicated. It depends upon (1) the particular metal, (2) the strength of the acid, (3) the temperature, and (4) the presence of salts resulting from the chemical action.

### THE GENERAL CHARACTER OF THE INTERACTION OF NITRIC ACID AND METALS.

*Perform this experiment in the hood.*

**Experiment 76.** — Stand four test tubes in the test tube rack and slip into each a few small pieces of one of the following metals: zinc, copper, tin, and iron. Add to each test tube in succession just enough concentrated nitric acid to cover the metal. Observe the changes in each case, particularly (1) the vigor of the action, (2) the nature and properties of all the products, especially color and solubility, and (3) evidence of presence or absence of hydrogen.

Tabulate these observations.

The interaction of nitric acid and copper typifies the chemical action in many cases and demands additional examination.

### THE INTERACTION OF NITRIC ACID AND COPPER.

*Perform this experiment in the hood.*

**Experiment 77.** — I. Construct an apparatus like Fig. 30 and arrange it to collect a gas over water as in the preparation of hydrogen (Exp. 19). Put from 10 to 15 gm. of copper borings in the flask, insert the stopper tightly, adjust the delivery tube, fill three bottles with water, and invert them in the trough. Pour just enough concentrated nitric acid through the safety tube into the flask to cover the copper, taking care to seal the bend of the safety tube with acid. Dense brown fumes are evolved. If the action is too vigorous, add a little water through the safety tube. Collect three bottles of the gas which bubbles from the delivery tube. Cover them with glass plates and stand them aside temporarily.

Pour the contents of the flask into a casserole or evaporating dish, and evaporate slowly to crystallization (not to dryness) on a water bath. If there is much of the original copper remaining, add more nitric acid at intervals. The crystals, after being dried between filter paper, should be preserved in a glass-stoppered bottle.

II. Meanwhile study the gas as follows:—

(*a*) Observe its general properties while covered.

(*b*) Uncover a bottle. Describe the result. Is the brown gas identical with the one observed in the generator at the beginning of the experiment?

(*c*) Uncover a bottle, pour in about 25 cc. of water, cover with the hand and shake vigorously, still keeping the bottle covered. Why has the brown gas disappeared? Uncover the bottle for an instant, then cover and shake again. Is the result the same? Repeat, if the result is not definite, or does not agree with previous observations.

(*d*) With the third bottle determine whether the two gases will burn or support combustion. A convenient flame is a burning match fastened to a stiff wire. Plunge it to the bottom at first and gradually raise it into the brown gas.

Answer the following:—

(1) What is the source of the colorless gas?

(2) What is the general chemical relation of the two gases to each other? To the air?

(3) Why is not the brown gas collected in the bottles by displacement of water?

(4) Will either gas burn, or support combustion?

(5) Which gas has been observed before? In what experiment (other than Exp. 76)?

(6) What is the general relation of these gases to nitric acid?

**III.** Study the properties of the crystals by determining: —

(*a*) Solubility in water (cold and hot).
(*b*) Action of heat.
(*c*) Action of their solution upon an iron nail.
(*d*) Action of their solution when added to ammonium hydroxide.

Compare the observed properties with those of copper nitrate obtained from the laboratory bottle. Are the two substances identical?

**Discussion of Experiment 77.** — The colorless gas is nitric oxide, the brown gas is nitrogen peroxide, and the blue solid is copper nitrate. The following equation is usually given as the simplest expression of the interaction of copper and nitric acid: —

$$\underset{\text{Copper}}{3\,Cu} + 8\,HNO_3 = \underset{\text{Copper Nitrate}}{3\,Cu(NO_3)_2} + 4\,H_2O + \underset{\text{Nitric Oxide}}{2\,NO}$$

When the nitric oxide is exposed to the air, it combines with oxygen. This reaction is most simply represented by the equation —

$$\underset{\text{Nitric Oxide}}{NO} + O = \underset{\text{Nitrogen Peroxide}}{NO_2}$$

Nitric oxide is not always the only gas produced by the interaction of nitric acid and copper, but it is most abundantly produced and is loosely regarded as the only oxide of nitrogen evolved. In the case of most metals, one or more of the numerous oxides of nitrogen are produced. Concentrated nitric acid and zinc yield ammonia, as well as nitric oxide, among the gaseous products, though the ammonia usually combines at once with the excess of acid to form ammonium nitrate. Some metals, such as tin and antimony, form an oxide and not a nitrate. Nitric acid is such a powerful oxidizing agent that hydrogen which is displaced from it by metals never appears as hydrogen, but is oxidized to water. (See page 117.)

**Nitrates.** — The salts of nitric acid have one common property, viz. extreme solubility in water, and hence many laboratory solutions are nitrates of the metallic elements. Nitrates behave in various ways when heated.

### ACTION OF NITRATES WITH HEAT.

**Experiment 78.** — I. Heat a little sodium nitrate in an ignition tube. A high temperature is necessary to produce any extensive chemical change. Insert a glowing match into the ignition tube. What gas is detected? Dissolve the residue in water and add a few drops of dilute sulphuric acid. Describe the result. Could the residue have been a nitrate? Why? If the experimental evidence is not definite enough for a final conclusion, proceed as follows: —

Heat a mixture of 10 gm. of sodium nitrate (or potassium nitrate) and 20 gm. of lead in a sand bath pan, which stands on a tripod, or similar support. Stir the melted mass with a stiff iron wire or blunt glass rod. Some of the lead will disappear and a yellowish brown powder will be seen in the molten mass. The action should proceed until most of the lead has disappeared. Allow the mass to cool, transfer to a mortar, pulverize, add hot water, and filter the clearer portion; add more hot water to the residue and filter this portion. Add to the combined filtrates a few drops of concentrated sulphuric acid. Describe the result. Compare the result with the action of concentrated sulphuric acid on a solution of sodium nitrite. Is the white residue (produced in the ignition tube) a nitrate? Why?

The yellowish product is lead oxide. What general chemical change led to its formation? How must the nitrate have been changed?

Draw a general conclusion from the observations made above regarding the action of heat on nitrates of the alkali metals.

II. Pulverize 8 or 10 gm. of lead nitrate, and heat the powder in an ignition tube or evaporating dish. Describe the result. How does it differ essentially from the action of heat on the alkali nitrates?

Two gaseous products are formed, though one — the oxygen — is not easily detected in the presence of the other. More definite results are obtained as follows: Heat the lead nitrate in an ignition tube provided with a delivery tube passing to the bottom of a test tube half full of sodium hydroxide solution and connected by a delivery tube with another test tube, or small bottle, arranged to collect a gas over water. The gas will be found to be oxygen. Add a few drops of concentrated sulphuric acid to the test tube which contained the sodium hydroxide solution. What compound is present? What chemical change produced it? The residue in the ignition tube is lead oxide. Compare these results with those obtained in Exp. 76, III. (*b*).

Draw a general conclusion regarding the action of heat on the nitrates of the heavier metals.

The exceptional behavior of ammonium nitrate when heated places it in a class by itself. An examination of its behavior involves also a study of the main product, which is another oxide of nitrogen.

### THE ACTION OF AMMONIUM NITRATE WITH HEAT.

**Experiment 79.** — The apparatus is shown in Fig. 63. The ammonium nitrate is heated in *A*, which is connected with an *empty* bottle, *D*; the latter is provided with a delivery tube, *E*, which dips into a pneumatic trough, filled, as previously described, with moderately warm water.

Fill *A* one-third full of crystals of ammonium nitrate, adjust the apparatus, and heat *A* gently with a low flame. The ammonium nitrate melts and appears to boil. Regulate the heat so that the evolution of gas will be slow. Notice the continuous formation of another product besides the gas. Collect three bottles of the gas, *free from air*, cover

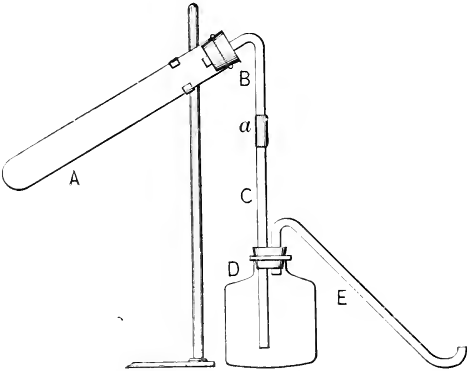

FIG. 63. — Apparatus for decomposing ammonium nitrate.

each with a glass plate, and stand them aside until needed. As soon as the last bottle has been removed from the trough, disconnect the apparatus at *a*, and stand the generator in the hood to cool.

Test the gas as follows: —

(*a*) Allow a jar to remain uncovered for a few seconds. How does this oxide of nitrogen differ from nitric oxide?

(*b*) Thrust a glowing stick of wood into the same bottle of gas. Describe the result. Is the gas combustible? Does it support combustion?

(*c*) The observations in (*b*) suggest that the gas is oxygen, but it is not, though this fact is not easily proved by a single experiment. Put a small piece of sulphur in a deflagrating spoon, light it, and lower the burning sulphur *at once* into another bottle of gas. If the experiment is conducted properly, the sulphur will not burn so brightly as it would in a bottle of oxygen.

(*d*) Stand the other bottle mouth downward in the pneumatic trough, or better, in a vessel of cold water. Describe the result. If the result is not conclusive, fill the bottle half full of water, cover with the hand, and shake. Would this observation help distinguish the gas from oxygen?

What in all probability is the other product of the chemical change in this experiment? Could it have been an impurity in the ammonium nitrate?

How would you distinguish ammonium nitrate from all other nitrates? How would you distinguish the third oxide of nitrogen — nitrous oxide — from (*a*) the other oxides of nitrogen, (*b*) air, (*c*) oxygen, (*d*) hydrogen, (*e*) nitrogen, (*f*) carbon dioxide?

### CLASS-ROOM EXERCISE. XXIV.

1. Summarize the properties of the three oxides of nitrogen studied.
2. State briefly the effect of heat on (*a*) potassium nitrate, (*b*) copper nitrate, (*c*) ammonium nitrate.
3. Predict the simplest equation for the reaction in the case of (*a*) sodium nitrate, (*b*) copper nitrate, (*c*) lead nitrate, (*d*) zinc nitrate, (*e*) ammonium nitrate, when heated. Verify the equations.
4. Select from Exp. 77, III. a test for copper.

The fact that all nitrates, except ammonium nitrate, yield oxygen when heated furnishes —

### A SPECIAL TEST FOR NITRATES.

**Experiment 80.** — Heat a piece of charcoal in the Bunsen flame, lay it on a board, or iron pan, and cautiously sprinkle powdered potassium nitrate upon the hot surface. Stand back when the action begins. Observe and describe the action, especially its violence and rapidity, also the color of the flame, the effect on the charcoal, and any other characteristic result.

This kind of chemical action is called *deflagration*. What causes it?

**Law of Multiple Proportions.** — There are two oxides of nitrogen besides the three already studied, and these five oxides aptly illustrate the Law of Multiple Proportions. This law is usually stated thus: —

*A fixed weight of one element so combines with different weights of another element that the relations between these different weights are expressed by small whole numbers.*

The composition of compounds is usually expressed in per cent. If, however, a definite weight is adopted as a unit for one component, and the composition is expressed in terms of this unit, the simple integral relation existing between the different proportions of the other element is clearly seen. The following table illustrates the law of multiple proportions: —

| Name. | Composition in Per Cent. | Unit Weight. | Ratio. |
|---|---|---|---|
| | N – O. | N. | N – O. |
| Nitrous oxide | 63.6 – 36.4 | 7 | 7 – 4 |
| Nitric oxide | 46.6 – 53.4 | 7 | 7 – 8 |
| Nitrogen trioxide | 36.8 – 63.2 | 7 | 7 – 12 |
| Nitrogen peroxide | 30.4 – 69.6 | 7 | 7 – 16 |
| Nitrogen pentoxide | 25.9 – 74.1 | 7 | 7 – 20 |

From this table it is clear that the proportions of oxygen in combination with a fixed weight of nitrogen are as 1 : 2 : 3 : 4 : 5. This law together with the law of definite proportions has profoundly influenced the development of the atomic theory of Dalton.

### CLASS-ROOM EXERCISE. XXV.

1. Gunpowder.
    - (*a*) Composition.
    - (*b*) General methods of manufacture.
    - (*c*) Kinds.
    - (*d*) Uses.
    - (*e*) Products of action.

2. Composition, manufacture, and uses of gun-cotton, nitroglycerine, and dynamite.

3. Miscellaneous.
    - (*a*) Define and illustrate nitrification.
    - (*b*) What are etchings and how are they made?

4. Law of Multiple Proportions.
    - (*a*) History.
    - (*b*) Exact meaning.
    - (*c*) Other illustrations.
    - (*d*) Theoretical significance.

5. What is the valence of nitrogen in $HNO_3$? In $Cu(NO_3)_2$? In $N_2O$, $NO$, $N_2O_3$, $NO_2$, and $N_2O_5$?

**Aqua Regia** is an old term still applied to a mixture of concentrated nitric and hydrochloric acids (one volume of the former to three of the latter).

### THE CHARACTERISTIC PROPERTY OF AQUA REGIA.

**Experiment 81.** — Touch a small piece of gold leaf with the end of a moist glass rod, and wash the gold leaf into a test tube by pouring a few cubic centimeters of concentrated hydrochloric acid down the rod. Heat gently until the acid just begins to boil. Does the gold dissolve?

Wash another piece of gold leaf into another test tube with concentrated nitric acid, and heat as before. Does the gold dissolve?

Pour the contents of one tube into the other, and warm gently. Does the gold dissolve? Draw a conclusion.

Answer the following: —

(1) What is the literal meaning and significance of the term *aqua regia?*

(2) What other metals does *aqua regia* dissolve?

(3) What is the chemical action of *aqua regia* on gold?

(4) Upon what property of nitric acid does the action of *aqua regia* depend?

## PROBLEMS. XVII.

1. Find the simplest formula of the substances having the indicated composition: —

(*a*) H = 1.58
N = 22.22
O = 76.19

(*b*) O = 47.52
N = 13.86
K = 38.61

2. What weight of pure nitric acid would yield 100 gm. of oxygen if completely decomposed?

3. The specific gravity of nitric acid is 1.522. (*a*) What will 100 cc. weigh? (*b*) What volume must be taken to weigh 100 gm.?

4. What volume is occupied by 10,000 gm. of nitric acid? (Assume density = 1.5.)

5. One gram of gunpowder yielded 280 cc. of gas at the normal temperature and pressure. At the instant of explosion the temperature rose to 2000° C. Calculate the volume occupied at this temperature (assuming an unchanged pressure).

# CHAPTER XII.

## CARBON AND ITS SIMPLEST COMPOUNDS.

CARBON is a constituent of every living thing. It forms a vast number of compounds. It is often the essential, and sometimes the only, constituent of the remains of animals and vegetables.

### DISTRIBUTION OF CARBON.

**Experiment 82.** — (*a*) Cover the bottom of a Hessian crucible with a thin layer of sand. Put on the sand a small piece of wood, a small, compact wad of cotton, a small bone, and a lump of starch. Fill the crucible loosely with dry sand, and slip it into the ring of an iron stand. Heat with a flame which extends just above the bottom of the crucible until the smoking ceases (approximately a half hour). After the crucible has cooled sufficiently to handle, pour the contents out upon a block of wood or an iron pan. Examine the contents. What is the residue? What is hereby shown about the distribution of carbon?

(*b*) Heat about 1 gm. of sugar in an old test tube until the vapors cease to appear. What is the most obvious product?

(*c*) Close the holes at the bottom of a lighted Bunsen burner, and hold a piece of crayon or glass tubing in the upper part of the flame long enough for a thin deposit to form. Examine it, name it, and state its source.

(*d*) Hold a piece of clean crayon in the flame of a candle which stands on a block of wood, and compare the result with that in (*c*).

(*e*) Light a wax taper and note the most obvious product of the flame. Explain. If time permits, heat in separate test tubes, or on an iron pan, a piece of meat, a little flour, a bit of albumen, a piece of bread What is the most obvious product in each case?

Draw a general conclusion regarding the distribution of carbon.

Carbon in the form of wood charcoal and of animal charcoal (often called *bone black*) possesses remarkable properties.

### DECOLORIZING ACTION OF CHARCOAL.

**Experiment 83.** — Fill a test tube one-fourth full of animal charcoal (see App. A, § 11), add 10 cc. of indigo solution, shake thoroughly for a minute, and then warm gently. Filter through a wet filter paper into a clean test tube. Compare the color of the filtrate with that of the indigo solution. Explain the change in color.

Other organic substances besides indigo are similarly changed. Draw a general conclusion regarding the decolorizing power of charcoal.

### DEODORIZING ACTION OF CHARCOAL.

**Experiment 84.** — Smell of a weak solution of hydrogen sulphide gas. Fill a test tube half full of powdered wood charcoal, add 5 cc. of hydrogen sulphide solution, and cork securely. If the tube leaks, make the opening gas-tight with vaseline. Shake thoroughly. After fifteen or twenty minutes, remove the stopper and smell of the contents. Is the odor much less offensive? Repeat, unless a definite result is obtained. Explain the change.

**The decolorizing and deodorizing power of charcoal** is largely a physical operation, and is mainly due to its porosity.

The attraction of carbon for oxygen — already shown by experiment and often observed in various forms of combustion — permits

### REDUCTION BY CARBON.

**Experiment 85.** — Prepare an apparatus like that shown in Fig. 23. Fill the ignition tube half full of a mixture of copper oxide (6 parts) and powdered wood charcoal (1 part); introduce the mixture by the method described in App. A, § 11. Arrange the apparatus as directed in Exp. 12. The delivery tube in this case, however, should dip into a small dish of lime water. Heat the whole ignition tube gently at first; increase the heat gradually, and finally heat strongly that part of the ignition tube containing the mixture. If the mixture "crawls" up the ignition tube, tap the tube gently near the top. A gas will be evolved. Heat until a decided change is produced in the lime water, and then immediately remove the end of the delivery tube from the liquid. Describe the change in the lime water. What caused the change? Draw a conclusion regarding chemical change which occurred in the ignition tube. Examine the contents of the tube to verify the conclusion. If the verification is not decisive, proceed as follows: —

Fill a porcelain crucible half full of the original mixture, and cover it with a thin layer of charcoal. Cover the crucible and stand it on a pronged tripod, or a triangle, and heat for five or ten minutes. Let the crucible cool, still covered, and, when cool enough to handle comfortably, turn out and examine the contents. Does the observation verify the previous conclusion? If the chemical change in the crucible had been complete, and the proportions absolutely correct, what would the residue have been?

### LABORATORY EXERCISE. IX.

1. Complete the equation —

$$CuO + C = CO_2 +$$

2. Define reduction in terms of Exp. 85.
3. What other element will reduce copper oxide?

### CLASS-ROOM EXERCISE. XXVI.

1. Allotropism or Allotropy.
   (*a*) Definition.
   (*b*) Illustration by means of the allotropic modifications of carbon
   (*c*) Reason for existence of allotropes.
   (*d*) Literal meaning and significance of the word *allotropy*.
2. Diamond.
   (*a*) Occurrence.
   (*b*) Artificial preparation.
   (*c*) Properties.
   (*d*) Famous diamonds.
3. Graphite.
   (*a*) Occurrence.
   (*b*) Preparation.
   (*c*) Properties.
   (*d*) Uses.
   (*e*) Literal meaning of the word.
   (*f*) What is "black lead"? Plumbago?

4. Animal charcoal.
    - (*a*) Preparation.
    - (*b*) Properties.
    - (*c*) Average composition.
    - (*d*) Uses.
    - (*e*) To what property is its efficiency due?
    - (*f*) What is bone black? Why so called?

5. Charcoal.
    - (*a*) Various kinds, and how obtained.
    - (*b*) General properties.
    - (*c*) Uses.

6. Coal.
    - (*a*) Kinds.
    - (*b*) Relation of kinds to each other.
    - (*c*) General mode of formation.
    - (*d*) What are peat and lignite?
    - (*e*) Composition of typical coals (see App. C, Table XI.).

7. Give the method of preparation (or source), properties, and uses of the following forms of carbon:—
    - (*a*) Lampblack.
    - (*b*) Coke.
    - (*c*) Gas carbon.
    - (*d*) Soot.

8. Experimental.
    - (*a*) Examine a section of coal through a microscope.

9. Miscellaneous.
    - (*a*) How would you prove the chemical identity of the three allotropic modifications of carbon?
    - (*b*) What is black smoke?
    - (*c*) What industries utilize the decolorizing power of animal charcoal?
    - (*d*) Use of charcoal as a disinfectant.
    - (*e*) What immense industry utilizes the principle illustrated in Exp. 85?

**Carbon Dioxide** is the best known of the simpler compounds of carbon.

### PREPARATION AND PROPERTIES OF CARBON DIOXIDE.

**Experiment 86.** — *Method:* Prepare carbon dioxide from hydrochloric acid and calcium carbonate, and study the gas collected over water.

*Apparatus:* Use the same apparatus as in the preparation of hydrogen (see Exp. 19 and Fig. 30). Other materials needed are lumps of marble, sand, concentrated hydrochloric acid, stick or splinter of wood, candle fastened to a wire, lime water, five bottles.

*Process:* Cover the bottom of the flask with sand, add a little water, and carefully slip into it a dozen or more small lumps of marble. Arrange the apparatus to collect the gas over water, as previously directed. Add through the safety tube just enough concentrated hydrochloric acid to cover the marble. Reject the first portion of the gas evolved. Collect five bottles, cover with glass plates or wet filter paper, and stand aside till needed.

Allow the action in the flask to continue, and preserve the contents for subsequent examination.

I. Study the properties of carbon dioxide gas as follows: —

(*a*) Plunge a burning stick into one bottle. Describe the result.

(*b*) Lower a lighted candle into a bottle of air, and invert a bottle of carbon dioxide over it, holding the bottles mouth to mouth. Describe the result. What does this result show about the specific gravity of carbon dioxide? Devise a simple experiment to verify this conclusion, using for that purpose one of the remaining bottles of carbon dioxide.

(*c*) Pour a little lime water into a bottle of carbon dioxide, cover with the hand, and shake vigorously. Describe and explain the result.

(*d*) Fill a bottle of carbon dioxide one-third full of water, cover with the hand, and shake vigorously. Invert, still covered, in a vessel of water. Does the result reveal any facts about the solubility of carbon dioxide?

II. Filter the contents of the flask into a casserole or evaporating dish, adding a little warm water beforehand, if the contents is solid. Evaporate to dryness in the hood over a free flame as long as much liquid remains. As the residue approaches pasty consistency, add a little water and continue the evaporation. If a casserole is used, move it about rapidly to avoid spattering. If an evaporating dish is used, stand it on a gauze-covered support and move the lighted burner underneath. Heat the residue until no fumes of hydrochloric acid are evolved. Dissolve some of the residue in distilled water and test portions for (*a*) a carbonate, (*b*) a chloride, (*c*) a calcium compound (see Exp. 29). If a calcium compound is found, confirm the observation thus: —

Dip a clean, moist platinum test wire into the solid residue, and hold it in the Bunsen flame. If calcium is present, the flame will be colored a yellowish red.

What is the residue? Verify the conclusion by a simple experiment.

The simplest equation for the interaction of hydrochloric acid and calcium carbonate is:—

$$\begin{array}{cccccccccc} 2\,HCl & + & CaCO_3 & = & CO_2 & + & CaCl_2 & + & H_2O & \\ & & \text{Calcium} & & \text{Carbon} & & \text{Calcium} & & & \\ & & \text{Carbonate} & & \text{Dioxide} & & \text{Chloride} & & & (20) \\ 73 & + & 100 & = & 44 & + & 111 & + & 18 & \end{array}$$

The qualitative composition of carbon dioxide may be shown by the

### SYNTHESIS OF CARBON DIOXIDE.

**Experiment 87.** — *Method:* Draw, simultaneously, purified air over hot charcoal and the product of the reaction through barium hydroxide solution or lime water.

*Apparatus:* The essential part of the apparatus is shown in Fig. 64. Short pieces of sodium hydroxide are placed in the left-hand limb of *A* and calcium chloride in the other, the two substances being separated by a wad of cotton; small wads of cotton are also placed just below the stopper in each limb. This U-tube removes the carbon dioxide and water vapor from the air. The tube *BB'* is hard glass and from 15 to 20 cm. long. A spiral of copper wire is attached at the point *B'*. The test tube *C* is connected directly with the combustion tube by a glass

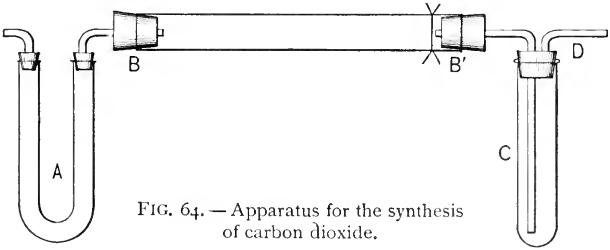

FIG. 64.—Apparatus for the synthesis of carbon dioxide.

tube passing to the bottom of the test tube; the tube *D* serves as an outlet for the excess of air and is attached directly to a filter pump or aspirator (see Fig. 107). The apparatus may be supported by clamps at the points *B* and *B'*.

*Process:* Place three or four lumps of dry wood charcoal in the combustion tube. Fill the absorption tube half full of barium hydroxide solution. Connect as shown in the figure. Draw a slow current of air through the apparatus to detect the leaks, if any. Readjust the connections, if necessary. Regulate the current of air so that a bubble a second, approximately, is drawn through the apparatus. Heat the whole combustion tube gently at first, and finally concentrate the heat at the place where the charcoal is located. A wing-top burner is well adapted for this operation. In a few minutes (approximately 10) a marked change is seen in the absorption tube. Describe it. Since the nitrogen of the air takes no part in the chemical change, explain the change. What does the experiment prove about the composition of carbon dioxide.

The simplest equation for the synthesis of carbon dioxide is:—

$$\begin{array}{ccccc} C & + & O_2 & = & CO_2 \\ 12 & + & 32 & = & 44 \end{array} \qquad (21)$$

**A liter of carbon dioxide** under standard conditions weighs 1.977 gm.

Carbon dioxide is vitally connected with the subject of combustion.

### CARBON DIOXIDE AND COMBUSTION.

**Experiment 88.** — (*a*) Exhale through a glass tube into a test tube half full of lime water. Describe and explain the result.

(*b*) Lower a lighted candle into a bottle and allow it to burn for a few minutes. Remove the candle, pour a little lime water into the bottle, and shake vigorously. Describe and explain the result.

(*c*) Allow a stick of wood to burn for a short time in a bottle, and then proceed as in (*b*). Describe the result. Does it confirm the results in (*a*) and (*b*)?

Answer the following: —

(1) What is the main product of the combustion of substances containing carbon?

(2) What gas was formed in Exp. 14 (*c*)?

(3) How does this experiment verify previous work on "burning in the air"?

(4) How does (*a*) show the relation between carbon dioxide and combustion?

### CLASS-ROOM EXERCISE. XXVII.

1. Additional study of carbon dioxide.
   (*a*) Occurrence. (*b*) Modes of formation.
   (*c*) Properties not shown experimentally.
   (*d*) Liquid carbon dioxide. (*e*) Solid carbon dioxide.
2. Review.
   (*a*) Combustion.
   (*b*) Carbonated mineral waters. (*c*) Respiration and decay.
3. Historical.
   (*a*) Van Helmont's work on carbon dioxide.
   (*b*) Black's work. (*c*) Lavoisier's work.
   (*d*) Why was carbon dioxide called *gas sylvestre* and *fixed air?*
4. Miscellaneous.
   (*a*) How would you distinguish carbon dioxide from all other gases?
   (*b*) Summarize the most important properties of carbon dioxide.
   (*c*) What is the decisive test for carbon dioxide?
   (*d*) What happens when a burning stick is plunged into oxygen?
   Into nitrogen? Into hydrogen? Into carbon dioxide?
   (*e*) What happens when lime water is poured into oxygen: Into hydrogen? Into nitrogen? Into carbon dioxide?

5. Deduce the molecular weight and simplest formula of carbon dioxide from the following data: —

   (*a*) A liter of carbon dioxide weighs 1.977 gm.
   (*b*) Carbon dioxide contains its own volume of oxygen.

6. Dumas and Stas found that 80 parts of oxygen by weight combined with 30 parts of carbon. Deduce the atomic weight of carbon from this fact together with the results obtained in 5 (*supra*).

## PROBLEMS. XVIII.

1. How many grams of calcium carbonate are necessary to produce 15 l. of carbon dioxide?

2. How much calcium carbonate and hydrochloric acid are needed to form 132 gm. of carbon dioxide?

3. What weight of carbon burned in air will produce 11 gm. of carbon dioxide?

4. Twelve grams of carbon were burned in the oxygen liberated from 122.5 gm. of potassium chlorate. How much carbon dioxide was formed, and what was the excess of oxygen?

5. An excess of air was passed over red-hot charcoal and thereby formed 21 l. of carbon dioxide. What volume of air was deprived of its oxygen?

6. A piece of pure graphite weighing 7 gm. is completely burned in oxygen? What volume of carbon dioxide is formed?

7. What volume of oxygen is necessary to burn a kilogram of carbon?

8. Eighteen grams of carbon are to be burned in air. How many liters of air are needed, and how many liters of carbon dioxide will be formed?

9. What volume of carbon dioxide at 12° C. and 750 mm. will be produced by the action of hydrochloric acid on 10 gm. of marble?

10. What weight of water must be decomposed to furnish sufficient oxygen to form, with pure carbon, 44 gm. of carbon dioxide?

11. How much oxygen by weight and by volume is required to unite with pure carbon to form 132 gm. of carbon dioxide?

12. How much carbon by weight is there in a liter of carbon dioxide?

13. How much carbon dioxide by weight and volume is in the air of a room 6 m. long, 4 m. wide, and 3 m. high, if there is one volume of carbon dioxide in 1000 volumes of air?

**Carbonic Acid.** — Carbon dioxide gas is often called *carbonic acid gas*, or simply *carbonic acid*. It is believed that when carbon dioxide is passed into water it combines with the water and forms a weak, unstable acid. The case is analogous to the formation of the base ammonium hydroxide by the solution of ammonia gas in water. Carbonic acid has never been isolated, but its salts, the carbonates, are numerous and well-defined compounds. The formula $H_2CO_3$ has been given to carbonic acid.

### CARBONIC ACID.

**Experiment 89.** — Construct a carbon dioxide generator like that shown in Fig. 46. *B* is filled nearly full of lumps of marble. Concentrated hydrochloric acid (1 vol. to 1 vol.) is put in *A*. The generator is operated as previously described. (See Exp. 44.) Attach *D* to a wash bottle (see Figs. 56, 65, and 114) containing water or concentrated sulphuric acid to free the gas from any hydrochloric acid carried over mechanically; connect the wash bottle with an absorption apparatus consisting of a large test tube or bottle; the tube from the wash bottle should reach to the bottom of the absorption apparatus. A simple combined generator and wash bottle is shown in Fig. 65.

Fill the absorption apparatus nearly full of water, add a few drops of a solution of phenolphthalein and just enough sodium hydroxide solution to color the liquid a faint pink. Allow a slow current of carbon dioxide to bubble through the apparatus until a definite change is produced in the absorbing liquid. Describe and explain it.

Sketch the essential part of the apparatus.

**Carbonates** are salts of the hypothetical carbonic acid. Many carbonates have already been studied. Additional experiments, however, are needed to illustrate the —

### FORMATION AND PROPERTIES OF CARBONATES.

**Experiment 90.** — (*a*) Pass carbon dioxide *free from acid* into a bottle full of lime water until considerable precipitate is formed. The generator shown in Fig. 46 may be used, if a wash bottle containing water or concentrated sulphuric acid is placed between the generator and the bottle of lime water. The combined generator and wash bottle shown in Fig. 65 also gives satisfactory results. The gas is generated in the large test tube, *A*, and washed in *B*, and any convenient delivery tube is attached to *C*; the apparatus may be stood in a test tube rack. When sufficient precipitate has formed, disconnect the generator, and allow the precipitate to settle. Decant the supernatant liquid, add a few drops of hydrochloric acid to its residue, and test the gas evolved with a burning match and with a rod moistened with lime water. What is the gas? What is the precipitate?

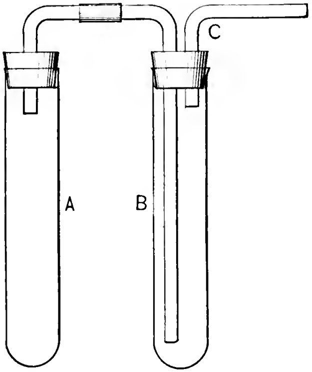

FIG. 65.—Apparatus for generating and washing carbon dioxide.

Complete the equations:—

$$CO_2 + Ca(OH)_2 = H_2O +$$
$$CaCO_3 + HCl = CaCl_2 +$$

(*b*) Pass carbon dioxide *free from acid* through a weak solution of sodium hydroxide (or potassium hydroxide) for several minutes. Add a few drops of hydrochloric acid to a portion of the liquid. Describe and explain the result, testing the main product of the action, if necessary. In what essential respect does the compound formed from the sodium hydroxide differ from the one formed from calcium hydroxide?

(*c*) Heat a little powdered magnesium carbonate in a test tube or ignition tube, and test (as in (*a*)) the gas evolved. Explain the result.

Complete the equation:—

$$\underset{\text{Magnesium Carbonate}}{MgCO_3} = MgO +$$

LABORATORY EXERCISE. X.

1. What is the decisive test for a carbonate?
2. In what two ways may carbonates be decomposed?
3. How are carbonates formed?
4. Complete the equations:—

(*a*) $KOH + CO_2 = K_2CO_3 +$

(*b*) $K_2CO_3 + HCl = CO_2 +$

5. How may lime water be easily distinguished from solutions of sodium and potassium hydroxide?

Carbonic acid is dibasic, hence, besides ordinary calcium carbonate, there is

### ACID CALCIUM CARBONATE.

**Experiment 91.** — Pass carbon dioxide *free from acid* into lime water until the precipitate disappears. Filter, if the liquid is not perfectly clear, and then heat. Describe the change. Why should the precipitate disappear before heating? Why reappear after heating? (If the latter question cannot be answered from evidence already given, consult a text-book.)

### LABORATORY EXERCISE. XI.

1. If $Na_2CO_3$ is the formula of sodium carbonate, what is the formula of acid sodium carbonate?
2. What is the formula of acid calcium carbonate?

### DETERMINATION OF CARBON DIOXIDE IN A CARBONATE.

**Experiment 92.** — *Method:* Decompose a known weight of a carbonate (preferably calcite, which is a pure variety of calcium carbonate) and calculate the loss in weight as carbon dioxide.

*Apparatus:* A 125 cc. Erlenmeyer flask is provided with a one-hole rubber stopper fitted with a small bulb trap as show in Fig. 66. The trap is about 45 mm. high, and is made by sealing a short tube into a bulb about 25 mm. in diameter. It is filled with fused calcium chloride held in place by absorbent cotton. Its object is to prevent the loss of water vapor which might be carried off mechanically by the escaping carbon dioxide. The whole apparatus should not weigh more than 60 gm. before adding the acid

FIG. 66. — Apparatus for determining the proportion of carbon dioxide in a carbonate.

*Process:* Clean and dry the flask. Fill the bulb trap as previously directed. Put in the flask 25 cc. of dilute hydrochloric acid (one volume of acid to two volumes of water). Slip into the flask a lump of calcite (crystallized calcium carbonate) weighing about one gram — its exact weight need not be known. Insert the stopper and allow the action to proceed without interruption. The object of this preliminary operation is (1) to fill the apparatus with carbon dioxide, since it is to be weighed finally in that condition, and (2) to allow the liquid to absorb carbon dioxide, so that only a very small amount will be retained in subsequent operations. Meanwhile weigh exactly to a centigram from 1.3 to 1.5 gm. of calcite. When the evolution of carbon dioxide has ceased, open the flask for an instant, close it, and weigh to a centigram. Slip in the weighed piece of calcite and insert the stopper. Allow the action to proceed until no more gas is evolved, then open the flask for an instant, close it, and weigh again.

Record the results as follows: —

| | Grams. |
|---|---|
| Weight of apparatus . . . . . . . . . . . . . | |
| Weight of calcite . . . . . . . . . . . . . . | |
| Total . . . . . . . . . . . . . . . . . . | |
| Final weight of apparatus . . . . . . . . . . | |
| Weight of carbon dioxide lost . . . . . . . . . | |
| Per cent of carbon dioxide in the calcite . . . . | |

SUMMARY.

Per Cent of Carbon Dioxide in Calcite.

| Found. | Theory. | Class Average. |
|---|---|---|
| 1. | | |
| 2. | | |
| 3. | | |
| Average. | | |

Several determinations may be made with the same apparatus without replacing the acid, if time permits a continuous operation. While the calcite is being decomposed, another piece may be weighed ready for use as soon as the "Final weight" is found, which then becomes, of course, the "Weight of apparatus" in the second determination.

**Carbon Monoxide** is another compound of carbon and oxygen. It may be produced by the

### ACTION OF HOT CHARCOAL ON CARBON DIOXIDE.

*The product of this experiment is a poisonous gas and should not be allowed to escape into the air.*

**Experiment 93.** — *Method*: Pass carbon dioxide over hot charcoal, collect over water the gaseous product, and test it.

*Apparatus:* The essential part of the apparatus is shown in Fig. 64. A carbon dioxide generator (see Fig. 46) is connected with the U-tube *A*, which is to be filled with calcium chloride. The combustion tube *BB′* is the same as that used in Exp. 87. The test tube *C* contains a solution of sodium hydroxide to absorb the excess of carbon dioxide. A delivery tube is connected with *D* and passes into a pneumatic trough arranged to collect a gas over water. The apparatus may be supported by clamps placed at *B* and *B′* or any other convenient points. It is advisable, though not absolutely necessary, to add a little sodium hydroxide solution to the water in the trough. Four or five bottles and several lumps of charcoal are also needed.

*Process:* Put three or four lumps of charcoal in the combustion tube and connect the different parts of the apparatus. Regulate the flow of carbon dioxide so that the rate is (approximately) a bubble a second. Stop the leaks, if any are detected. Heat the whole combustion tube at first, and finally concentrate the heat where the charcoal is located. Collect all the gas evolved, but reject the first two bottles, as they contain air (and possibly carbon dioxide). Collect two or three more bottles, cover with glass plates, and set them aside temporarily. Stop the generator, remove the delivery tube from the trough, and gradually cool the combustion tube. Test the gas thus: —

(*a*) Notice that it is colorless.

(*b*) Hold a lighted match at the mouth of a bottle for an instant. Note the flame, especially its color and how it burns. After the flame has disappeared, drop a lighted match into the bottle. Describe the result. Draw a conclusion and verify it by (*c*).

(*c*) Burn another bottle of gas, and after the flame has disappeared, pour a little lime water into the bottle and shake. Does the result verify the conclusion made in (*b*)? If not, repeat with another bottle of gas.

### LABORATORY EXERCISE. XII.

1. Summarize the observed properties of carbon monoxide.

2. What is the chemical relation of carbon monoxide to carbon dioxide?

3. How can each be changed into the other? What two general processes do the changes illustrate?

The simplest equation for the action of carbon dioxide on hot charcoal is

$$\begin{array}{ccccccc} CO_2 & + & C & = & 2\,CO & & \\ & & & & \text{Carbon Monoxide} & & (22) \\ 44 & + & 12 & = & 56 & & \end{array}$$

**A liter of carbon monoxide** under standard conditions weighs 1.25 gm.

If the properties of carbon monoxide cannot be conveniently studied by Exp. 93, they may be studied from the gas prepared by the usual method, as follows: —

### PREPARATION AND PROPERTIES OF CARBON MONOXIDE.

(*Optional.*)

**Experiment 94.** — Construct an apparatus similar to that used for the preparation of chlorine (see Fig. 52). The flask should be smaller, though this modification is not necessary, and the tube *B* should pass to the bottom of a wash bottle fitted with a two-hole rubber stopper. The wash bottle is filled two-thirds full of sodium hydroxide solution. A delivery tube passes from the wash bottle into a pneumatic trough arranged to collect a gas over water.

Put 10 gm. of crystallized oxalic acid (*oxalic acid is poisonous*) in the flask, insert the stopper with its tubes, see that all joints are tight, and then pour 30 cc. of concentrated sulphuric acid through the safety tube into the flask. Heat the flask gently, and carbon monoxide will be evolved. A small flame must be used, because the gas is rapidly evolved as the heat increases. It is advisable to remove or lower the flame as bubbles appear in the flask — *regulate the heat by the effervescence*. Collect all the gas, but do *not* use the first two bottles, covering the bottles with glass plates as they are filled, and setting them aside temporarily. Disconnect the generator at *a*, and stand it in the hood to cool. Test the gas as directed in Exp. 93.

Add a little acid to the wash bottle. What gas besides carbon monoxide was produced?

Sketch the essential part of the apparatus.

## CLASS-ROOM EXERCISE. XXVIII.

1. Additional study of carbon monoxide.
   - (*a*) Various methods of formation.
   - (*b*) Properties besides those revealed by experiment.
   - (*c*) Its reducing power.
   - (*d*) What is water gas?
   - (*e*) Danger from carbon monoxide.
2. Combustion and the oxides of carbon.
   - (*a*) Which oxide is formed in an excess of air? Which in an excess of carbon?
   - (*b*) State and explain the various chemical changes which occur from the entrance of oxygen (in the air) below the grate of a red-hot coal fire to the end of the burning of the carbon monoxide at the top of the coal.
3. Miscellaneous.
   - (*a*) How may a mixture of the oxides of carbon be separated into its constituents?
   - (*b*) What is the valence of carbon in carbon monoxide and in carbon dioxide? Is there any satisfactory explanation of this fact?
   - (*c*) Illustrate the Law of Multiple Proportions by the two oxides of carbon.
   - (*d*) Carbon dioxide is often called carbonic anhydride. Why? What experiment proves this?
   - (*e*) Calculate the molecular weight of calcium carbonate from the result obtained in Exp. 92. Compare with the theoretical weight.
   - (*f*) Devise an experiment to verify the result obtained in Exp. 92.

## PROBLEMS. XIX.

1. If 20 gm. of carbon are heated in the presence of 44 gm. of carbon dioxide, what weight of carbon monoxide is formed, and what weight, if any, of carbon remains?

2. What volume of carbon dioxide must be passed over red-hot charcoal to yield 159 l. of carbon monoxide?

3. How many liters of carbon dioxide must be passed over red-hot charcoal to yield 84 gm. of carbon monoxide?

4. What volume of carbon dioxide must be passed over what weight of carbon to produce 10 l. of carbon monoxide at 14° C. and 760 mm.?

5. What per cent of carbon by weight is contained in carbon monoxide? In carbon dioxide?

6. Carbon dioxide is passed into lime water and forms a precipitate weighing 5 gm. How much carbon dioxide can be obtained from this precipitate?

7. Calculate the percentage composition of—

(*a*) Calcium carbonate.
(*b*) Carbon monoxide.
(*c*) Carbon dioxide.
(*d*) Magnesium carbonate.

# CHAPTER XIII.

## COMPOUNDS OF CARBON AND HYDROGEN — ILLUMINATING GAS — FLAMES — OXIDATION AND REDUCTION.

**Hydrocarbons.** — Three important compounds of carbon and hydrogen are methane, ethylene, and acetylene. They are called hydrocarbons, and are the first members of three large classes of similar compounds.

### PREPARATION AND PROPERTIES OF METHANE, OR MARSH GAS.

**Precaution.** — *A mixture of methane and air explodes violently when ignited. Guard against flames and leaks.*

**Experiment 95.** — *General Directions:* Use the same apparatus as in Exp. 12. Fill the ignition tube nearly full of the mixture described below and insert a plug of glass wool, or asbestos between top of the mixture and the inner end of the delivery tube. Clamp the apparatus in the position shown in Fig. 23, and collect the gas over water.

*Supplies:* Heat a few grams of sodium acetate in a porcelain dish or iron pan until the water of crystallization has been expelled. Use a small flame and stir constantly toward the end of the operation. Prepare one of the following mixtures: —

(*a*) Three parts of dehydrated sodium acetate, and two parts soda lime.

(*b*) Equal parts of dehydrated sodium acetate, dry sodium hydroxide, and quicklime.

*Process:* Fill the tube with either one of the above mixtures and arrange the apparatus as previously described. Heat the whole ignition tube gently at first, increasing the heat gradually until a gas is evolved. Reject the first bubbles. Collect two small bottles of the marsh gas, and then remove the end of the delivery tube from the water. Hold a lighted match at the mouth of one of the bottles, and observe the nature of the flame, especially its color, luminosity, visible products (if any), temperature, etc. Repeat with the other bottle and carefully observe any properties needing confirmation. Pour a little lime water into one of the bottles and shake. Explain the result. What evidence of the composition of methane is presented by this experiment?

When methane burns in air, the combustion is represented by the equation: —

$$\begin{array}{ccccccc} CH_4 & + & 2O_2 & = & CO_2 & + & 2\,H_2O \\ \text{Methane} & & \text{Oxygen} & & \text{Carbon} & & \text{Water} \\ & & & & \text{Dioxide} & & \\ 1\text{ vol.} & + & 2\text{ vols.} & = & 1\text{ vol.} & + & 2\text{ vols.} \\ 16 & + & 64 & = & 44 & + & 36 \end{array} \tag{23}$$

### PREPARATION AND PROPERTIES OF ETHYLENE.

**Precautions.** — (a) *A mixture of ethylene and air explodes violently, if ignited. Guard against flames and leaks.*

(b) *Hot sulphuric acid burns frightfully. Guard against breakage.*

**Experiment 96.** — *Apparatus:* Use the apparatus shown in Fig. 63. Arrange it to collect a gas over water.

*Supplies:* Put 5 cc. of water in a test tube and slowly pour upon it 15 cc. of concentrated sulphuric acid. Cool the acid by holding the test tube in a stream of cold water. Put 5 to 7 cc. of alcohol in the test tube *A*, add a little clean sand, or a few glass beads, and then slowly pour in the *cold* acid. Two or three small bottles will be needed. A dish should stand under *A* to catch the contents, in case of accident.

*Process:* Adjust the apparatus as shown in the figure, taking care not to crush the test tube. Heat the test tube gently between the bottom and the surface of the contents to detect any leaks in the apparatus. Readjust, if necessary. Heat gently to drive out the air, and when it is judged that the gas which is being evolved is ethylene, collect two or three bottles. As the heat increases the mixture is apt to froth or "bump"; sometimes the gas is evolved suddenly. Hence, the heat must be so regulated that the evolution of gas is slow. Especial care must be taken not to heat the test tube above the surface of the contents, otherwise a sudden movement of the hot liquid might crack the tube. As soon as the gas has been collected, disconnect at *a*, and stand the generator in the hood to cool. When cool enough to handle, pour the contents down the sink or into a receptacle especially provided for dangerous mixtures.

Test the gas by holding a lighted match at the mouth of a bottle. Observe and record the color and temperature of the flame, its luminosity, rapidity of combustion, visible products, and any other characteristic properties. Add a little lime water to one of the bottles in which the gas was burned, shake, and explain the result.

What evidence does this experiment present regarding the composition of ethylene?

When ethylene burns, the combustion is represented by the equation: —

$$\begin{array}{ccccccccc} C_2H_4 & + & 3\,O_2 & = & 2\,CO_2 & + & 2\,H_2O & & \\ \text{Ethylene} & & & & & & & & (24) \\ 1 \text{ vol.} & + & 3 \text{ vols.} & = & 2 \text{ vols.} & + & 2 \text{ vols.} & & \\ 28 & + & 96 & = & 88 & + & 36 & & \end{array}$$

### PREPARATION AND PROPERTIES OF ACETYLENE.

**Experiment 97.** — Put about 10 cc. of water in a test tube, stand the test tube in a rack, and drop two or three very small pieces of calcium carbide into the test tube. Acetylene is evolved. After the action has proceeded long enough to expel the air, light the gas by holding a lighted match at the mouth of the tube. Observe and record the nature of the flame, especially its color, intensity, visible products (if any), temperature, etc. Hold a cold glass plate over the flame. What does the result suggest about the composition of acetylene? What other evidence of its composition is revealed by the properties previously observed?

### CLASS-ROOM EXERCISE. XXIX.

1. Hydrocarbons.
    - (*a*) Exact meaning of term.
    - (*b*) Reason for the vast number.
    - (*c*) Meaning of the term "homologous series."
    - (*d*) Illustrate by hydrocarbons the terms: *saturated, unsaturated, isomerism, polymerism, substitution product.*

2. Methane.
    - (*a*) Names, and why so called.
    - (*b*) Occurrence.
    - (*c*) Properties (additional).
    - (*d*) What is fire damp? choke damp?
3. Ethylene.
    - (*a*) Names, and why so called.
    - (*b*) Equation for the reaction in Exp. 96.
    - (*c*) Properties (additional).
    - (*d*) Liquid ethylene and its uses.
    - (*e*) One essential difference between ethylene and methane.
4. Acetylene.
    - (*a*) Synthesis, and historical significance of this method of formation.
    - (*b*) Industrial preparation from calcium carbide.
    - (*c*) Properties.
    - (*d*) Uses.
    - (*e*) Essential difference between methane, ethylene, and acetylene.
5. Miscellaneous.
    - (*a*) Illustrate the Law of Gay-Lussac by the facts revealed by the combustion of methane and ethylene.

## PROBLEMS. XX.

1. What volume of oxygen will be needed and what volume of carbon dioxide will be formed in the combustion of (*a*) 2 l. of methane, and (*b*) 2 l. of ethylene?

2. How many volumes of oxygen are required and how many volumes of carbon dioxide will be formed in burning 500 cc. of (*a*) ethylene and (*b*) methane?

3. What weight of air is necessary for the complete combustion of 1000 cc. of marsh gas at 10° C. and 752 mm.? What are the volumes of the products measured at 300° C. and 752 mm.?

4. What weight of air is needed for the complete combustion of 4 gm. of ethylene?

5. Calculate the percentage composition of—

(*a*) Marsh gas, $CH_4$.
(*b*) Ethylene, $C_2H_4$.
(*c*) Acetylene, $C_2H_2$.

**Illuminating Gas** is formed by the destructive distillation of coal. It is a mixture of many gases, and its composition varies with the coal employed. The constituents are divided into three classes, viz. illuminants, diluents, and impurities. The first class contains, among other gases, ethylene and acetylene. Marsh gas, hydrogen, and carbon monoxide, all of which burn with a feeble — non-yellow — flame, are the most common diluents. The common impurities are carbon dioxide, hydrogen sulphide, and ammonia; some of these, however, are entirely removed from the purified gas. Illuminating gas prepared as above is often called coal gas.

## PREPARATION AND PROPERTIES OF ILLUMINATING GAS.

**Experiment** 98. — *Apparatus:* The apparatus is shown in Fig. 67. *AA'* is an ignition tube from 10 to 15 cm. long. A spiral of copper wire is placed near *A'*, and the tube is supported by a clamp as near this end of the tube as convenient. An empty test tube or bottle is connected with the combustion tube by a bent tube passing to the bottom of *B*; this vessel retains tarry matter which comes from the ignition tube. The U-tube contains moistened pink litmus paper in the limb *C* and a narrow strip of filter paper moistened with a lead compound (nitrate or acetate) in the limb *C'*, the latter serving to detect hydrogen sulphide. The bottle *D* is connected as shown in the figure, and is to be

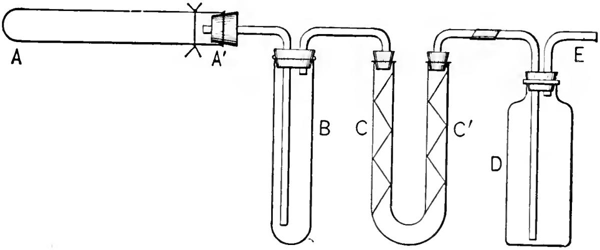

FIG. 67. — Apparatus for preparing illuminating gas.

one-third full of lime water. The tube *E* is to be connected with a delivery tube passing into a pneumatic trough arranged to collect a gas over water. Three or four bottles will be needed.

*Process:* Fill *AA′* two-thirds full of coarsely powdered soft coal which should be held in place with a loose plug of glass wool, or shredded asbestos. See that all connections are gas tight by heating the ignition tube gently; if the apparatus is tight, the expanded air will bubble through the bottle *D*. Readjust, if necessary.

Heat the whole ignition tube gently at first, and gradually increase the heat, but avoid heating either end very hot, otherwise the closed end may soften and burst or the rubber stopper may melt. As the heat increases, watch for marked changes in *B*, *CC′*, and *D*. As soon as the slow bubbling shows that all air has been driven out of the apparatus, collect, as previously directed, two or three bottles of the gas evolved. Cover the bottles with wet filter paper or glass plates as fast as they are removed from the trough. When the last bottle has been removed, disconnect the apparatus at any convenient point between *A′* and *C′* Let the ignition tube cool. Test the gas by bringing a lighted match near the mouth of a bottle. Observe and record the color and heat of the flame, the rapidity of combustion of the gas, whether or not smoke is formed, and any other characteristic property. Repeat with the remaining gas and observe more closely any facts suggested but not clearly revealed by the first observations.

Examine the contents of the ignition tube. Does it suggest coke or some form of carbon? Examine the bottle *B* for tarry matter. Does the paper in *C* reveal the presence of any ammonia? If the paper in *C′* is brown or black, it is due to lead sulphide which is formed by the interaction of hydrogen sulphide and lead compounds. Did the gas contain hydrogen sulphide? Did it contain carbon dioxide? Record the result of each observation. Sketch the essential parts of the apparatus.

### COMBUSTION OF ILLUMINATING GAS.

**Experiment 99.** — Attach a pointed glass tube (see Exp. 57 (*c*)) to the rubber tube connected with the gas jet and lower a small flame of illuminating gas into a cold, dry bottle. Observe the most definite result. Remove and extinguish the flame, add a little lime water, and shake. Describe the result. What do these two observations show about the composition of the main constituents of illuminating gas?

## CLASS-ROOM EXERCISE XXX.

1. Illuminating gas (coal gas).
    - (*a*) Industrial preparation, including a description of the apparatus for purification and storage.
    - (*b*) Constituents.
    - (*c*) Removal of impurities.
    - (*d*) By-products — use and value.
2. Water gas.
    - (*a*) Industrial preparation.
    - (*b*) Two main constituents.
    - (*c*) Why is it often "enriched"?
    - (*d*) Dangers from its use.
3. Miscellaneous.
    - (*a*) Upon what does the value of an illuminating gas depend?
    - (*b*) How is its value determined?
    - (*c*) Explain the statement, "This is a 20 candle-power gas."
    - (*d*) Describe the Welsbach light.
4. Review.
    - (*a*) Combustion.
    - (*b*) Distribution of carbon.
    - (*c*) Sources of ammonia gas and ammonium compounds.
    - (*d*) Source and use of (*a*) gas carbon, (*b*) coke.
    - (*e*) Law of Multiple Proportions in the light of the facts revealed by the composition of hydrocarbons.

## PROBLEMS. XXI.

1. How much illuminating gas at 10° C. and 530 mm. is required to fill a tank having a capacity of 800 cu. m.? (Specific gravity of illuminating gas is 0.5 referred to air, and 14.43 referred to hydrogen.)

2. A bottle contains 53.2 cc. of moist illuminating gas at 760 mm. and 18.5° C. What is the volume of the dry gas under standard conditions?

3. When steam is passed over red-hot coal in the preparation of water gas, the simplest equation for the reaction is —

$$C + H_2O = CO + H_2.$$

How many liters of hydrogen and of carbon monoxide at 10° C. and 750 mm. will be formed from 100 gm. of steam?

**Bunsen Burner.** — The Bunsen burner which is used so constantly in chemical operations was devised by the German chemist, Bunsen. The form varies, but the principle is the same in all burners. The parts of a common form are shown in Fig. 68.

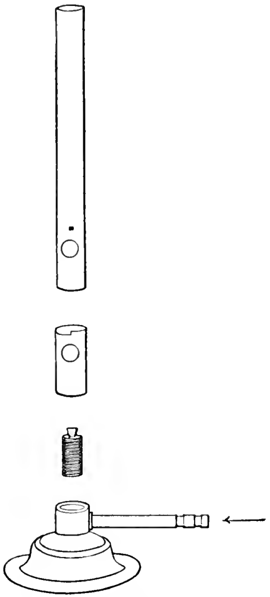

FIG. 68. — Parts of a Bunsen burner.

### CONSTRUCTION OF A BUNSEN BURNER.

**Experiment 100.** — Take apart a Bunsen burner and study the construction. Sketch the essential parts. Write a short description of the burner.

The theory and practical advantages of a Bunsen burner are only partly shown by its structure. It is also necessary to study the

### BUNSEN BURNER FLAME.

**Experiment 101.** — I. (*a*) Close the holes at the bottom of a Bunsen burner and hold a piece of crayon in the upper part of the flame. Note the black deposit. What is it? Where did it come from? Open the holes and hold the blackened crayon in the colorless flame. What becomes of the deposit? How is the flame changed, if at all? Does the experiment suggest a cause of the luminosity of a flame? What is it?

(*b*) Dip a glass tube a short distance into powdered wood charcoal, place the end containing the charcoal in one of the holes at the bottom of the burner, and blow gently two or three times into the other end. Describe and explain the result. Does it verify the answer to the last question in (*a*)?

(*c*) Open and close the holes of a lighted burner several times. Describe the result. Pinch the rubber tube to extinguish the flame, then light the gas at the holes. What change is produced in the flame? What causes the change?

Answer the following: —

(1) What is the object of the holes?
(2) Why does the gas burn at the top and not inside of the burner?
(3) Why does it sometimes "strike back" and burn inside?
(4) Why is the Bunsen flame non-luminous?

II (*a*) Hold a match across the top of the tube of a lighted Bunsen burner. When it begins to burn, remove and extinguish it. Note where it is charred, and explain the result. The same fact may be shown by sticking a pin through a (sulphur) match 15 mm. from the tip, suspending it across the burner, and then lighting the gas. The position of the match is shown in Fig. 69. Turn on a full current of gas before lighting it. What does this experiment show about the structure of the Bunsen flame? Verify your answer by (*b*).

(*b*) Bend a glass tube about 15 cm. long into the shape shown in Fig. 70. Hold the shorter arm in the flame about 2 cm. from the top of the burner tube. Hold a lighted match for an instant at the upper end of the tube. What does the result show about the structure of the Bunsen flame? Does it verify (*a*)?

FIG. 69. — A (sulphur) match suspended across the top of a Bunsen burner.

(*c*) Find the hottest part of the flame, when a full current of gas is burning, by holding a platinum or copper wire in the flame. Measure its distance, approximately, from the top of the burner tube.

(*d*) Examine a typical Bunsen flame — one which shows clearly the outlines of the inner part. What is the general shape of each main part? Draw a vertical and a cross section of the flame.

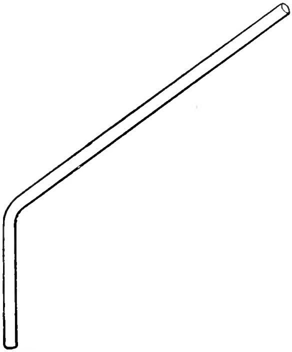

FIG. 70. — Bent tube for examining the structure of a Bunsen flame.

## CANDLE FLAME.

**Experiment 102.** — Attach a candle to a block of wood by means of a little melted candle wax, and proceed as follows: —

(*a*) Hold a cold, dry bottle over the lighted candle. Describe the result. What is the product? What is its source? Remove the bottle, pour a little lime water into it, and shake. Describe and explain the result. What are the two main products formed by a burning candle?

(*b*) Blow out the candle flame, and quickly hold a lighted match in the escaping smoke. Does the candle relight? Why? What is the general nature of this smoke? How is it related to the candle wax? How does (*b*) contribute to the explanation of (*a*)?

(*c*) Stand a lamp chimney over the lighted candle. How is the flame affected? Prop up the chimney on two blocks of wood. Does the candle continue to burn? Why? Now cover the top of the chimney with a piece of tin or similar covering. What is the result? Why?

### LABORATORY EXERCISE. XIII.

1. Sketch a candle flame.
2. Examine a lamp burner. Is it constructed according to the principles revealed in Exp. 102 (*c*)?
3. What is the essential difference between a Bunsen and a candle flame?
4. Is there any essential difference between a candle flame and a gas or lamp flame?
5. Explain the statement: "Large lamps have a powerful central draft."

### CLASS-ROOM EXERCISE. XXXI.

1. Essential facts in the life of Bunsen.
2. What is meant by the ignition point or kindling temperature of a flame?
3. Describe a miner's safety lamp, and state the exact principle on which its use depends.
4. Give additional facts about the Bunsen burner flame.
5. Give additional facts about the luminosity of flames.
6. Why does a draught of cool air often cause a lamp flame to smoke?

**Oxidizing and Reducing Flames.** — It is convenient to consider the Bunsen flame from two standpoints, viz. its power to give or to take oxygen, or in other words its power to oxidize or to reduce. That portion of the flame giving oxygen freely is called the **oxidizing flame**, and the part taking oxygen is called the **reducing flame**. A diagram of the general relation of these two flames is shown in Fig. 71 *A* is the most effective part of the oxidizing flame, and *B* of the reducing flame. At *A* metals are oxidized, and at *B* oxides are reduced.

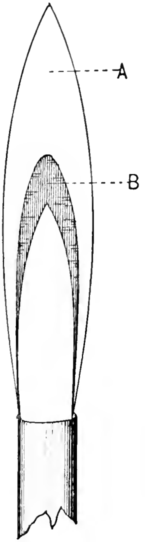

FIG. 71. — The oxidizing (*A*) and reducing (*B*) flames.

### LABORATORY EXERCISE. XIV.

1. Sketch the oxidizing and reducing flames.
2. Why do oxidation and reduction occur respectively at *A* and *B* (Fig. 71)?

**Borax Bead.** — A practical application of the oxidizing and reducing flames is often made by "testing with a borax bead." Borax, when heated, melts and forms a clear, transparent globule, which resembles a glass bead. Many metallic substances dissolve in melted borax and form colored beads. These beads differ in color and thus serve to identify the substance or at least its essential elementary constituent.

### TESTS WITH BORAX BEADS.

**Experiment 103.** — Make a small loop on the end of the platinum test wire (see App. A, § 14), moisten it, and dip it into powdered borax. Heat it in the flame, rotating it slowly; at first the borax swells, but finally shrinks to a small, transparent bead. If the bead is too small, add more borax and heat again. After use, the bead may be removed by dipping it, while hot, into water; the sudden cooling shatters the bead, which may then be easily rubbed or scraped from the wire.

(*a*) *Cobalt compounds.* Touch a transparent borax bead with a glass rod which has been moistened with cobalt nitrate or chloride solution. Heat the bead in the oxidizing flame. Notice the color when cold. If it is black, melt a little more borax into the bead; if faintly colored, moisten again with the cobalt solution. The color is readily detected by looking at the bead against a white object in a strong light, or by examining it with a lens. When the color has been definitely determined, heat again in the reducing flame. Compare the color of the cold bead with the previous observation.

(*b*) *Copper compounds.* Make another transparent bead, moisten it with copper sulphate solution, and heat it first in the oxidizing flame, and then in the reducing flame. Compare the colors of the cold beads, and draw a conclusion.

(*c*) *Manganese compounds.* Make another transparent bead, moisten it with manganese sulphate solution, or touch it with a minute quantity of manganese dioxide, and proceed as in (*b*). Compare the colors of the cold beads, and draw a conclusion.

(*d*) Tabulate the results of this experiment.

Differently colored hot and cold beads are characteristic of the compounds of different elements, and the bead test is often used to confirm other observations or to suggest a further special examination.

**The Blowpipe and its Use.** — Sometimes it is not convenient to apply the bead test, nor to hold a substance directly in the flame. It is then customary to study the action of a substance when heated on charcoal in a small flame, much like the Bunsen flame and produced by a blowpipe (Fig. 72). A blowpipe tube is put inside the burner tube (Fig. 73) to produce a luminous flame. The tip of the blowpipe rests on this tube (Fig. 74), and when air is gently blown through the blowpipe, a long, slender flame is produced (Fig. 75), exactly like a Bunsen

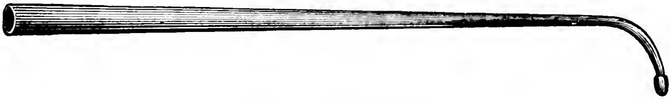

FIG. 72. — A blowpipe.

flame as far as its oxidizing and reducing powers are concerned. In Fig. 75, *A* is the oxidizing and *B* is the reducing part, *i.e.* these points are the most effective for these respective operations.

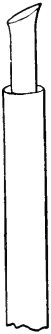

FIG. 73. — A blowpipe tube in position.

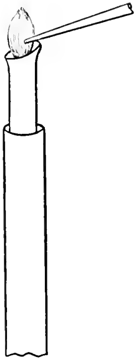

FIG. 74. — The blowpipe tip in position.

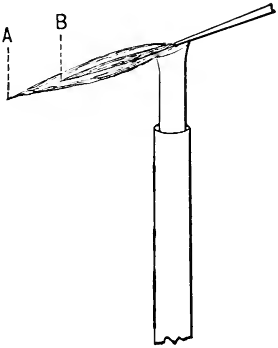

FIG. 75. — The blowpipe flame.

### USE OF THE BLOWPIPE.

**Experiment 104.** — Slip the blowpipe tube into the burner, light the gas and lower the flame until it is about 4 cm. high. Rest the tip of the blowpipe on the top of the tube, as shown in Fig. 74, placing the tip just within the flame. Put the other end of the blowpipe between the lips, puff out the cheeks, inhale through the nose, and exhale into the tube, using the cheeks somewhat as a bellows. Do not blow in puffs, but produce a continuous flow of air by steady and easy inhaling and exhaling. The operation is natural and simple, and, if properly performed, will not make one out of breath. The flame should be an inner blue cone surrounded by an outer and almost invisible cone, somewhat as shown in Fig. 75, though its shape varies with the method of production. Practice until the flame is produced voluntarily and without exhaustion. Watch the flame and learn to distinguish the two parts, so that they may be intelligently utilized.

Sketch a blowpipe and a blowpipe flame.

### REDUCTION WITH THE BLOWPIPE.

**Experiment 105.** — (*a*) *Lead compounds.* Make a shallow hole at one end of the flat side of a piece of charcoal. Fill the hole with a mixture of equal parts of powdered sodium carbonate and lead oxide, and heat the mixture in the reducing flame. The sodium carbonate melts and assists the fusion of the oxide, but the former is not changed chemically. In a short time bright, silvery globules will appear on the charcoal. Let the mass cool, and pick out the largest globules. Put one or two in a mortar, and strike with a pestle. Are they soft and malleable, or brittle and hard? State the result when a globule is drawn across or rubbed upon a white paper. How do the properties compare with those of metallic lead? What has become of the oxygen? Of what chemical use is the charcoal?

(*b*) Grind together in a mortar a little sodium sulphate and wood charcoal, adding at intervals just enough water to hold the mass together. Heat this paste for a few minutes in the reducing flame as in (*a*). Scrape the fused mass into a test tube, boil in a little water, and put a drop of the solution on a bright silver coin. If a dark brown stain is produced, it is evidence of the formation of silver sulphide. Repeat, if no such stain is produced. State all the chemical changes which led to the production of the silver sulphide, explaining at the same time how the experiment illustrates reduction.

### OXIDATION WITH THE BLOWPIPE.

**Experiment 106.** — (*a*) Heat a small piece of zinc on charcoal in the oxidizing flame. What is the product? Observe and record the color of the product, and the color of the coating on the charcoal when both hot and cold. (See (*d*).)

(*b*) Heat a piece of lead as in (*a*), and note and record the presence or absence of fumes, as well as the color of the coating when hot and cold.

(*c*) Heat a small piece of tin in the oxidizing flame, and observe and record the presence or absence of fumes, and the color of the coating, if any.

(*d*) Tabulate the above results, stating in each case (1) color of the coating on the hot charcoal, (2) color of the coating on the cold charcoal, (3) presence or absence of fumes, (4) name of product.

### CLASS-ROOM EXERCISE. XXXII.

1. What industrial use is made of the principle shown in Exp. 105 (*b*)?
2. Does the chemical action in Exp. 106 differ essentially from combustion in the air? Then why use a blowpipe?
3. Review.
   (*a*) Oxidation and reduction.
   (*b*) Compound blowpipe.

# CHAPTER XIV.

## SOME COMMON ORGANIC COMPOUNDS.

CARBON forms a vast number of compounds. Many of them were first obtained directly or indirectly from living things, and they naturally came to be considered under the comprehensive title of "Organic Chemistry." Such a subdivision is now made for convenience, not for any intrinsic chemical properties of these carbon compounds, or organic compounds, as they are often called.

Several organic compounds have already been studied; these will be reviewed from a broader standpoint, and a few others which are familiar will be examined.

The common organic compounds are members, or derivatives, of the classes:—

1. Hydrocarbons.
2. Alcohols.
3. Acids.
4. Carbohydrates.

**The Composition of Organic Compounds** has already been partially shown. They usually contain carbon and hydrogen, with or without oxygen, or nitrogen, or both; a few contain sulphur, phosphorus, and other elements. When burned or heated they yield carbon, or carbon dioxide, water vapor, ammonia gas, sulphur compounds, etc., according to their composition. The general character of typical organic compounds may be shown by a study of the

## COMPOSITION OF ORGANIC COMPOUNDS.

**Experiment 107.** — (*a*) *Carbon.* (1) Recall or repeat the experiments which showed that carbon is a constituent of wood, cotton, bone, starch, sugar, illuminating gas, candle wax, meat, flour, bread, albumen. (2) Heat 2 or 3 cc. of turpentine in a porcelain or iron dish, and then set fire to it. Does it contain carbon? Hold a bottle over the flame long enough to collect any product, and then test the contents for carbon dioxide; does the observation verify the previous conclusion? (3) Repeat with alcohol. Does it contain carbon? Burn a small lump of camphor in a dish or on a block of wood. Does it contain carbon? (4) Hold a bottle over a burning kerosene lamp long enough to collect any product, and test as in (2). Does kerosene contain carbon?

(*b*) *Hydrogen.* (1) Set fire to 1 or 2 cc. of the following liquids in a porcelain dish (or crucible), and hold over the flame a *cold* dry bottle long enough to allow the condensation of the water vapor which is always one product of the combustion of organic compounds which contain hydrogen: alcohol, turpentine, kerosene. (2) Heat in separate test tubes the following dry solids, and if they contain hydrogen, a little water vapor will condense on the upper part of the test tube: sugar, starch, flour, wood, paper, hair. (3) Hold a cold, dry bottle, for a few seconds over a burning kerosene lamp, a Bunsen flame, an ordinary gas flame, a burning candle, a burning taper. Is hydrogen a component of kerosene, illuminating gas, and wax?

The oxygen, which unites with the hydrogen to form the water, may come from the substance, as in the case of sugar, starch, wax, wood, paper, or it may come from the air. No simple experiment will determine the source of the oxygen.

(*c*) *Nitrogen.* Mix a little granulated gelatine (one part) with dry soda lime (two parts) and heat the mixture in a test tube. Hold a piece of moist red litmus paper in the escaping vapor. It will be turned blue by escaping ammonia gas. Gelatine (also horn, glue, and leather) contains nitrogen which is liberated in combination with hydrogen as ammonia gas.

(*d*) *Sulphur.* (1) Put a little mustard paste on a clean silver coin. The brown stain is silver sulphide. Explain. (2) Why is a silver spoon tarnished by a cooked egg?

Draw a general conclusion regarding the composition of organic compounds.

**Hydrocarbons.** — These compounds of hydrogen and carbon have already been considered in Chapter XIII.

### LABORATORY EXERCISE. XV.

1. Recall or repeat the experiments illustrating the preparation and properties of methane, ethylene, and acetylene.
2. What are the products of the combustion of these hydrocarbons?
3. Write the equations for the combustion of methane and ethylene.

### CLASS-ROOM EXERCISE. XXXIII.

1. Review Class-Room Exercise XXIX., Chapter XIII., page 223.
2. Give the components (if a compound) or the main constituents (if a mixture) of kerosene, petroleum, natural gas, illuminating gas, naphtha, gasoline, turpentine, benzine, benzene.

### PROBLEMS. XXII.

1. Review Problems XX., Chapter XIII., page 224.

**Alcohols.** — These are compounds of carbon, hydrogen, and oxygen. Ethyl alcohol is the best known member of this class, and is therefore called *alcohol*, just as sodium chloride is called *salt*. Alcohol is formed by the fermentation of glucose, or grape sugar.

### PREPARATION OF ALCOHOL.

**Experiment 108.** — Prepare alcohol from one of the following mixtures: —

(*a*) 300 gm. of grape sugar.
2 l. of water.
500 cc. of yeast.

(*b*) 50 gm. of grape sugar.
1 l. of water.
60 cc. of yeast.

(*c*) 25 gm. of grape sugar.
150 cc. of water.
$\frac{1}{4}$ of a compressed yeast cake.

I. Dissolve the grape sugar in the water and add the yeast. Put the mixture in a large bottle or flask provided with a one-hole rubber stopper fitted with a delivery tube which reaches to the bottom of a small bottle half full of lime water.

Pour enough kerosene down a rod upon the lime water to prevent it from interacting with the carbon dioxide of the air. Stand the apparatus in a dark, or moderately dark place, where the temperature is 25°–30° C.

Fermentation begins at once, and carbon dioxide, one of the products, bubbles through the lime water. Examine the stopper for a leak, if no change occurs in the lime water. The operation should be allowed to continue at least a day, and longer if possible. The flask will then contain mainly water, unchanged grape sugar, alcohol, and some products of minor importance. Decant the liquid, agitate it with a little bone black to remove the odor and color, and filter. The alcohol, which varies in quantity with the conditions, is dissolved in a large excess of water and must be separated by distillation.

II. The distillation is performed with the apparatus used in Exp. 30. Fill the flask half full of the liquid from I., add a few pieces of pipestem (or granulated zinc, or glass tubing) to prevent "bumping," and distil about 50 cc. Save the distillate. Replace the residue in the flask by more liquid from I., distil again, and repeat this operation until all the liquid has been used. Replace the one-hole stopper with a two-hole stopper, insert a thermometer in one hole so that the bulb just touches the surface of the combined distillates which should now be distilled. Heat gently, and collect in a separate receiver the distillate which is formed when the liquid boils between 80° and 93° C. This distillate contains most of the alcohol.

Note the odor. Drop a little into a warm dish, and hold a lighted match over it. If it does not burn, it shows that the alcohol is too dilute. Put a little in a dish, warm gently, and light the vapor. Describe the result.

If this experiment is indifferently successful, the general character of alcohol may be learned by a study of

## THE PROPERTIES OF ALCOHOL.

**Experiment 109.** — (*a*) Determine cautiously the odor and taste of alcohol. Drop a little on a glass plate or on a piece of paper, and watch it evaporate. Is its rate of evaporation more rapid than that of water?

(*b*) Weigh a measured quantity (about 25 cc.) of 95 per cent alcohol and calculate its specific gravity.

(*c*) Alcohol dissolves many organic substances. Try camphor, powdered shellac, or rosin.

(*d*) Burn a little alcohol in a dish and observe the nature of the flame. What are the products of combustion?

(*e*) Cautiously add a few drops of concentrated sulphuric acid to equal volumes (about 5 cc. each) of acetic acid and alcohol. Shake carefully, and then warm gently. The pleasant fruit-like odor is due to the vapor of ethyl acetate, a volatile liquid which is always formed under these circumstances. Its formation is a simple test for either alcohol or acetic acid. Its nature will be subsequently considered.

### LABORATORY EXERCISE. XVI.

1. Summarize the properties of alcohol.
2. Complete the equation —

$$C_2H_6O \quad + \qquad = \quad CO_2 \quad + \quad H_2O.$$

**Fermentation** is a destructive process caused, probably, by various organisms, and the products vary with the organism. The equation for the reaction when sugar ferments is

$$\underset{\text{Glucose}}{C_6H_{12}O_6} \quad = \quad \underset{\text{Alcohol}}{2\,C_2H_6O} \quad + \quad 2\,CO_2.$$

Cane sugar (ordinary sugar) does not ferment. If boiled with acid, it changes thus: —

$$\underset{\text{Cane Sugar}}{C_{12}H_{22}O_{11}} \quad + \quad H_2O \quad = \quad \underset{\text{Glucose}}{C_6H_{12}O_6} \quad + \quad \underset{\text{Fructose}}{C_6H_{12}O_6},$$

and fermentation may then occur as usual.

**The Formula of Alcohol** is often written $C_2H_5.OH$, because conclusive evidence shows that this general grouping of atoms exists in the molecule. The radical $C_2H_5$ is called *ethyl*. Like ammonium, it exists only in combination. The ethyl group is present in many organic compounds, and its presence in alcohol gives rise to the term *ethyl alcohol.*

Alcohols are analogous to metallic hydroxides. Thus $C_2H_5.OH$ corresponds to KOH and NaOH. Both form salts. Hence we may write, as in the case of the experiment on neutralization : —

$$C_2H_5.OH + HCl = \underset{\text{Ethyl Chloride}}{C_2H_5.Cl} + H_2O.$$

$$C_2H_5.OH + \underset{\text{Acetic Acid}}{C_2H_4O_2} = \underset{\text{Ethyl Acetate}}{C_2H_5.C_2H_3O_2} + H_2O.$$

Ethyl acetate is analogous to sodium acetate. It is an organic salt, and it is often called an *ethereal salt,* or *ester.*

### CLASS-ROOM EXERCISE. XXXIV.

1. Additional study of alcohol.
    - (*a*) Occurrence.
    - (*b*) Various methods of preparation.
    - (*c*) Properties.
    - (*d*) Purification.
    - (*e*) Test.
    - (*f*) Constitution.
    - (*g*) Uses.
    - (*h*) Determination of alcohol in liquors.
2. Fermentation.
    - (*a*) General nature.
    - (*b*) Alcoholic fermentation.
    - (*c*) Applications.
    - (*d*) Nature of ferments.
3. Composition of brandy, whiskey, wines, and beer.
4. Methyl alcohol.
    - (*a*) Preparation.
    - (*b*) Properties.
    - (*c*) Uses.
    - (*d*) What is methylated spirit ?
5. Miscellaneous.
    - (*a*) What is "proof spirit" ?
    - (*b*) History of alcohol.
    - (*c*) Derivation of the word *alcohol.*
    - (*d*) Absolute alcohol.
    - (*e*) Tinctures.

6. Ethereal salts.
    - (*a*) Occurrence in flowers and fruits.
    - (*b*) General properties.
    - (*c*) Artificial preparation. (*d*) Uses.

**Ether** is a compound of carbon, hydrogen, and oxygen, and is made from alcohol by heating with sulphuric acid. In its narrow sense the term ether means the best known member of an homologous series of organic compounds which is analogous to the metallic oxides. Ordinary ether is ethyl oxide, $(C_2H_5)_2O$ or $C_2H_5 . O . C_2H_5$.

### PROPERTIES OF ETHER.

**Ether vapor is easily ignited, and should never be brought near a flame, unless special directions are so given.**

**Experiment 110.** — (*a*) Pour a little ether into a dish or test tube and observe the odor and volatility. Taste cautiously. Pour a drop upon a glass plate or a block of wood. How does its rate of evaporation compare with that of alcohol?

(*b*) Recall or repeat the experiment (see Exp. 33 (*b*)) illustrating the solubility relations of ether and water.

(*c*) Add a bit of wax to a few cubic centimeters of ether. The result is typical; draw a conclusion.

### LABORATORY EXERCISE. XVII.

1. From what has already been shown about ether, predict its approximate boiling point.

2. Explain the statement, "ether is miscible with alcohol in all proportions."

3. Complete the equation for the combustion of ether: —

$$(C_2H_5)_2O + 6\,O_2 = CO_2 + H_2O.$$

### CLASS-ROOM EXERCISE. XXXV.

1. Additional study of ether.
    - (*a*) Industrial preparation. (*b*) Properties. (*c*) Uses.
    - (*d*) Formula, constitution, and equations for its preparation.

2. Miscellaneous.
   - (*a*) Ethers.
   - (*b*) Significance of the terms *ethyl ether* and *sulphuric ether*.
   - (*c*) Explain the expression, "ether is prepared by a continuous process."
   - (*d*) Anhydrous ether.

## PROBLEMS. XXIII.

1. Ether boils at 34.9° C. and alcohol at 78.4° C. What are the corresponding points on the Fahrenheit scale?

2. Calculate the weight of alcohol which may be obtained from a kilogram of cane sugar.

3. Calculate the percentage composition of—
   - (*a*) Alcohol, $C_2H_6O$.
   - (*b*) Acetic acid, $C_2H_4O_2$.
   - (*c*) Cane sugar, $C_{12}H_{22}O_{11}$.

4. If 10 gm. of pure alcohol are burned, what weight of each product is formed?

5. One and a half kilograms of absolute alcohol are burned.
   - (*a*) How many liters of oxygen are needed?
   - (*b*) How many liters (at standard conditions) of carbon dioxide and of water vapor are formed?

6. Calculate the simplest formula of the substances possessing the indicated composition: —

| | | |
|---|---|---|
| (*a*) | Carbon | 40.00 |
| | Hydrogen | 6.67 |
| | Oxygen | 53.33 |
| (*b*) | Carbon | 15.80 |
| | Hydrogen | 5.26 |
| | Nitrogen | 36.84 |
| | Sulphur | 42.10 |
| (*c*) | Carbon | 54.55 |
| | Hydrogen | 9.09 |
| | Oxygen | 36.36 |
| (*d*) | Carbon | 91.30 |
| | Hydrogen | 8.69 |
| (*e*) | Carbon | 49.05 |
| | Hydrogen | 5.14 |
| | Oxygen | 17.20 |
| | Nitrogen | 28.61 |
| (*f*) | Carbon | 74.07 |
| | Hydrogen | 8.64 |
| | Nitrogen | 17.29 |
| (*g*) | Carbon | 35.71 |
| | Hydrogen | 2.38 |
| | Nitrogen | 33.33 |
| | Oxygen | 28.58 |

**Aldehyde** is a compound of carbon, hydrogen, and oxygen, which is formed by the oxidation of alcohol. Aldehyde, like alcohol and ether, is the name both of an homologous series and of its best known member.

### PREPARATION AND PROPERTIES OF ALDEHYDES.

**Experiment 111.** — (*a*) *Acetic Aldehyde.* Add a little concentrated hydrochloric acid and a few drops of alcohol to a few cubic centimeters of potassium dichromate solution. Warm gently, and observe the peculiar-smelling gaseous product. It is aldehyde vapor, aldehyde itself being a colorless, extremely volatile liquid which boils at 20.8° C. See Exp. 183.

(*b*) *Formic Aldehyde* or *Formaldehyde.* Put a few cubic centimeters of methyl alcohol in a test tube and stand the test tube in a rack. Wind a piece of copper wire into a spiral around a glass rod or lead pencil. Slip the spiral from the rod, grasp one end in the forceps, and heat the wire red-hot in the flame. Then quickly drop it into the methyl alcohol. The pungent vapor which is suddenly produced is the vapor of formaldehyde.

The preparation of formaldehyde is represented by the equation —

$$\underset{\text{Methyl Alcohol}}{CH_3 . OH} + O = \underset{\text{Formaldehyde}}{H . CHO} + H_2O$$

The formula of acetic aldehyde is $CH_3 . CHO$.

### LABORATORY EXERCISE. XVIII.

1. Complete the equation for the formation of acetic aldehyde:

$$C_2H_5 . OH + O =$$

2. Derivation of the term *aldehyde.*
3. Why may acetic aldehyde be called ethaldehyde?

### CLASS-ROOM EXERCISE. XXXVI.

1. Formaldehyde.
   (*a*) Preparation. (*b*) Properties.
   (*c*) Uses as disinfectant and germicide.
2. Use of aldehydes in formation of mirrors.

**Acetic Acid** is an exceedingly important organic compound. The acid and its derivatives find numerous applications.

### PROPERTIES OF ACETIC ACID.

**Experiment 112.** — Recall or determine again the general properties of acetic acid.

Warm a little in a test tube and compare the odor with that of other acids.

### TEST FOR ACETIC ACID.

**Experiment 113.** — Repeat Exp. 109 (*e*), and describe the result. What is the name of the sodium compound which corresponds to the main product of the reaction?

**Vinegar** is dilute, impure acetic acid.

### PROPERTIES OF VINEGAR.

**Experiment 114.** — Show experimentally that vinegar is (1) an acid, and (2) contains acetic acid.

**Acetates** are salts of acetic acid.

### PREPARATION OF ACETATES.

*Most acetates are poisonous.*

**Experiment 115.** — Prepare one or all of the following acetates: —

(*a*) *Sodium acetate.* Dissolve 20 gm. of sodium carbonate in 10 cc. of water in a large porcelain or agate dish, and slowly add 30 cc. of commercial acetic acid, with constant stirring. If the solution is not acid, add a little more acetic acid. Filter the solution, if not clear. Evaporate to crystallization. When the crystals have formed, remove and dry them. Describe the crystals. Prove that they contain water of crystallization. Test the acetate as follows: (1) Dissolve a little in water, add a few drops of concentrated sulphuric acid, and boil. What does the odor show is present? What other acids have been similarly prepared? (2) Dissolve as in (1), add a few drops of alcohol and of sulphuric acid, and boil. What does the odor conclusively prove? Preserve the crystals, finally, in a glass-stoppered bottle, or in one having a cork covered with paraffine.

(*b*) *Lead acetate* (*poisonous*). To 10 gm. of litharge add 18 cc. of commercial acetic acid in small portions. Stir the mixture constantly during the addition of acid. After all the acid has been added, heat gently until the action ceases. If the solution is green or bluish, it is due to a copper compound. The copper may be precipitated and removed mechanically by standing a strip of lead in the solution for an hour or more. Decant the clearer liquid and then filter. Evaporate to crystallization on a water bath. Remove the crystals from the liquid, and dry at a moderate temperature. Preserve the crystals finally as in (*a*). Describe the crystals. Test them for lead (see Exp. 98, last paragraph, and Exp. 134 (*c*) (1)), and for an acetate.

**The Formula of Acetic Acid** is $C_2H_4O_2$ or $CH_3.COOH$. This acid is formed by the oxidation of ethyl alcohol, thus:—

$$C_2H_6O + O_2 = C_2H_4O_2 + H_2O.$$

The change is due to fermentation. Acetic acid is monobasic.

CLASS-ROOM EXERCISE. XXXVII.

1. Additional study of acetic acid.
   - (*a*) Industrial preparation.
   - (*b*) Glacial acetic acid.
   - (*c*) Properties.
   - (*d*) Uses.
   - (*e*) Test.
   - (*f*) Constitution.
2. Acetates.
   - (*a*) General method of preparation.
   - (*b*) Chemical name and formula of iron liquor, red liquor, verdigris, sugar of lead.
   - (*c*) Formula of sodium, lead, silver, chromium, ammonium, aluminium, and calcium acetates.
   - (*d*) Uses of acetates.
3. Vinegar.
   - (*a*) What is it?
   - (*b*) Chemistry of its formation.
   - (*c*) Industrial preparation
   - (*d*) Uses.
4. Oxalic acid.
   - (*a*) Preparation.
   - (*b*) Properties.
   - (*c*) Formula and constitution.
   - (*d*) What experimental use has been made of oxalic acid? What do the results show about its composition?

5. Show by formulas the relation between ethyl alcohol, acetic aldehyde, and acetic acid, and by equations the possibility of passing from one to the other.

**Soap** is a mixture of alkaline salts of organic acids, mainly palmitic and stearic acids. It is made by boiling caustic soda or potash with fats. The process is called saponification, and consists simply in decomposing an organic salt (fat) into the corresponding alcohol (glycerine) and an alkaline salt (soap). Fats are analogous to ethyl acetate, ordinary fats being mixtures of glyceryl palmitate, stearate, and oleate.

### PREPARATION OF SOAP.

**Experiment 116.** — Prepare soap in an iron or a tin dish by one of the following methods: —

(*a*) Dissolve 10 gm. of sodium hydroxide in 75 cc. of water, add 30 gm. of lard, and boil until the mixture begins to solidify. Then add 20 gm. of fine salt in small portions. Stir constantly during the addition of the salt. Boil a few minutes. Let the mass cool, and then remove the soap, which will form in a cake at the surface.

(*b*) Dissolve 13 to 15 gm. of sodium hydroxide in 100 cc. of water, add 100 cc. of castor oil, and boil for about half an hour. Add 20 gm. of salt, and then proceed as in (*a*).

(*c*) Dissolve 8 gm. of potassium hydroxide in 150 cc. of alcohol, add 10 gm. of lard, and stir constantly while the mixture is being heated on a water bath to syrupy consistency. Allow the contents to cool. The jelly-like product is soap.

Preserve a sample.

### PROPERTIES OF SOAP.

**Experiment 117.** — (*a*) Leave soap shavings exposed to the air for several days. What does the result show about the presence of water in the soap?

(*b*) Test a yellow soap for free alkali. Is such a soap carefully prepared?

(*c*) Add considerable dilute sulphuric acid to a soap solution. The greasy precipitate, which finally rises to the top, is a mixture mainly of palmitic and stearic acids.

(*d*) To a little soap solution in separate test tubes add calcium sulphate and magnesium sulphate solutions. Describe the result. Boil for a few minutes and describe the result. Prepare a solution of acid calcium carbonate by passing carbon dioxide into lime water until the precipitate is redissolved. Add some of the solution to a soap solution, and describe the result. Boil, as above, and describe the result.

**Hardness of Water** is illustrated by Exp. 117 (*d*). Magnesium and calcium sulphates and calcium carbonate (in solution as the acid carbonate) form insoluble salts with soap. As long as water contains these calcium and magnesium compounds, the soap is useless as a cleansing agent, since it is immediately transformed into a sticky solid. If the hardness cannot be removed by boiling, as in the case of water containing magnesium and calcium sulphates, the hardness is called *permanent*. Heat decomposes acid calcium carbonate, the calcium carbonate is precipitated and carbon dioxide is liberated; hence hardness due to calcium carbonate can be removed by boiling, and is therefore called *temporary*. *Soft water* contains no magnesium or calcium salts.

### CLASS-ROOM EXERCISE. XXXVIII.

1. Additional study of soap.
    - (*a*) Industrial preparation.
    - (*b*) Chemistry of its preparation.
    - (*c*) Kinds.
    - (*d*) Uses.
2. Glycerine (or Glycerol).
    - (*a*) Industrial preparation.
    - (*b*) Properties.
    - (*c*) Relation to soap industry.
    - (*d*) Formula and composition.
    - (*e*) Uses.
3. Miscellaneous.
    - (*a*) Stearin. — preparation and uses.
    - (*b*) What is saponification?
    - (*c*) Organic acids related to stearic acid.

**Carbohydrates** are compounds of carbon, hydrogen, and oxygen. Sugar and starch are familiar representatives.

### LABORATORY EXERCISE. XIX.

1. Examine sugar and summarize its general properties.

2. What happens when it is heated? When treated with concentrated sulphuric acid? What do these experiments show about its composition?

3. Examine sugar through a lens. What is its crystal form?

### FEHLING'S TEST FOR SUGAR.

**Experiment 118.** — Perform Exp. 153.

### LABORATORY EXERCISE. XX.

1. Examine starch through a microscope, and describe the result. How does it differ from sugar?

2. What is the effect of heat on starch? Of concentrated sulphuric acid? What does it contain, according to these experiments?

3. Summarize the general properties of starch.

4. Is starch soluble in cold water?

### CLASS-ROOM EXERCISE. XXXIX.

1. Sugars.
    - (*a*) Distribution of sugars.
    - (*b*) Industrial preparation and properties of glucose.
    - (*c*) Cane sugar, — distribution, source, industrial preparation, properties.
    - (*d*) Review "Fermentation."
    - (*e*) Formula and composition of several sugars.
2. Miscellaneous.
    - (*a*) What is molasses? Sugar of milk? Fructose? Glucose?
    - (*b*) Beet root sugar industry.
3. Starch.
    - (*a*) Distribution and source.
    - (*b*) Industrial preparation.
    - (*c*) Properties.
    - (*d*) Uses.
    - (*e*) Formula.
    - (*f*) Kinds.
4. Chemistry of bread-making.
5. Dextrine — preparation, properties, uses, and relation to starch and dextrose.
6. Cellulose — distribution, properties, and uses.

# CHAPTER XV.

## BROMINE, IODINE, AND HYDROFLUORIC ACID.

THE elements hitherto studied suggest little or no relation to each other. Chlorine, however, bears a strong resemblance to bromine and iodine, and the experiments in this chapter are intended to reveal, among other things, this relation as well as to present the special facts about bromine and iodine and their compounds.

### PREPARATION AND PROPERTIES OF BROMINE.

**Precaution.**—*Bromine is a corrosive liquid which forms, at the ordinary temperature, a suffocating vapor. Perform in the hood all experiments which use or evolve bromine.*

**Experiment 119**.—Bend a glass tube about 30 cm. long into the shape shown in Fig. 76. The lower end passes through the hole of a rubber stopper which should tightly fit a large test tube. Put five or six crystals of potassium bromide in the test tube, add an equal quantity of manganese dioxide and 10 cc. of dilute sulphuric acid. Insert the stopper and its tube securely, and boil gently. Do not hold the test tube in the hand, but use the test tube holder. Brown fumes soon appear in the test tube and pass out of the delivery tube. Regulate the heating so that this vapor will condense and collect in the lower bend of the delivery tube. Both vapor and liquid are bromine. When no further boiling produces bromine vapor in the test tube, pour the bromine into a bottle of water. Observe and record the physical properties of this bromine, especially the color, solubility in water, specific gravity, volatility, and physical state. Try the action of the contents of

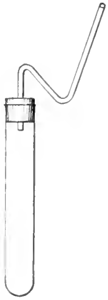

FIG. 76.—Apparatus for preparing bromine.

the bottle on litmus paper; if the action is not marked, push the paper down near the bromine. Determine the odor by smelling cautiously of the water in the bottle. As soon as these observations have been made, pour the contents of the bottle into the sink and flush with water, or pour into a jar in the hood. Wash the test tube free from all traces of bromine, taking care to get none on the hands.

Answer the following: —

1. In what ways does bromine physically resemble chlorine? In what ways does it differ from chlorine?

2. How is it essentially different from all other elements previously studied?

The simplest equation for the reaction in the preparation of bromine is —

$$\underset{\substack{\text{Potassium}\\ \text{Bromide}}}{2\,KBr} + 2\,H_2SO_4 + MnO_2 = \underset{\text{Bromine}}{Br_2} + \underset{\substack{\text{Manganese}\\ \text{Sulphate}}}{MnSO_4} + \underset{\substack{\text{Potassium}\\ \text{Sulphate}}}{K_2SO_4} + 2\,H_2O \tag{25}$$

$$238 + 196 + 87 = 160 + 151 + 174 + 36$$

PREPARATION AND PROPERTIES OF HYDROBROMIC ACID.

*Perform this experiment in the hood.*

**Experiment 120.** — Put three or four crystals of potassium bromide in a test tube, and add a few drops of concentrated sulphuric acid. The white product is hydrobromic acid. Observe its most apparent properties — cautiously. Test the gas with both kinds of litmus paper. Does this product resemble hydrochloric acid? Are they similar enough to be closely related? Give any evidence of a secondary reaction.

Since $2\,NaCl + H_2SO_4 = 2\,HCl + Na_2SO_4$ is the equation for the preparation of hydrochloric acid, what is the equation for the preparation of hydrobromic acid?

Hydrobromic acid forms **bromides**, just as hydrochloric acid forms chlorides.

## PROPERTIES OF POTASSIUM BROMIDE.

**Experiment 121.** — Examine a crystal of potassium bromide. Compare it with potassium or sodium chloride in regard to crystal form, color, and solubility in water.

Add a few drops of silver nitrate solution to potassium bromide solution. Is the solid product soluble in dilute ammonium hydroxide? Compare the interaction with the interaction of silver nitrate and chlorides. Do the properties of bromides, typified by potassium bromide, suggest any marked relation to chlorides?

## PREPARATION AND PROPERTIES OF IODINE.

**Experiment 122.** — Mix and grind together in a mortar three or four crystals of potassium iodide and an equal quantity of manganese dioxide. Put the mixture in a 250 cc. Erlenmeyer flask, add about 8 cc. of water, and about 5 cc. of concentrated sulphuric acid, drop by drop. Invert a dry funnel upon the flask, as shown in Fig. 77, and stand the flask on a sand bath. Heat with a low, steady flame. Too much heat will vaporize the water, or crack the flask — perhaps both. The vapor of iodine will fill the flask, and crystals of iodine will collect on the sides of the funnel. Regulate the heat so that the flask is filled with vapor, but so that little or none escapes from the funnel. If vapor escapes from the funnel, plug the stem with cotton or soft paper. If crystals of iodine collect on the upper portions of the flask, a gentle heat will remove them. Continue to heat until no more iodine is evolved, or until sufficient iodine has been collected for the subsequent experiments. Scrape the crystals into a dish. Observe and record the physical properties of iodine, especially the color of the solid and of the vapor, volatility, and odor (cautiously). Heat a crystal in a dry test tube, and when the tube is half full of vapor invert it. What does the result show about the specific gravity of iodine vapor? Touch a crystal with the finger. What color is the stain? Will water remove it? Will alcohol? Will a solution of potassium iodide? What do these results show about the solubility of iodine? Preserve the crystals.

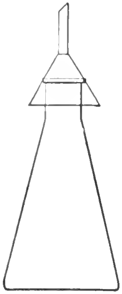

FIG. 77. — Apparatus for preparing iodine.

Answer the following: —

1. In what ways does iodine resemble chlorine and bromine? In what ways does it differ from them?

2. How does it differ essentially from all other elements previously studied?

The simplest equation for the chemical action in the preparation of iodine is —

$$\begin{array}{ccccccccccccc} 2\,KI & + & 2\,H_2SO_4 & + & MnO_2 & = & I_2 & + & MnSO_4 & + & K_2SO_4 & + & 2\,H_2O \\ \text{Potassium Iodide} & & & & & & \text{Iodine} & & & & & & \\ 332 & + & 196 & + & 87 & = & 254 & + & 151 & + & 174 & + & 36 \end{array} \qquad (26)$$

### THE CARBON BISULPHIDE TEST FOR IODINE.

**Precaution.** — *Carbon bisulphide is inflammable. It should not be used near flames.*

**Experiment 123.** — (*a*) *Free iodine.* Add a few drops of carbon bisulphide to a very dilute solution of iodine, made by dissolving a crystal of iodine in a solution of potassium iodide or in alcohol, and observe the color of the carbon bisulphide, which, being much heavier than water, will sink to the bottom of the test tube. How does it resemble the color of iodine vapor?

(*b*) *Combined iodine.* Add a few drops of carbon bisulphide to a very dilute solution of potassium iodide. Is there positive evidence of iodine? Now add a few drops of chlorine water (see App. C, REAGENTS), and shake. How does this result compare with the final result in (*a*)? The result is due to the fact that chlorine liberates iodine from its compounds, and the iodine, being free, forms the characteristic color.

### THE STARCH TEST FOR IODINE.

**Experiment 124.** — Grind a gram of starch in a mortar with a little water to creamy consistency. Pour this into about 100 cc. of boiling water, and stir the hot liquid. Allow it to cool, or cool it by holding the vessel in a stream of cold water, and then decant the clear liquid. Use this starch solution to test for iodine.

(*a*) *Free iodine.* Add a few cubic centimeters of the starch solution to a test tube nearly full of water, and then add a few drops of a solution of iodine. The deep blue color is due to the presence of a compound which is always formed under these circumstances, but the composition of which is unknown. If the color is black, pour out half of the liquid and add more water, or pour some of the liquid into a vessel of water.

(*b*) *Combined iodine.* Add a few cubic centimeters of the starch solution to a very dilute solution of potassium iodide. Is the blue compound formed? Add a few drops of chlorine water and shake. Compare with the final result in Exp. 123 (*b*).

The facts revealed by Exp. 124 are practically utilized in —

### THE DETECTION OF STARCH.

**Experiment 125.** — Test the following for starch by grinding a small portion of each with water in a mortar, and then adding a *few drops* of the extract to an exceedingly dilute solution of iodine: (*a*) wheat flour, (*b*) oatmeal, (*c*) potato, (*d*) rice, (*e*) beans, (*f*) bread, (*g*) cracker, (*h*) pepper, (*i*) baking powder.

Use, if possible, about the same amount in each case, and the results will then be an approximation of the relative amount of starch. Tabulate the results. If the starch is detected in (*h*) and (*i*), give a reason for its presence.

### PREPARATION AND PROPERTIES OF HYDRIODIC ACID.

**Experiment 126.** — Put three or four crystals of potassium iodide in a test tube and add a few drops of concentrated sulphuric acid. Fumes of hydriodic acid are formed among other products. Observe the most obvious properties of these fumes as in Exp. 120, and compare with the fumes of hydrochloric and hydrobromic acids. Are the acids similar?

Record the evidence of secondary reactions in this experiment. If nothing but hydriodic acid were formed, what would be the simplest equation? Why would this method be impracticable for the preparation of hydriodic acid on a large scale?

Hydriodic acid forms **iodides,** which correspond to chlorides and bromides.

### PROPERTIES OF POTASSIUM IODIDE.

**Experiment 127.** — Examine a crystal of potassium iodide. Compare it with potassium chloride and potassium bromide in regard to crystal form, color, and solubility in water.

Add a few drops of silver nitrate solution to potassium iodide solution. Compare the precipitate with the one obtained in the case of chlorides and of bromides. Test its solubility in dilute ammonium hydroxide. Does potassium iodide exhibit any marked resemblance to potassium chloride and potassium bromide?

### CLASS-ROOM EXERCISE. XL.

1. Additional study of bromine.
    (*a*) Occurrence.
    (*b*) Sources of commercial bromine.
    (*c*) Industrial preparation.
    (*d*) Summary of properties experimentally observed.
    (*e*) Additional properties.
    (*f*) Preparation and properties of bromine water.
    (*g*) Uses.
2. Additional study of iodine.
    (*a*) Occurrence.
    (*b*) Industrial preparation from seaweed and from Chili saltpeter.
    (*c*) Summary of properties experimentally observed.
    (*d*) Additional properties. (*e*) Vapor density.
    (*f*) Solubility and color of various solutions. (*g*) Uses.
3. Historical.
    (*a*) Discovery of bromine by Balard. Liebig's blunder.
    (*b*) Contributions of Courtois, Davy, and Gay-Lussac to the discovery of iodine.

4. Prepare a table showing the relation between chlorine, bromine, and iodine. The table should include the atomic weight, physical state, color, typical method of preparation, properties, and analogous compounds.

5. Tabulate the properties of hydrochloric, hydrobromic, and hydriodic acids.

6. What chemical properties are possessed in common by chlorine, bromine, and iodine?

7. Miscellaneous.
   (*a*) What is "drug store iodine"?
   (*b*) How is Exp. 124 utilized in chemical analysis?
   (*c*) What is a "bromide" print?
   (*d*) What is iodoform, and for what is it used?
   (*e*) Meaning and significance of the terms "halogen," "haloid," and "halide."

8. Prepare a short essay on "The Stassfurt Deposits."

### PROBLEMS. XXIV.

1. Bromine boils at 138.7° F. What is its boiling point on the centigrade scale?

2. The specific gravity of bromine is 3. What volume does a pound occupy?

3. Calculate the percentage composition of—
   (*a*) Potassium bromide.
   (*b*) Potassium iodide.
   (*c*) Silver bromide.
   (*d*) Hydrobromic acid.
   (*e*) Hydriodic acid.
   (*f*) Iodoform, $CHI_3$.

4. Marignac found that 3.946 gm. of silver dissolved in pure nitric acid were precipitated by 4.353 gm. of potassium bromide. What is the atomic weight of bromine? (Assume $K = 39$.)

5. Iodine vapor is 8.716 times heavier than air. What is its molecular weight?

6. How much potassium iodide is necessary to prepare 63.5 gm. of iodine?

7. How much iodine is liberated from potassium iodide by all the chlorine obtained by boiling 6 gm. of pure manganese dioxide with hydrochloric acid? (Equation is $2\ KI + MnO_2 + 4\ HCl = I_2 + MnCl_2 + 2\ KCl + 2\ H_2O$.)

8. How much potassium bromide is necessary to prepare 10 gm. of bromine?

**The Element Fluorine** is not adapted to examination in the laboratory on account of its corrosive properties. Its acid, however, may be studied, if it is prepared with caution.

### PREPARATION AND PROPERTIES OF HYDROFLUORIC ACID.

**Precaution.** — *Hydrofluoric acid gas is a corrosive poison. An aqueous solution of the gas — commercial hydrofluoric acid — burns the flesh frightfully.*

**Experiment 128.** — Warm a glass plate about 10 cm. square by dipping it into hot water or by standing it near a warm object, such as a radiator or drying oven. If it is held over a flame it is liable to crack. Coat one surface with beeswax or paraffine. The surface should be uniformly covered with a thin layer. Scratch letters, figures, or a diagram through the wax with a pointed instrument. Be sure the instrument removes the wax through to the glass, and that the lines are not too fine.

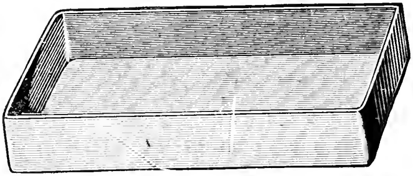

FIG. 78. — A lead dish.

Put 5 gm. of calcium fluoride in a lead dish (Fig. 78), and add just enough concentrated sulphuric acid to form a thin paste. Stir the mixture with a file or nail. Place the glass plate, wax side down, upon the lead dish and stand the whole apparatus in the hood for several hours, or until some convenient time. Remove the plate. Scrape the contents of the dish, immediately, into a waste jar in the hood, and wash the dish free from acid. Most of the wax can be scraped from the glass plate with a knife. The last portions can be removed by rubbing with a cloth moistened with alcohol or turpentine. If time permits, the wax can be removed by supporting the plate on nails in a dish of water and boiling the water for a few minutes. Do not attempt to melt off the wax over the flame. If the experiment has been properly performed, the plate will be etched where the glass was exposed to the hydrofluoric acid gas. Figure 79 shows an etched plate.

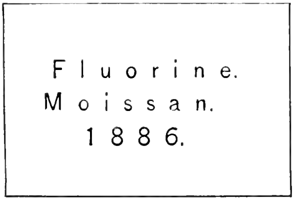

FIG. 79. — Glass plate etched by hydrofluoric acid gas.

The simplest equations for the two reactions in Exp. 128 are —

$$\begin{array}{ccccccc} CaF_2 & + & H_2SO_4 & = & 2\,HF & + & CaSO_4 \\ \text{Calcium} & & & & \text{Hydrofluoric} & & \text{Calcium} \\ \text{Fluoride} & & & & \text{Acid (gas)} & & \text{Sulphate} \\ 78 & + & 98 & = & 40 & + & 136 \end{array} \qquad (27)$$

and —

$$\begin{array}{ccccccc} SiO_2 & + & 4\,HF & = & SiF_4 & + & 2\,H_2O \\ \text{Silicon} & & & & \text{Silicon} & & \\ \text{Dioxide} & & & & \text{Tetrafluoride} & & \\ 60 & + & 80 & = & 104 & + & 36 \end{array} \qquad (28)$$

It will be learned later that silicon dioxide is the essential constituent of the mixture which is made into glass. Silicon tetrafluoride is a gas.

### CLASS-ROOM EXERCISE. XLI.

1. Fluorine.
   - (*a*) Occurrence.
   - (*b*) Isolation by Moissan.
   - (*c*) Properties.
   - (*d*) Liquefaction of fluorine.
2. Further study of hydrofluoric acid.
   - (*a*) Industrial preparation of the commercial acid.
   - (*b*) Properties.
   - (*c*) Preservation and transportation.
   - (*d*) Vapor density.
   - (*e*) Uses.
3. Miscellaneous.
   - (*a*) Compare the properties of fluorine with those of chlorine, bromine, and iodine.
   - (*b*) Compare the properties of hydrofluoric acid with the corresponding acid of the other halogens.
4. The periodic classification of the elements.
   - (*a*) Natural tendency of elements to assume groups.
   - (*b*) The Periodic Law stated, expanded, and illustrated by the halogen groups of elements.
   - (*c*) Value of the law in correcting atomic weights and in predicting new elements.
   - (*d*) Explain the statement: "Argon may cause a rearrangement of the periodic classification."

## PROBLEMS. XXV.

1. What is the percentage composition of—

   (*a*) Fluorspar, $CaF_2$?

   (*b*) Cryolite, $Na_3AlF_6$?

2. If 100 gm. of fluorspar are heated with sulphuric acid, how much calcium sulphate and hydrofluoric acid are formed ?

3. If 100 parts of calcium fluoride yield 174.5 parts of calcium sulphate, what is the atomic weight of fluorine ? (Assume $CaF_2$ and $Ca = 40$.)

## CHAPTER XVI.

## SULPHUR AND ITS COMPOUNDS.

SULPHUR is one of the most important of the non-metallic elements.

### PHYSICAL PROPERTIES OF SULPHUR.

**Experiment 129.** — (*a*) Examine a piece of roll sulphur and write a short account of its most obvious physical properties.

(*b*) Weigh a lump of roll sulphur to a decigram. Slip it carefully into a graduate previously filled with water to a known point — about half full — and note the increase in the volume of water. This increase in volume is equal to the volume of the sulphur. Calculate the specific gravity of sulphur from the observed data.

### AMORPHOUS SULPHUR.

**Experiment 130.** — Put a few pieces of roll sulphur in an old test tube. Heat carefully until the sulphur boils, and then quickly pour the contents of the test tube into a dish of cold water. This is amorphous sulphur. Note its properties. Preserve, and examine it after twenty-four hours. Describe the change, if any.

Define amorphous, and illustrate it by this experiment.

Sulphur exists not only as the amorphous variety, but it crystallizes under favorable conditions in two distinct crystal systems.

### CRYSTALLIZED SULPHUR.

**Experiment 131.** — (*a*) *Monoclinic.* Fill a small Hessian crucible nearly full of roll sulphur. Support the crucible in a ring, and heat until all the sulphur is melted. Let it cool, and the instant crystals shoot out from the walls just below the surface, pour the remaining melted sulphur into a dish of cold water. When the crucible can be handled without discomfort, crack it open lengthwise. The contents should resemble the conventional sketch of a typical mass of crystals shown in Fig. 80. Observe and record the properties of the crystals, especially the shape, size, color, luster, brittleness or elasticity, and any other characteristic property. Allow the best crystals to remain undisturbed for a day or two; then reëxamine, and record any marked changes.

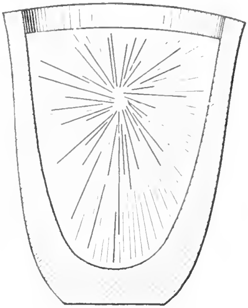

FIG. 80. — Section of a crucible showing monoclinic crystals of sulphur.

(*b*) *Orthorhombic*. Put 3 gm. of flowers of sulphur in a test tube and add about 5 cc. of carbon disulphide — remember the **precaution** to be observed in using this liquid (see Exp. 123). Shake until all the sulphur is dissolved, then pour the clear solution into a small dish to crystallize. It is advisable, though not absolutely necessary, to stand the dish in the hood, where there is no flame and where the offensive vapor will be quickly removed. Watch the crystallization toward the end, and, if perfect crystals form, remove them with the forceps (see Fig. 81). Allow the liquid to evaporate almost entirely, then remove and dry the crystals. Examine them as in (*a*) and record their properties.

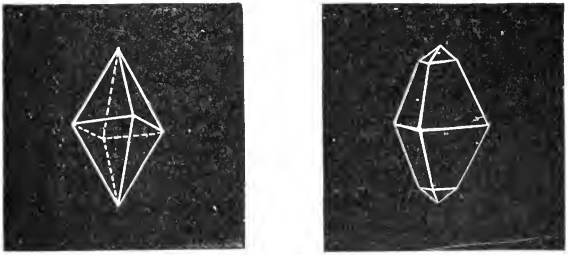

FIG. 81. — Orthorhombic crystals of sulphur.

(*c*) Tabulate the essential results in (*a*) and (*b*).

(*d*) Make an outline sketch of a crystal of orthorhombic sulphur.

**Sulphur combines** with both non-metals and metals, and forms *sulphides* in most cases.

### COMBINING POWER OF SULPHUR.

**Experiment 132.** — (*a*) Set fire to a little sulphur in a deflagrating spoon, and lower the spoon into a bottle. *Cautiously* waft the fumes toward the nose, and observe and describe the odor. The product is a compound of sulphur. From what has already been shown experimentally about combustion, state the components of the fumes.

(*b*) Mix 3 gm. of flowers of sulphur with 3 gm. of iron powder or clean iron filings, and heat half of the mixture in an old test tube. As soon as there is decided evidence of chemical action, remove the test tube from the flame. When the tube has cooled, break the end by rapping it sharply on a hard surface. Examine the contents. What is the evidence of the formation of a new compound? Verify the conclusion by adding a little dilute hydrochloric acid to the product and then to the remainder of the original mixture, testing the gaseous product in each case by the odor.

Summarize the evidence regarding the power of sulphur to combine with iron.

Similar results are obtained with copper and other metals. Draw a general conclusion regarding the combining power of sulphur.

### CLASS-ROOM EXERCISE. XLII.

1. Additional study of sulphur.
    - (*a*) Occurrence of free sulphur.
    - (*b*) Occurrence of combined sulphur.
    - (*c*) Theories of the volcanic and the organic formation of sulphur.
    - (*d*) Extraction of sulphur from native sulphur.
    - (*e*) Other sources of sulphur.
    - (*f*) Purification.
    - (*g*) What is roll sulphur? Brimstone? Flowers of sulphur?
    - (*h*) Additional properties, *e.g.* solubility.
    - (*i*) Illustrate allotropy by means of sulphur.
    - (*j*) What is milk of sulphur? (*k*) Uses.
2. Miscellaneous.
    - (*a*) Review of the orthorhombic and monoclinic systems of crystallization. (See CLASS-ROOM EXERCISE. VII., 7 (*d*).)
    - (*b*) Illustrate the term *amorphous* by sulphur.
    - (*c*) What part has sulphur played in the history of chemistry?
    - (*d*) In what crystal form does free crystallized sulphur occur?

**Hydrogen Sulphide.** — An important compound of sulphur is the gas, hydrogen sulphide. It is formed in many chemical changes, it possesses characteristic properties, it forms many compounds, and is an indispensable laboratory reagent. It is a *poisonous* gas, and should never be liberated in the laboratory.

### PREPARATION AND PROPERTIES OF HYDROGEN SULPHIDE.

*Perform this experiment in the hood, and do not inhale the gas.*

**Experiment 133.**—Construct a generator like that shown in Fig. 82. The test tube *A* is clamped to an iron stand at such a height that the end of *C* reaches to the bottom of a bottle in which the gas is collected by downward displacement and the solution subsequently prepared. The essential part of the apparatus is also shown in Fig. 54, the tube *C* being joined to *B* by the rubber connector *a*. Fill the test tube one-third full of coarsely powdered ferrous sulphide, cover with dilute hydrochloric acid, and quickly insert the stopper. Hydrogen sulphide gas is rapidly evolved. Diffusion of the gas may be prevented by covering the bottle with the block of wood used in the collection of chlorine, or with a piece of stiff paper shaped as shown in Fig 83, and cut along the line *ab* to allow it to be slipped

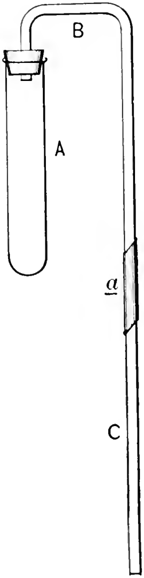

FIG. 82. — Apparatus for generating hydrogen sulphide gas.

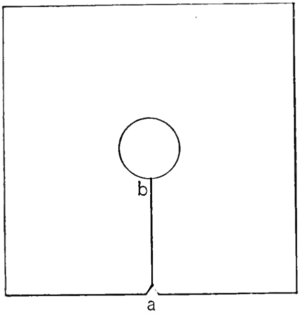

FIG. 83. — Perforated card to prevent the diffusion of hydrogen sulphide gas.

easily upon the delivery tube. Collect three small bottles, removing each, and covering with a glass plate as soon as full. If the evolution of gas stops or slackens, warm gently or add more hydrochloric acid. After the bottles of gas have been collected, allow the gas from the generator to bubble through a bottle half full of water. Hydrogen sulphide water will be formed, which should be preserved in a stoppered bottle for future use; or, if more convenient, it may be poured into the stock bottle.

Meanwhile test the hydrogen sulphide gas as follows: —

(*a*) Waft a *little* of the gas cautiously toward the nose, and then describe the odor. This is characteristic of hydrogen sulphide, and is a decisive test. Has the gas color?

(*b*) Test the gas from the same bottle with both kinds of litmus paper. Is it acid or alkaline?

(*c*) Bring a lighted match to the mouth of a bottle. Observe the properties of the flame as in previous experiments. Observe cautiously the odor of the product of the burned gas; to what compound is the odor due? What, then, is *one* component of hydrogen sulphide?

(*d*) Burn another bottle of hydrogen sulphide and hold a cold bottle over the burning gas. What additional experimental evidence does this result give regarding the composition of hydrogen sulphide?

### LABORATORY EXERCISE. XXI.

1. Summarize the properties of hydrogen sulphide and the experimental evidence of its composition.

2. Complete the equation for the combustion of hydrogen sulphide: —

$$2\,H_2S \;+\; O \;=\; 2\,SO_2 \;+$$

The simplest equation for the preparation of hydrogen sulphide gas is —

$$\underset{\text{Ferrous Sulphide}}{FeS} + 2\,HCl = \underset{\text{Hydrogen Sulphide}}{H_2S} + \underset{\text{Ferrous Chloride}}{FeCl_2}$$

$$88 + 73 = 34 + 127$$

**A liter of dry hydrogen sulphide gas** under standard conditions weighs 1.542 gm.

**Sulphides.** — Hydrogen sulphide is a weak acid and forms salts called sulphides. Many of these sulphides have characteristic colors and they vary in solubility, hence they are easily detected and separated. They are readily formed by adding to solutions of compounds of metals a solution of hydrogen sulphide gas, — usually called hydrogen sulphide water, — or by passing the gas through the solution of the metal; the moist gas itself acts immediately on some metals.

### PREPARATION AND PROPERTIES OF SOME SULPHIDES.

**Experiment 134.** — (*a*) Shake the bottle of hydrogen sulphide water prepared in Exp. 133 (or a similar solution), and hold successively at the mouth or in the neck of the bottle (1) a clean copper wire or piece of sheet copper, (2) a bright strip of lead, and (3) an untarnished silver coin. Describe the result in each case. These compounds are sulphides of the respective metals.

(*b*) Put a gram of litharge — the brownish-yellow oxide of lead — in a test tube, cover it with hydrogen sulphide water, and warm gently. Describe and explain the change. The formula of pure litharge is PbO. Complete the equation: —

$$PbO + H_2S = H_2O +$$

(*c*) Prepare the following solutions in separate test tubes: —

(1) Dissolve a small crystal of lead nitrate in water.

(2) Boil a minute quantity of arsenic trioxide in dilute hydrochloric acid.

(3) Dissolve a minute quantity of tartar emetic — a compound of the metal antimony — in water.

(4) Dissolve a small crystal of zinc sulphate in water.

All the above solutions are *poisonous*. The test tubes should be about half full. Add about one-fourth its bulk of hydrogen sulphide water to each solution, and shake thoroughly. Observe the color of each sulphide, and tabulate the results.

The varying solubility of sulphides permits —

### THE SEPARATION OF SULPHIDES.

**Experiment 135.** — Mix equal quantities of dilute solutions of (1) copper sulphate, (2) ferrous sulphate, and (3) sodium chloride. Add dilute hydrochloric acid. Pass hydrogen sulphide gas through the mixture (use the same generator as in Exp. 133), or add considerable hydrogen sulphide water. In either case, warm gently, allow the precipitate to settle, and then add a few drops of hydrogen sulphide water to ensure complete precipitation. If precipitation is not complete, add more hydrogen sulphide water, and proceed as above. If complete, decant the clear liquid into a clean vessel by pouring it down a glass rod, and save it for subsequent examination. Add dilute hydrochloric acid to the precipitate, warm gently, and filter. There are now three parts to test. Proceed as follows: —

(1) *Precipitate.* Stand the funnel in a clean test tube, make a hole with a pointed glass rod in the apex of the filter paper, and wash the precipitate into the test tube with dilute nitric acid by pouring the acid down a glass rod directly upon the solid. Use as little acid as possible. Boil until the black precipitate is dissolved, and then add considerable ammonium hydroxide. The solution should become a distinct blue, which is evidence that a copper compound is present. Had an iron compound been present, the result would be entirely different.

(2) *Filtrate.* Add ammonium hydroxide to the filtrate, and a black precipitate will be formed. This is iron sulphide. Dissolve it in hydrochloric acid, add a few drops of nitric acid, boil, and then add ammonium hydroxide. A rusty-red precipitate of ferric hydroxide is formed, showing that an iron compound was the main constituent of the filtrate.

(3) *Clear liquid.* Filter, if this liquid is not free from particles. Apply the flame test for sodium to this liquid.

Summarize the results of the whole experiment.

The above experiment illustrates briefly the general method of separation and detection of metals employed in Qualitative Analysis. Sulphides are not the only compounds, however, which permit systematic separation. A more explicit treatment of this subject will be found in Chapter XIX., page 322.

### CLASS-ROOM EXERCISE. XLIII.

1. Additional study of hydrogen sulphide.
    - (*a*) History.
    - (*b*) Names and their significance.
    - (*c*) Occurrence.
    - (*d*) Properties not experimentally shown.
    - (*e*) Uses.
2. Sulphides.
    - (*a*) Natural sulphides.
    - (*b*) Exact name and color of sulphides of arsenic, antimony, zinc, lead.
    - (*c*) Why does a silver spoon blacken in mustard? In a cooked egg?
    - (*d*) Why does white paint often turn brown or black?
    - (*e*) Why does silverware blacken?
    - (*f*) Name and formula of cinnabar, stibnite, copper pyrites, and pyrite.
3. Composition of hydrogen sulphide gas.

(1) If a spiral of platinum wire is heated in a bottle of hydrogen sulphide gas, a yellow powder and an invisible gas are formed; the powder burns with a pale blue flame, producing a choking vapor, and the invisible gas burns with a colorless, hot flame. What evidence do these facts present about the composition of hydrogen sulphide?

(2) Sulphur is often deposited in bottles of hydrogen sulphide water. Chlorine passed into hydrogen sulphide forms sulphur and hydrochloric acid. What do these facts show about the composition of hydrogen sulphide?

(3) Give additional evidence of the composition of hydrogen sulphide.

### PROBLEMS. XXVI.

1. How much does a stick of brimstone 1 m. long and 4 cm. in diameter weigh?

2. Dumas heated 10.000 grains of silver in sulphur vapor and obtained 11.481.5 grains of silver sulphide. What is the atomic weight of sulphur? (Assume Ag = 108, and $Ag_2 + S = Ag_2S$.)

3. How many grams of sulphur can be obtained from 600 cc. of hydrogen sulphide gas?

4. How many grams of ferrous sulphide are necessary to yield a liter of hydrogen sulphide gas?

5. Calculate the percentage composition of —

(*a*) Galena, $PbS$.
(*b*) Zinc blende, $ZnS$.
(*c*) Iron pyrites, $FeS_2$.
(*d*) Ferrous sulphide, $FeS$.

6. What weight and what volume of hydrogen can be obtained from 102 gm. of hydrogen sulphide?

**The two oxides of Sulphur**, sulphur dioxide and sulphur trioxide, and their related acids, are the most important compounds of sulphur.

### PREPARATION AND PROPERTIES OF SULPHUR DIOXIDE.

*Perform this experiment in the hood.*

**Experiment 136.** — Construct a generator like that shown in Fig. 82. The test tube should be clamped high enough to be conveniently heated with a free flame and the delivery tube must be long enough to reach to the bottom of an empty bottle. Put a few grams of copper borings in the test tube, cover with concentrated sulphuric acid, and arrange the apparatus to collect a gas as in Exp. 133. Heat gently with a low flame, and sulphur dioxide gas is evolved. As the reaction proceeds, the evolution of gas may become violent. Remove the flame at intervals, regulating the heat so that the evolution of gas is slow and steady. Remember that hot sulphuric acid is a dangerous liquid. Collect four small bottles of gas, covering each with a glass plate as soon as full. Then remove the stand with its attached generator to the hood or some place where it will not be disturbed. Study the sulphur dioxide gas as follows: —

(*a*) Observe the most obvious physical properties, *e.g.* color, odor (cautiously), specific gravity.

(*b*) Prove that it is the same gas as was formed in Exp. 132 (*a*).

(*c*) Hold a blazing stick in a bottle of the gas. Will the gas burn or support combustion? What previously acquired facts would have enabled you to predict this result?

(*d*) Pour 10 or 15 cc. of water into a bottle of sulphur dioxide, cover with the hand, and shake. What is the evidence of solution? Is the resulting liquid acid, alkaline, or neutral?

(*e*) Moisten a pink or purple flower with a few drops of water, hang it in the remaining bottle of sulphur dioxide, holding it in place by putting the stem between the glass and a cork. Observe and describe any change in the color of the flower. If the action is slow or indefinite, stand a deflagrating spoon containing burning sulphur in the bottle for a minute or two. What is this operation called?

### LABORATORY EXERCISE. XXII.

1. Light a sulphur match and observe and describe the entire action, as far as the sulphur is concerned.
2. What is the function of the sulphur in a burning match?

**Sulphurous Acid** is formed when sulphur dioxide dissolves in water. Sulphur dioxide is, therefore, sulphurous anhydride. The simplest equation expressing this fact is —

$$SO_2 + H_2O = \underset{\text{Sulphurous Acid}}{H_2SO_3}$$

It is an unstable acid, and gradually forms sulphuric acid by combining with oxygen from the air. Its salts — the *sulphites* — yield sulphur dioxide when treated with acids.

### CLASS-ROOM EXERCISE. XLIV.

1. Additional study of sulphur dioxide.
    - (*a*) Occurrence.
    - (*b*) Additional properties, *e.g.* solubility, specific gravity.
    - (*c*) Relation to combustion.
    - (*d*) Explanation of bleaching power.
    - (*e*) Liquid sulphur dioxide.
    - (*f*) Composition and formula.

**Sulphur Trioxide,** $SO_3$, is formed when a mixture of sulphur dioxide and oxygen is passed over heated platinized asbestos. It is the anhydride of sulphuric acid. The equation expressing the latter fact is —

$$SO_3 + H_2O = H_2SO_4$$

Sulphuric acid has not, until recently, been industrially prepared in this way. Its preparation is not well adapted to the laboratory, but its properties may be profitably studied.

### ACTION OF SULPHURIC ACID WITH WATER.

**Experiment 137.** — Recall or repeat that part of Exp. 36 which shows the action of sulphuric acid with water. Record the result.

This experiment illustrates the necessity of always *adding the acid to the water*. The most obvious result is due to the formation of a compound of sulphuric acid and water. The experiment also proves the avidity with which sulphuric acid unites with water. This affinity often decomposes compounds which contain hydrogen and oxygen in the proportion to form water.

### ACTION OF SULPHURIC ACID WITH ORGANIC MATTER.

**Experiment 138.** — (*a*) Write some letters or figures with dilute sulphuric acid on a sheet of white paper, and move the paper back and forth over a low flame, taking care not to set fire to the paper. As the water evaporates the dilute acid becomes concentrated. Observe and describe the result. Paper is largely a compound of carbon, hydrogen, and oxygen, and the hydrogen and oxygen are present in the proportion to form water. Explain the general chemical change in this experiment.

(*b*) Dissolve 3 or 4 gm. of sugar in a test tube half full of warm water, stand the test tube in the rack, and add *cautiously* several drops of concentrated sulphuric acid. If there is no decided result, add more acid. What is the black product? Compare the final result with that obtained in Exp. 82 (*b*). Is the chemical action the same in each experiment? Are the statements made in (*a*) about paper also true of sugar?

(*c*) Repeat (*b*), using powdered starch instead of sugar. Describe the result. How does the result resemble that in (*b*) and in Exp. 82 (*a*)? Predict the components of starch. In what simple way may the prediction be verified?

(*d*) Stand a stick of wood in a test tube one-fourth full of concentrated sulphuric acid. Allow it to remain in the acid for fifteen minutes, then remove the stick and wash off the acid. Describe the change. Does it resemble that in (*a*), (*b*), and (*c*), and in Exp. 82 (*a*)?

### LABORATORY EXERCISE. XXIII.

1. Can sulphuric acid be used to dry gases? Can it be used to dry ammonia gas? Why?

2. Is sulphuric acid volatile? What experiment proves this?

3. What happens when concentrated sulphuric acid is exposed to moist air?

4. Determine the specific gravity of sulphuric acid by weighing a measured volume (25 cc.) in a graduate. Compare the result with the specific gravity of the commercial acid.

### TEST FOR SULPHURIC ACID.

**Experiment 139.** — Devise a test for sulphuric acid. Verify it. Is the same test applicable to soluble sulphates? Devise, or find in a text-book, a test for sulphates almost insoluble in water, *e.g.* calcium sulphate. Verify it. How can a sulphate be distinguished from a sulphite? Is barium sulphate absolutely insoluble in all reagents?

### CLASS-ROOM EXERCISE. XLV.

1. Additional study of sulphuric acid.
    - (*a*) Complete account of its industrial preparation.
    - (*b*) Properties not experimentally found.
    - (*c*) Uses. (*d*) History.
    - (*e*) Significance of the term *oil of vitriol.*
2. Miscellaneous.
    - (*a*) Explain and illustrate the statement, "sulphurous and sulphuric acid are dibasic."
    - (*b*) Valence of sulphur in $H_2S$. $SO_2$, $SO_3$, $H_2SO_3$, $H_2SO_4$.
    - (*c*) Give the chemical name and formula of gypsum, heavy spar, plaster of Paris, white vitriol, blue stone, green vitriol, Glauber's salt, kieserite.
    - (*d*) Illustrate the Periodic Law by sulphur, selenium, and tellurium.
    - (*e*) What is "hypo"? What is sodium thiosulphate? Use of the latter.

## PROBLEMS. XXVII.

1. What volume of air is necessary to burn 1000 gm. of sulphur?

2. How many grams of sulphur can be obtained from a liter of sulphuric acid? (Assume the commercial density.)

3. Calculate the percentage composition of—

(*a*) Barium sulphate.
(*b*) Zinc sulphate.
(*c*) Copper sulphate.
(*d*) Sodium sulphate.
(*e*) Chamber crystals. $SO_2.NO_2.HO$.
(*f*) Nordhausen acid, $H_2S_2O_7$.

# CHAPTER XVII.

## SILICON AND PHOSPHORUS AND THEIR COMPOUNDS.—REVIEW.

SILICON does not occur in the free state, but its compounds are abundant and widely distributed. Approximately one-fourth of the earth's crust is silicon. Its most important compound is silicon dioxide, $SiO_2$, which is familiar as sand and quartz.

The properties of some compounds of silicon are shown by the

### RELATION OF SILICON DIOXIDE TO OTHER COMPOUNDS OF SILICON.

**Experiment 140.**—Fill a porcelain or platinum crucible half full of a fusion mixture—sodium carbonate (4 parts) and potassium carbonate (5 parts). Support the crucible on a tripod, or a triangle, and heat gently, until the fusion mixture is melted. Then add a pinch of clean, exceedingly fine sand, and heat with the hottest available flame—a blast lamp flame, if possible. Add three or four pinches of sand from time to time, and finally heat until no particles of sand can be seen suspended in the molten mass. The heat must be intense and steadily applied, but so regulated that there is no undue frothing. Cool the crucible gradually, especially if it is porcelain, and then place it in a beaker or casserole; cover with water and boil gently until the mass is detached. Remove the crucible, and heat until the residue is dissolved. Filter, if the solution contains any sand. Add dilute hydrochloric acid slowly to the clear liquid until the carbonates are decomposed. The gelatinous precipitate is silicic acid. Rub some between the fingers and describe the result. Evaporate the precipitate to dryness upon a water bath in a porcelain dish, or casserole, in the hood. As the mass hardens, stir it with a glass rod. Toward the end, add more hydrochloric acid and evaporate to complete dryness. Then stand the dish on a gauze-covered tripod and heat strongly for five minutes. The residue is silicon dioxide mixed with chlorides of sodium and potassium. Rub some between the fingers or across a glass plate. Is any grit detected?

If the apparatus required for Exp. 140 is not available, the same results may be shown by the

### PREPARATION AND PROPERTIES OF SILICIC ACID.

**Experiment 141.** — Add hydrochloric acid to a solution of sodium silicate — "water glass." The gelatinous precipitate is silicic acid. Proceed from this point as in Exp. 140.

**Discussion of Experiment 140.** — When silicon dioxide is fused with alkaline carbonates, an alkaline silicate is formed. The alkaline silicates are soluble in water, and when hydrochloric acid is added to such a solution, the gelatinous precipitate formed is a silicic acid having the formula $H_2SiO_3$ (probably). This acid is decomposed by heat into silicon dioxide and water, thus: —

$$\underset{\text{Silicic Acid}}{H_2SiO_3} = \underset{\text{Silicon Dioxide}}{SiO_2} + H_2O$$

The silicic acids are numerous and complex, and the silicates which they form constitute a large class of important minerals and rocks.

### CLASS-ROOM EXERCISE. XLVI.

1. Additional study of silicon.
   - (*a*) Occurrence and distribution.
   - (*b*) Preparation and properties of amorphous and of crystallized silicon.
   - (*c*) Resemblance to carbon.

2. Silicon dioxide.
   - (*a*) Occurrence and distribution.
   - (*b*) Properties of various forms.
   - (*c*) Geological significance.
   - (*d*) Uses.
   - (*e*) Reason for abundance.
   - (*f*) Meaning of the term *silica.*

3. Other compounds of silicon.
   - (*a*) Silicon tetrafluoride.
   - (*b*) Silicic acids.
   - (*c*) Silicates.
   - (*d*) Carborundum.

4. Miscellaneous.
   - (*a*) What is diatomaceous earth? Examine it through a microscope.
   - (*b*) What is "electro-silicon"?
   - (*c*) What is opal?
   - (*d*) Crystallization of quartz.
   - (*e*) Significance of the term *rock crystal.*
   - (*f*) What is the essential chemical constituent of many rocks?
   - (*g*) Write an essay on "Glass."
   - (*h*) Prepare a short account of the manufacture of pottery and porcelain.

## PROBLEMS. XXVIII.

1. Scheele found that 0.6738 gm. of silicon tetrachloride gave 2.277 gm. of silver chloride. Calculate the atomic weight of silicon. (Assume the equation, $SiCl_4 + 4\,AgNO_3 + 2\,H_2O = SiO_2 + 4\,AgCl + 4\,NH_4NO_3$.)

2. How much silicon can be obtained from 119 gm. of potassium silicofluoride? (Assume the equation, $K_2SiF_6 + 4\,K = Si + 6\,KF$.)

3. Calculate the percentage composition of —
   - (*a*) Willemite, $Zn_2SiO_4$.
   - (*b*) Steatite, $Mg_3Si_4O_{11} \cdot H_2O$.

**Phosphorus** itself is not adapted to extended laboratory examination on account of its properties. It takes fire suddenly and produces dangerous burns. The element has some uses, and its compounds have many industrial applications.

### LABORATORY EXERCISE. XXIV.

1. Recall or read again the notes made about the properties of phosphorus observed in Exp. 45.

2. Smell of the end of a phosphorus-tipped match, then smell of the fumes arising from a bottle containing phosphorus under water. Do they suggest each other?

3. If convenient, rub the end of a phosphorus-tipped match in the dark, and observe and describe the result.

### CLASS-ROOM EXERCISE. XLVII.

1. A study of yellow phosphorus.
    - (*a*) History.
    - (*b*) Significance of name.
    - (*c*) Occurrence and sources.
    - (*d*) Manufacture.
    - (*e*) Properties.
    - (*f*) Uses.
    - (*g*) Danger in its use.
    - (*h*) How it should be kept.
2. A study of red phosphorus.
    - (*a*) Preparation.
    - (*b*) Properties.
    - (*c*) Uses.
3. Compounds of phosphorus.
    - (*a*) Preparation and properties of phosphorus pentoxide.
    - (*b*) Orthophosphoric acid, $H_3PO_4$, and its salts.
    - (*c*) Metaphosphoric acid, $HPO_3$, and its salts.
4. Miscellaneous.
    - (*a*) Use of phosphates as food of plants.
    - (*b*) Fertilizer — manufacture, constituents, use.
    - (*c*) Test for phosphorus.
    - (*d*) Write an essay on "Matches."
    - (*e*) What is bone black? Bone ash? A cupel? Apatite? Phosphate rock? Microcosmic salt? Disodium phosphate?
    - (*f*) Illustrate allotropy by phosphorus.
    - (*g*) To what group in the periodic classification does phosphorus belong?

5. A study of arsenic.
   - (*a*) Occurrence. (*b*) Preparation. (*c*) Properties.
   - (*d*) Preparation, properties, and uses of arsenic trioxide.
   - (*e*) What is Marsh's test for arsenic?
   - (*f*) What is "arsenic," white arsenic, arsenious oxide?
   - (*g*) Antidote for arsenic poisoning.

**Review.** — The foregoing part of this book contains most of the fundamental principles of chemistry illustrated by experiments as far as such a method permits. It is advisable at this stage to glance backward and review these principles in the light of the facts gleaned from all sources.

### CLASS-ROOM EXERCISE. XLVIII.

1. Discuss the atomic theory as an expression of the Laws of Multiple and Definite Proportions.

2. Discuss molecules and molecular weight in the light of the Laws of Avogadro and Gay-Lussac, vapor density, vapor tension, elevation of the boiling point, and depression of the freezing point.

3. Illustrate Gay-Lussac's Law by the volumetric formation of water, ammonia gas, hydrochloric acid gas, hydrogen sulphide, sulphur dioxide, and by the oxides of carbon.

4. Illustrate the Law of Multiple Proportions by the oxides of nitrogen, carbon, and sulphur, and by the hydrocarbons, methane, ethylene, and acetylene.

5. Illustrate the Law of Definite Proportions by Exps. 15, 16, 22, 26, 44, 54, 55, 92.

6. Discuss the causes, accompaniments, and results of chemical action.

7. Discuss: —
   - (*a*) Heat of combustion.
   - (*b*) Solution.
   - (*c*) Allotropism.
   - (*d*) Valence.

8. Discuss with illustrations: —
   - (*a*) The Periodic Law.
   - (*b*) Reduction and oxidation.
   - (*c*) Equations.
   - (*d*) Elements and compounds.
   - (*e*) Symbols and formulas.

9. State the test for a chloride, sulphate, sulphide, sulphite, nitrate, carbonate, and for Zn, Na, Ca, Cu, Pb, Ag, C. O, H, Sb, As, $CO_2$, N, Cl, Br, I, S, $H_2S$, $SO_2$, $SiO_2$, P, an alkali, an acid, and a neutral salt.

10. State the chemical name and formula of lime, lime water, limestone, chloride of lime, bleaching powder, marble, muriatic acid, oil of vitriol, blue vitriol, ammonia, sal ammoniac, salt, soda, potash, caustic soda, quicklime, slaked lime, cream of tartar.

11. Prepare a chronological table of the following chemists who may be regarded as the founders of modern chemistry: Boyle, Becher, Stahl, Black, Cavendish, Priestley, Dalton, Davy, Faraday, Graham, Liebig, Lavoisier, Berthollet, Proust, Gay-Lussac, Dumas, Scheele, Stas, Berzelius.

12. Discuss the principles underlying the nomenclature of acids, bases, and salts.

## PROBLEMS. XXIX. (Review.)

1. How many grams of hydrogen may be obtained by the interaction of zinc and 196 gm. of pure sulphuric acid?

2. How much mercury will remain after obtaining 48 gm. of oxygen by heating mercuric oxide?

3. What weight of potassium chlorate is necessary to produce 144 gm. of oxygen?

4. What weight of potassium chloride remains after obtaining 8 gm. of oxygen from potassium chlorate?

5. How much zinc is necessary to obtain 2 l. of hydrogen?

6. How many grams of oxygen can be obtained from 490 gm. of potassium chlorate?

7. If the density of ice is 0.92, what volume will a liter of water at 4° C. occupy when frozen?

8. The formula of crystallized gypsum is $CaSO_4 . 2\,H_2O$. What volume of steam at 300° C. is liberated by heating 100 gm. of gypsum?

9. Ten liters of hydrogen at 30° C. are passed over hot copper oxide. How much copper is set free, and what volume of steam at 300° C. is formed? (The equation is $CuO + H_2 = Cu + H_2O$.)

10. A quantity of carbon monoxide was passed through a red-hot tube containing ferric oxide ($Fe_2O_3$). The resulting gas was absorbed by caustic potash. The gain in weight of the caustic potash was 0.86 gm. What was the volume of the carbon monoxide at the standard temperature and pressure?

## CHAPTER XVIII.

## SODIUM — POTASSIUM — COPPER — SILVER — MAGNESIUM — CALCIUM — ZINC — MERCURY.

**Metals.** — Most of the elements hitherto studied are gases. The few solids — carbon, sulphur, phosphorus, and iodine — bear almost no resemblance to the elements called metals. The popular conception of a metal is that of a lustrous solid, possessing certain characteristics, *e.g.*, malleability or ductility, tenacity, varying degrees of hardness, etc. It is more satisfactory to make no sharp division into non-metals and metals, but to regard all the elements as a series, broken into related groups the members of which have a marked gradation of properties, some elements emphasizing basic, some acid, and some both properties. If an element emphasizes basic properties at any or all times, it is then a metal. (See page 128.)

### PROPERTIES OF SODIUM.

*Remember the precautions to be observed in using sodium.* (See page 89.)

**Experiment** 142. — (*a*) Examine a small piece of sodium, and record its most obvious physical properties, *e.g.* color, luster, whether hard or soft, etc.

(*b*) Repeat Exp. 42, and read its accompanying text.

(*c*) Answer the following: —

(1) Is sodium heavier or lighter than water?
(2) What properties suggest that it is a metal?
(3) Is it harder or softer than most metals?

**The Principal Compounds of Sodium** are the chloride, sulphate, carbonate, nitrate, and hydroxide.

### GENERAL PROPERTIES OF SODIUM COMPOUNDS.

**Experiment 143.** — Recall (or determine again) the flame test, solubility in water, and color of the most common sodium compounds. Draw legitimate conclusions.

**Sodium Chloride** is the most abundant compound of sodium. It is familiar to all under the name salt, or common salt.

### PREPARATION OF PURE SODIUM CHLORIDE.

**Experiment 144.** — Grind 50 gm. of salt with 150 cc. of water in a mortar, adding the salt in small portions. Filter this saturated solution into a dish or beaker, and conduct hydrochloric acid gas into it. Generate the gas by the method employed in Exp. 61 or Exp. 63. A funnel, or calcium chloride tube, or a thistle tube, should be attached to the end of the delivery tube so that it will dip just below the surface of the salt solution. As the operation proceeds pure sodium chloride separates, but the magnesium chloride and other impurities remain in solution. When considerable sodium chloride has separated, remove the generator, allow the solid to settle, decant the clear liquid, add 15 cc. of cold water, and decant again. If a filter pump is available, remove the last traces of water with it; if not, wash and decant once or twice as before. Finally dry the salt by heating in a porcelain dish; stir constantly with a glass rod during the heating. Describe it.

Preserve a specimen in a small tube or bottle.

**Sodium Hydroxide** is a useful compound of sodium. Its properties have already been studied.

### PREPARATION OF SODIUM HYDROXIDE.

**Experiment 145.** — Dissolve 25 gm. of sodium carbonate in 150 cc. of water and heat gently in an iron dish (an ordinary iron spider is well adapted for this work). Meanwhile slake 10 gm. of lime and add just enough water to make a milky liquid — "milk of lime." Add the milk of lime to the sodium carbonate solution and boil for several minutes, stirring constantly with an iron wire or file. Let the precipitate subside, remove a little liquid with a pipette or small tube, and if it effervesces with hydrochloric acid, add more milk of lime and boil; if not, decant the clear liquid into a convenient vessel, let it stand for a few minutes or until the solid settles, and siphon off the liquid through a glass siphon. This solution of sodium hydroxide may be evaporated to dryness and the solid product tested and the remainder preserved, or the solution may be tested at once as follows: —

(*a*) Rub a little between the fingers and describe the feeling.

(*b*) Add to a small portion a solution of zinc sulphate. Add an excess of sodium hydroxide. These changes have already been explained (Exp. 21). Reconsider them, and state what they prove about this solution.

(*c*) Try the flame test.

(*d*) Determine the specific gravity by weighing a measured volume.

## CLASS-ROOM EXERCISE. XLIX.

1. Additional study of sodium.
    (*a*) Occurrence. (*b*) Discovery.
    (*c*) Industrial preparation.
    (*d*) Properties not experimentally shown. (*e*) Uses.
    (*f*) Derivation and significance of the words *natrium* and *sodium*.
2. Oxides of sodium.
    (*a*) Monoxide.
    (*b*) Peroxide — preparation, properties, and uses.
3. Sodium hydroxide.
    (*a*) Names and formula. (*c*) Additional properties.
    (*b*) Industrial preparation (*d*) Uses.
4. Sodium chloride.
    (*a*) Sources. (*b*) Purification.
    (*c*) Properties not experimentally found.
    (*d*) Uses. (*e*) Physiological value.
5. Sodium carbonate.
    (*a*) Names.
    (*b*) Industrial preparation, including the most important equations in each process. (*c*) Uses.
    (*d*) Chemical name and formula of soda, washing soda, soda crystals, alkali, soda ash.

6. Sodium bicarbonate.
   - (*a*) Formula.
   - (*b*) Names and their significance.
   - (*c*) Properties and uses.
7. Other sodium compounds.
   - (*a*) Sodium sulphate — source, preparation, properties, uses, historical significance.
   - (*b*) Sodium nitrate — names, source, properties, and uses.
8. Miscellaneous.
   - (*a*) Valence of sodium in NaOH, $Na_2SO_4$, $HNaSO_4$, $Na_2O$, $Na_3PO_4$ (sodium phosphate), $H_2NaPO_4$ (dihydrogen sodium phosphate), $HNa_2PO_4$ (disodium phosphate), $HNa(NH_4)PO_4$ (microcosmic salt).

### PROPERTIES OF POTASSIUM.

*Observe the same precautions as in the use of sodium.*

**Experiment 146.** — (*a*) Examine a very small piece of freshly cut potassium, and record its most obvious physical properties. Touch it slightly. Does it suggest caustic potash and soda?

(*b*) Scrape the coating from a piece of potassium, press it between filter paper to remove the oil, cut off a very small piece, and drop it on the water in a pneumatic trough or similar vessel. Stand just near enough to see the action. Describe the action. How does it differ from the action of sodium? Test the water as in Exp. 142.

From what has already been learned about sodium and potassium, predict the main chemical change observed in (*b*) and write the equation.

(*c*) Answer the following: —

(1) Is potassium heavier or lighter than water?
(2) What properties suggest that it is a metal?
(3) How does it resemble and differ from sodium?
(4) What is the usual test for potassium?

(*d*) If possible, examine the flame of a potassium compound (*e.g.* potassium chloride) through a spectroscope, and describe the result.

**The Principal Compounds of Potassium** are the chlorate, chloride, hydroxide, nitrate, carbonate, bromide, and iodide.

### GENERAL PROPERTIES OF POTASSIUM COMPOUNDS.

**Experiment 147**. — Recall (or determine again) the solubility of these compounds. Subject them to the flame test. How does each color the flame? Draw a legitimate conclusion.

### LABORATORY EXERCISE. XXV.

1. What experimental use has been made of potassium chlorate, potassium bromide, and potassium iodide?
2. What happens to potassium hydroxide and potassium carbonate when exposed to the air?
3. What is the thermal effect produced by dissolving potassium nitrate and potassium hydroxide separately in water?

**Potassium Hydroxide** has properties which are analogous to those of sodium hydroxide. It is similarly prepared.

### PREPARATION OF POTASSIUM HYDROXIDE.

**Experiment 148**. — Proceed as in Exp. 145, but use potassium carbonate instead of sodium carbonate. Test as in the case of sodium hydroxide.

**Potassium Carbonate** is the historical *potash*, though that term is now usually applied to potassium hydroxide. Some of its properties have already been considered.

### PREPARATION AND PROPERTIES OF POTASSIUM CARBONATE.

**Experiment 149**. — (*a*) Heat strongly 5 gm. of cream of tartar — acid potassium tartrate — in an iron dish in the hood until the residue is white. Grind this solid with water in a mortar, and filter. Test the filtrate (1) with both kinds of litmus paper, (2) for potassium, and (3) for a carbonate. Record the results.

(*b*) Fill a test tube half full of wood ashes, add half the volume of water, shake, and warm gently. Filter, and test the filtrate as in (*a*). If test (3) is not decisive, repeat the experiment on a larger scale. Record the results.

Answer the following: —

(1) What is the source of cream of tartar?

(2) What do (*a*) and (*b*) show about the distribution of potassium? Of its assimilation by plants?

(3) What is the literal meaning of the word *potash?*

(*c*) Mix 5 gm. of powdered potassium nitrate and 1 gm. of powdered wood charcoal. Spread the mixture on an iron dish or pan, place the pan on a tripod or ring, stand a lighted burner under it, and step aside. Wait for a decided result. As soon as the mass is cool, grind it with a little water in a mortar, filter, and test the filtrate as in (*a*). How has the potassium nitrate been changed?

## CLASS-ROOM EXERCISE. L.

1. Additional study of potassium.
   (*a*) Occurrence and sources.
   (*b*) Discovery. (*c*) Industrial preparation.
   (*d*) Properties not experimentally shown. (*e*) Uses.
   (*f*) Derivation and significance of the words *kalium* and *potassium.*
2. Potassium hydroxide.
   (*a*) Names and formula. (*b*) Industrial preparation.
   (*c*) Explain the statement, "the reaction in the preparation of potassium hydroxide is reversible."
   (*d*) Additional properties. (*e*) Uses.
3. Halides of potassium.
   (*a*) Sources, preparation, properties, and uses of each.
   (*b*) Their similarity.
4. Potassium chlorate.
   (*a*) Industrial preparation. (*b*) Properties. (*c*) Uses.
5. Potassium carbonate.
   (*a*) Sources. (*c*) Properties.
   (*b*) Preparation (several methods). (*d*) Uses.
   (*e*) What is pearlash? Potash?
6. Potassium nitrate.
   (*a*) Names and their significance.
   (*b*) Sources. (*d*) Properties.
   (*c*) Preparation. (*e*) Uses.
   (*f*) Review the subject of gunpowder.

7. Sodium and potassium.
    - (*a*) How do they illustrate the Periodic Law?
    - (*b*) "Their molecules are monatomic." Explain this statement.
    - (*c*) Why are they called "alkali metals"?
    - (*d*) Significance of their discovery by Davy.
8. Miscellaneous.
    - (*a*) What is "lye"?
    - (*b*) Valence of potassium in KOH, $K_2SO_4$, $HKSO_4$, $K_2O$, $K_2CO_3$, KBr.
    - (*c*) Elements related to sodium and potassium — especially their properties.
    - (*d*) The terms *volatile*, *fixed*, *mineral*, and *vegetable*, have all been applied to alkali. Explain each term.
    - (*e*) The spectroscope — its construction and use.
    - (*f*) Review "The Stassfurt Deposits" with special reference to potassium compounds.
    - (*g*) Review "Fertilizer" from the standpoint of potash.
9. Ammonium compounds.
    - (*a*) Review the compounds previously mentioned or used.
    - (*b*) Why may they be considered with the "alkali" compounds?

## PROBLEMS. XXX.

1. Stas found that 100 gm. of pure silver (dissolved in nitric acid) required 54.2075 gm. of sodium chloride for complete precipitation. Calculate the atomic weight of sodium. (Assume atomic weights of silver and chlorine and $AgNO_3 + NaCl = AgCl + NaNO_3$.)

2. One hundred gm. of potassium chloride give 52.42 gm. of potassium and 47.58 gm. of chlorine. Calculate the atomic weight of potassium — making two assumptions.

3. How much sodium can be obtained from 2 kg. of sodium carbonate, if the latter is heated with carbon? The equation is —

$$Na_2CO_3 + C_2 = Na_2 + 3\ CO.$$

4. How much potassium carbonate is necessary to prepare a kilogram of potassium hydroxide? The equation is —

$$K_2CO_3 + Ca\ (OH)_2 = 2\ KOH + CaCO_3.$$

5. How much sodium carbonate is necessary to prepare a kilogram of sodium hydroxide?

6. What per cent of Glauber's salt, $Na_2SO_4 . 10 H_2O$ is sodium sulphate?

7. How many pounds of salt are needed to make one-fourth of a ton of salt cake? The equation is —

$$2 NaCl + H_2SO_4 = Na_2SO_4 + 2 HCl.$$

8. How much sodium carbonate can be made from 500 kg. of salt? The first equation is given in problem 7, the other is —

$$Na_2SO_4 + CaCO_3 + 2 C_2 = Na_2CO_3 + CaS + 4 CO.$$

9. How much sodium sulphide is formed by heating 100 gm. of sodium sulphate with charcoal? The equation is —

$$Na_2SO_4 + 2 C_2 = Na_2S + 4 CO.$$

**Copper** is an abundant element, and its compounds are numerous. Its properties, which are familiar, adapt the metal to various uses.

### GENERAL PROPERTIES OF COPPER.

**Experiment 150.** — (*a*) Examine several forms of copper — wire, sheet, filings, etc. — and observe and state the most obvious physical properties.

(*b*) Determine the specific gravity by weighing copper in air and then in water, or by weighing it in air and comparing this weight with the volume of water it displaces when immersed in a known volume in a graduate.

### LABORATORY EXERCISE. XXVI.

1. Is copper a good conductor of heat? Of electricity? On what evidence is your answer based?

2. Is copper ductile? Malleable? Brittle? Tough? Hard or soft?

3. What happens to copper when heated? When exposed to the air?

### TESTS FOR COPPER.

**Experiment 151.** — (*a*) Heat a copper wire in the Bunsen flame. The color is characteristic of copper and its compounds, though not a conclusive test, since the same color is produced by other substances.

(*b*) Add a few drops of ammonium hydroxide to copper sulphate solution, and observe the result; now add an excess of ammonium hydroxide. The blue solution is a characteristic and decisive test for copper.

(*c*) Add to a test tube half full of water a drop of copper sulphate solution and shake; then add a few drops of acetic acid and of potassium ferrocyanide solution. The brown precipitate is copper ferrocyanide.

**Copper replaces** some metals, if suspended in solutions of their compounds, and is in turn replaced by other metals. Both operations precipitate or deposit a metal.

### INTERACTION OF METALS — COPPER.

**Experiment 152.** — (*a*) Put a clean copper wire into a solution of any compound of mercury. After a short time, remove the wire and wipe it with a soft cloth or paper. Describe the change. What has become of some of the copper?

(*b*) Put in separate test tubes half full of copper sulphate solution a bright nail and a strip of clean zinc. After a short time remove the metals and examine them. What is the deposit? What has become of some of the zinc and iron? Does the final color of the solution indicate any chemical change? How would you prove the answer to the last question?

**The Important Compounds of Copper** are the copper oxides, copper sulphate, and copper nitrate.

### LABORATORY EXERCISE. XXVII.

1. What experimental use has been made of one copper oxide?
2. Examine a crystal of copper sulphate and observe its most obvious properties. What happens to it when heated?
3. How has copper nitrate been prepared? Summarize its properties. What is its formula?
4. Prove that each of the above important compounds contains copper.
5. What is the color of many copper compounds?

**Oxides of Copper.** — There are two oxides of copper, the black oxide, already studied (see Exps. 9 and 151 (*a*)), and the red oxide. The formulas are respectively $CuO$ and $Cu_2O$, and their exact names are cupric and cuprous oxides. These oxides are the first members of two classes of copper compounds — the *-ic* and *-ous*. In the former the metal is combined with a larger proportion of oxygen than in the latter, or is in a higher state of oxidation — "farther on in the series." Other metals form similar classes.

### PREPARATION AND PROPERTIES OF CUPROUS OXIDE.

**Experiment 153.** — Dissolve 1 gm. of copper sulphate in 15 cc. of water, and 5 gm. of Rochelle salt (sodium potassium tartrate) in 10 cc. of water; mix these solutions, warm, and filter. Add to this filtrate about 5 cc. of a dilute solution of grape sugar (or of cane suga. which has been boiled at least five minutes with dilute sulphuric acid), and enough sodium hydroxide solution to dissolve any precipitate. The solution must be strongly alkaline. Boil until a decided change is produced. The product is cuprous oxide. Describe it. Show by a comparative table the general properties of the two oxides of copper.

The mixture of copper sulphate, Rochelle salt, and sodium hydroxide is called Fehling's solution, and is widely used to detect the presence of sugar.

**Alloys** are mixtures or compounds of two or more metals. The properties of the alloy vary with the constituents and their proportions. If mercury is one constituent, the alloy is called an *amalgam*. The important alloys in which copper is the larger constituent are brass, bronze, German silver, and their allied alloys.

### CLASS-ROOM EXERCISE. LI.

1. Additional study of copper.
    - (*a*) Occurrence of native and combined copper.
    - (*b*) Preparation from its ores.
    - (*c*) Additional properties.
    - (*d*) Uses.
    - (*e*) History.
    - (*f*) What is electrolytic copper and how is it prepared?
2. Oxides.
    - (*a*) Review the preparation and properties of cuprous and cupric oxides.
    - (*b*) Significance of the terminations *-ous* and *-ic*.

3. Copper sulphate.
   - (*a*) Names.
   - (*b*) Preparation.
   - (*c*) Properties.
   - (*d*) Uses.
4. Copper sulphides.
   - (*a*) Occurrence.
   - (*b*) Properties.
   - (*c*) Names and formulas.
5. Copper compounds.
   - (*a*) Give the chemical name and symbol of ruby ore, copper pyrites, copper glance.
   - (*b*) Give formula and color of azurite and malachite.
6. Miscellaneous.
   - (*a*) What is the valence of copper in $CuO$, $Cu_2O$, $CuSO_4$, $Cu(NO_3)_2$, $Cu_2S$?
   - (*b*) Derivation of the word *cuprum*.
7. Alloys of copper.
   - (*a*) Composition, preparation, properties, and uses of common brass.
   - (*b*) Alloys allied to brass.
   - (*c*) Composition, properties, and uses of (1) gun metal, (2) bell metal, (3) speculum metal, (4) antique bronze, (5) aluminium bronze.
   - (*d*) Composition, properties, and uses of (1) ordinary German silver, (2) Chinese Pakfong, (3) white German silver, (4) casting German silver.
8. Experimental.
   - (*a*) Starting with copper, how would you prepare successively copper nitrate, cupric oxide, and cuprous oxide?
   - (*b*) Give the steps for the transformation of copper into copper sulphate and cuprous oxide?

## PROBLEMS. XXXI.

1. When 1.3305 gm. of copper were heated in oxygen, 1.6675 gm. of cupric oxide were formed. Calculate the atomic weight of copper. (The equation is $Cu + O = CuO$.)

2. How much cupric oxide is formed by heating 1467 gm. of copper in air?

3. How much cuprous oxide is formed by boiling 100 gm. of copper sulphate with potassium hydroxide and grape sugar? (The equation is $2\,CuSO_4 \cdot 5\,H_2O + 4\,KOH = Cu_2O + O + 12\,H_2O + 2\,K_2SO_4$.)

4. Calculate the percentage composition of —
   (*a*) Malachite, $CuCO_3 . Cu(OH)_2$.
   (*b*) Azurite, $2 CuCO_3 . Cu(OH)_2$.

**Silver** is familiar as the alloy of silver in coins. Two of its compounds, silver nitrate and silver chloride, have already been considered in testing for chlorine in the form of chlorides.

### PREPARATION OF SILVER.

**Experiment 154.** — Prepare silver by one or both of the following methods: —

(*a*) Fill a small dish (*e.g.* a porcelain crucible) half full of silver nitrate solution, and add a few drops of mercury. Allow the whole to stand undisturbed for a day or more, and then examine. The delicate crystals attached to the mercury are silver. Pick them out with the forceps, wash well with water, and preserve them.

(*b*) Dissolve a ten-cent piece in 10 cc. of concentrated nitric acid, dilute with a little water, and add hydrochloric acid until the precipitation is complete. Let the precipitate settle, and wash several times by decantation. Then filter and wash until the filtrate is neutral. If convenient, let the precipitate dry; if not, scrape half from the opened paper with a knife or spatula, put it in a shallow dish, cover with dilute sulphuric acid, and add a piece of zinc; put the other half in a cavity at the end of a piece of charcoal, cover with sodium carbonate, and reduce it with a blowpipe flame. In the first case, the silver will collect as a grayish powder; remove any excess of zinc, filter, wash with water and dry the residue. It may be preserved as a powder, or fused into a bead with a blowpipe flame. In the second case, minute globules of silver will appear on the charcoal; scrape them together and fuse into a single bead.

### PROPERTIES OF SILVER.

**Experiment 155.** — Examine the silver formed in Exp. 154, and state briefly its most obvious properties.

### TEST FOR SILVER.

**Experiment 156.** — Devise a test for combined silver, based upon previous experiments. Verify it.

### LABORATORY EXERCISE. XXVIII.

1. What caused the blue color of the filtrate in Exp. 154 (*b*)?

2. Name the haloid compounds of silver. State their resemblances and differences.

3. Explain the chemical change in the formation of silver in the first case in Exp. 154 (*b*). Does it differ fundamentally from the change in the second case?

4. How may silver sulphide be formed? What is its color? What happens chemically when silver "tarnishes"?

### CLASS-ROOM EXERCISE. LII.

1. Additional study of silver.
   (*a*) Occurrence of native and combined silver. What is argentiferous lead?
   (*b*) The various industrial methods of preparation.
   (*c*) Properties.
   (*d*) Uses.
   (*e*) History.
   (*f*) Derivation of the word *argentum*.
2. Halides of silver.
   (*a*) Occurrence, properties, and uses of silver chloride.
   (*b*) Preparation and uses of silver bromide and iodide.
   (*c*) Properties common to this group.
3. Silver nitrate.
   (*a*) Names.
   (*b*) Preparation.
   (*c*) Properties.
   (*d*) Uses.
4. Miscellaneous.
   (*a*) Describe the process of silver plating.
   (*b*) What is silver glance? Horn silver? Lunar caustic?
   (*c*) Compare copper and silver in the light of the Periodic Law.
   (*d*) In what respect does gold resemble copper and silver?
5. Alloys of silver.
   (*a*) Composition of standard silver coin.
   (*b*) What is sterling silver?
   (*c*) Why are coins stamped and not cast?

### PROBLEMS. XXXII.

1. Calculate the percentage composition of —

(*a*) Silver chloride, AgCl. (*b*) Silver sulphide, $Ag_2S$.

2. If 53.1958 gm. of silver form 92.6042 gm. of silver bromide, AgBr, what is the atomic weight of silver ?

3. If 1 gm. of silver chloride yields 0.7526 gm. of silver, what is the atomic weight of silver ?

4. If 480 gm. of silver interact with nitric acid, how much silver nitrate is formed ?

**Magnesium.** — The conduct of magnesium has been studied when the element was heated in air, in oxygen, and in ammonia, and when it was treated with hydrochloric acid.

### LABORATORY EXERCISE. XXIX.

1. What happens to magnesium when it is heated in air, in oxygen, and in ammonia gas ?

2. What is the equivalent of magnesium ? The atomic weight ? Summarize the experimental evidence on which the answers are based.

### GENERAL PROPERTIES OF MAGNESIUM.

**Experiment 157.** — (*a*) Examine a piece of magnesium, and write a short account of its most obvious physical properties.

(*b*) If a sufficient quantity is available, determine its specific gravity. Weigh it first in air, then in water, and compute the specific gravity from these data. Compare the result with the specific gravity of other metals.

**Magnesium Oxide** is the only important compound of magnesium previously considered.

### LABORATORY EXERCISE. XXX.

1. When magnesium is burned in air, what compound is formed ? What is its formula ?

2. What is magnesia ? State its observed properties. Predict an important application of magnesia.

## TESTS FOR MAGNESIUM.

**Experiment 158.** — (*a*) To a solution of magnesium sulphate or chloride add successively solutions of ammonium chloride, ammonium hydroxide, and disodium phosphate. A precipitate of ammonium magnesium phosphate is formed, $NH_4MgPO_4$. It is voluminous at first, but finally crystalline. It is soluble in acids. Try it.

(*b*) Put a little powdered magnesium carbonate or oxide in a cavity at the end of a piece of charcoal, moisten with water, and heat intensely in a blowpipe flame. Cool, and moisten with a drop of cobaltous nitrate solution. Heat again, and when cool observe the color. If the experiment has been conducted properly, a pink or pale flesh-colored residue coats the charcoal.

## CLASS-ROOM EXERCISE. LIII.

1. Additional study of magnesium.
    (*a*) Occurrence of its compounds.
    (*b*) Preparation.
    (*c*) Additional properties.
    (*d*) Uses.
    (*e*) How was magnesium utilized in the discovery of argon?

2. Magnesium oxide.
    (*a*) Names.
    (*b*) Preparation.
    (*c*) Properties.
    (*d*) Uses.
    (*e*) What is magnesia mixture? Calcined magnesia? Fluid magnesia? Magnesium citrate?

3. Magnesium sulphate.
    (*a*) Names.
    (*b*) Preparation.
    (*c*) Properties.
    (*d*) Use.

4. Miscellaneous.
    (*a*) Chemical name and formula of *magnesia alba*, magnesite, brucite, periclase, dolomite.
    (*b*) Valence of magnesium in $MgO$, $Mg_2CO_3$, $MgCl_2$, $Mg(OH)_2$.
    (*c*) Review "The Stassfurt Deposits" with reference to magnesium compounds.

5. Black's work on Magnesia alba.

### PROBLEMS. XXXIII.

1. What volume of nitrogen at 20° C. and 780 mm. will combine with 50 gm. of magnesium? (Equation is $3\ Mg + N_2 = Mg_3N_2$.)

2. How many grams of magnesium will be formed by heating 100 gm. of potassium with magnesium chloride? (Equation is $K_2 + MgCl_2 = Mg + 2\ KCl$.)

3. How much sodium is necessary to decompose a kilogram of magnesium chloride, and how much magnesium will be formed? (Equation is similar to 2.)

4. How much magnesium can be obtained from 100 gm. of (*a*) magnesite, $MgCO_3$, (*b*) dolomite, $MgCa(CO_3)_2$, (*c*) Epsom salts, $MgSO_4 \cdot 7\ H_2O$?

**Calcium** is an uncommon metallic element, but its compounds are exceedingly widespread and important.

### TESTS FOR CALCIUM.

**Experiment 159.** — (*a*) Subject calcium chloride to the flame test. Record the result.

(*b*) Repeat Exp. 29 (*d*).

(*c*) If possible, examine a calcium flame through a spectroscope. Compare the result with the position of the potassium lines.

**The Principal Compounds of Calcium** are the hydroxide, oxide, carbonate, chloride, sulphate, fluoride, and hypochlorite (bleaching powder).

### LABORATORY EXERCISE. XXXI.

1. What experimental use has been made of calcium hydroxide, calcium chloride, calcium fluoride, and calcium hypochlorite?

2. What happens to calcium oxide and calcium chloride when exposed to the air? What thermal effect is produced by dissolving each in water?

3. Give the common name of each of the above compounds (except calcium chloride).

4. How is lime water prepared? How should it be kept?

5. What happens to crystallized calcium sulphate (gypsum) when heated? To calcium carbonate when treated with acids?

6. Examine specimens of selenite, gypsum, and plaster of Paris. Describe them. Recall or repeat the experimental study of the solubility of calcium sulphate. Mix a little plaster of Paris with enough water on a block of wood to form a thin paste. Let it stand undisturbed for ten or fifteen minutes, and then examine. Describe the change. What is meant by the "set" of plaster of Paris?

7. Starting with limestone, how would you prepare lime, lime water, milk of lime, and chloride of lime?

8. Calcium sulphate is nearly insoluble in water; how can it be proved experimentally that it is a sulphate?

### CLASS-ROOM EXERCISE. LIV.

1. Additional study of calcium.
    - (*a*) Occurrence of combined calcium.
    - (*b*) Discovery.
    - (*c*) Properties.
2. Calcium oxide.
    - (*a*) Names.
    - (*b*) Industrial preparation.
    - (*c*) Properties.
    - (*d*) Uses.
    - (*e*) Explain and illustrate the statement, "the term *lime* is ambiguous."
3. Calcium hydroxide.
    - (*a*) Names.
    - (*b*) Properties of actual substance.
    - (*c*) Uses.
    - (*d*) What is "milk of lime"? Mortar? Cement? Plaster? Slaked lime? Explain the hardening of mortar, cement, and plaster.
4. Calcium carbonate.
    - (*a*) Names.
    - (*b*) Occurrence.
    - (*c*) Crystallization.
    - (*d*) Properties.
    - (*e*) Uses of various varieties.
    - (*f*) What are stalactites and stalagmites? How are they formed?
5. Calcium chloride.
    - (*a*) Natural sources.
    - (*b*) Preparation.
    - (*c*) Properties.
    - (*d*) Uses.
6. Calcium fluoride.
    - (*a*) Occurrence.
    - (*b*) Crystallization
    - (*c*) Properties.
    - (*d*) Uses.

7. Calcium sulphate.
    - (*a*) Names.
    - (*b*) Occurrence.
    - (*c*) Properties.
    - (*d*) Uses.
    - (*e*) Plaster of Paris — source, properties, and uses. Why so called?

8. Calcium hypochlorite.
    - (*a*) Review the subject as considered under chlorine
    - (*b*) Formula.
    - (*c*) What is "bleach"? "Chloride of lime"?

9. Calcium phosphates.
    - (*a*) Review the subject as considered under phosphorus.

10. Miscellaneous.
    - (*a*) Elements related to magnesium and calcium.
    - (*b*) Review "Hardness of water." (See CLASS-ROOM EXERCISE IX, 4 (*c*) and page 248.)
    - (*c*) What is the chemical name of Iceland spar? Of satin spar? Of selenite?
    - (*d*) What is the "lime" light?
    - (*e*) Valence of calcium in $CaO$, $CaCl_2$, $CaCO_3$, $Ca(OH)_2$, $CaSO_4$.
    - (*f*) Discuss magnesium and calcium in the light of the Periodic Law.

### PROBLEMS. XXXIV.

1. How much calcium may be obtained from 100 gm. of (*a*) marble, (*b*) gypsum, $CaSO_4 \cdot 2\,H_2O$, (*c*) fluor spar, $CaF_2$, (*d*) superphosphate of lime, $CaH_4(PO_4)_2$?

2. How many tons of limestone must be heated to form 100 tons of quicklime? The equation for the reaction is $CaCO_3 = CO_2 + CaO$.

3. Find the simplest formula of a compound having the percentage composition $Ca = 40$, $C = 12$, $O = 48$.

4. If 100 gm. of calcium carbonate yield 56 gm. of calcium oxide, what is the atomic weight of calcium?

5. How many grams does a pound of selenite lose when heated to constant weight? The equation is $CaSO_4 \cdot 2\,H_2O = CaSO_4 + 2\,H_2O$.

**Zinc** is a familiar metal, and the compounds, zinc sulphide and zinc sulphate, have been experimentally studied.

### GENERAL PROPERTIES OF ZINC.

**Experiment 160.** — (*a*) Examine a piece of zinc and record its most obvious properties.

(*b*) Determine the specific gravity by the method previously described. (See Exp. 150 (*b*).) Compare the result with the specific gravity of other metals.

### LABORATORY EXERCISE. XXXII.

1. What happens to zinc when heated? Describe and name the product. When it is exposed to air?

2. Is zinc hard or soft? Malleable? Ductile? Brittle? Tough? Does it melt easily?

### TESTS FOR ZINC.

**Experiment 161.** — (*a*) Recall or devise a simple test for combined zinc.

(*b*) Recall or repeat the action of zinc when heated in the oxidizing flame.

(*c*) Fill a small cavity at one end of a piece of charcoal with zinc oxide, moisten with water, and heat strongly in the blowpipe flame. Cool, and moisten with a drop of cobaltous nitrate solution, then heat again. Cool and examine. A green incrustation is caused by zinc compounds.

**Zinc interacts** with other metals similarly to copper.

### INTERACTION OF METALS — ZINC.

**Experiment 162.** — (*a*) Put a strip of clean zinc in copper sulphate solution, as in Exp. 152. Describe and explain the result which soon occurs.

(*b*) Put a strip of clean zinc in lead nitrate or lead acetate solution, and describe and explain the result.

(*c*) Clean a strip of zinc and put it in a solution of any salt of mercury. Examine after a short time. What is amalgamated zinc, and for what is it used?

### LABORATORY EXERCISE. XXXIII.

1. What happens when zinc is added to sulphuric acid? Describe the purified solid product. Illustrate by an equation the chemical change.

2. What happens when a little sodium hydroxide solution is added to a zinc compound (in solution)? When an excess is added? Name all the compounds of zinc formed in these changes.

3. What happens when a sulphide, *e.g.* hydrogen sulphide or ammonium sulphide, is added to a zinc compound (in solution)? What characteristic property has the solid product? If the zinc solution were acid (instead of neutral), how would the result be changed?

4. How may zinc be transformed into zinc nitrate and then into zinc oxide?

### CLASS-ROOM EXERCISE. LV.

1. Additional study of zinc.
   - (*a*) Occurrence of combined zinc.
   - (*b*) Industrial preparation.
   - (*c*) Properties.
   - (*d*) Uses.
2. Compounds.
   - (*a*) Chemical name and formula of calamine, zinc blende, gahnite, zinc white, white vitriol, sphalerite.
3. Alloys.
   - (*a*) What alloys contain zinc?
4. Miscellaneous.
   - (*a*) What is galvanized iron?
   - (*b*) Valence of zinc in $ZnSO_4$, $ZnCl_2$, $ZnO$, $ZnS$.

### PROBLEMS. XXXV.

1. Review Problems 1, 2, 4, and 6 in Chapter III., page 52.
2. Review Problems 2, 3, 4, 11, 12, and 15 in Chapter IV., page 62.
3. Review Problem 3 in Chapter IX., page 154.

**Mercury** is the only metallic element which is liquid at the ordinary temperature, and it is more or less familiar from its extensive use in thermometers and barometers.

### GENERAL PROPERTIES OF MERCURY.

*Mercury and its compounds are poisonous.*

**Experiment 163.** — (*a*) Examine some mercury, and state its characteristic physical properties. Pour a drop or two into a large shallow dish, and agitate the dish; does the result suggest the reason for the name "quicksilver"?

(*b*) Lift a bottle of mercury. Estimate its specific gravity. Verify the estimate by consulting a book, or by weighing a measured small quantity, as in the case of sulphuric acid (Exp. 4, Lab. Ex. XXIII., page 271).

**Mercury is prepared** from its compounds by roasting them alone, or with sodium carbonate or lime. Cinnabar, mercuric sulphide, is the natural compound used in the industrial preparation of mercury.

### PREPARATION OF MERCURY.

*Perform in the Hood.*

**Experiment 164.** — (*a*) Recall or repeat the preparation of mercury from mercuric oxide.

(*b*) Put a little vermilion or powdered cinnabar near one end of a glass tube open at both ends, and heat slowly at first, but finally strongly where the powder is located. The tube should be held at a slight angle to the flame and rotated. Fumes of sulphur dioxide (and possibly of mercury) will escape from the upper end of the tube, and mercury will be deposited in the colder portion somewhat as in (*a*).

(*c*) Mix any dry salt of mercury, *e.g.* mercuric chloride, with three times its weight of dry sodium carbonate, and heat strongly in an ignition tube. A test tube may be used, but it is liable to crack or melt. Mercury will be deposited on the upper part of the tube.

**Definitions.** — The layer of mercury obtained in Exp. 164 is a **sublimate.** The operation is called **sublimation**, or **subliming.** The preparation of iodine illustrated sublimation. A sublimate is usually purer than its source.

### TESTS FOR MERCURY.

**Experiment 165.** — (*a*) What is the simplest test for free mercury?
(*b*) Recall or devise a test for combined mercury. Verify it.

**Mercuric Oxide** is the only compound of mercury previously studied.

### LABORATORY EXERCISE. XXXIV.

1. Describe briefly but accurately the effect of heat on mercuric oxide. Write the equation for the change.
2. What historical interest has this compound ?

**Other Important Compounds of Mercury** belong to the two series analogous to the copper compounds previously described, viz., mercurous and mercuric nitrate and chloride.

### PROPERTIES OF MERCUROUS AND MERCURIC COMPOUNDS.

**Experiment 166.** — (*a*) *Mercurous.* Add a few drops of hydrochloric acid to a little mercurous nitrate solution. The white precipitate is mercurous chloride. Note its insolubility in water and in dilute hydrochloric acid. Add a few drops of ammonium hydroxide. The black precipitate is mainly mercurous ammonium chloride. Its formation is a delicate test for mercury in mercurous compounds.

(*b*) *Mercuric.* Add a few drops of hydrochloric acid to a little mercuric nitrate solution. Compare the result with that in (*a*). Add a few drops of ammonium hydroxide, or enough to produce a decided change. Compare with (*a*). The precipitate is mercuric ammonium chloride.

The formula of mercurous nitrate is $Hg_2(NO_3)_2$, and of mercurous chloride is $Hg_2Cl_2$; some authorities write these formulas $HgNO_3$ and HgCl. The formula of mercuric nitrate is $Hg(NO_3)_2$, and of mercuric chloride is $HgCl_2$. These compounds are analogous to the corresponding *-ous* and *-ic* oxides. Thus,

| | |
|---|---|
| Mercuric oxide is . . . . . . | HgO |
| Mercuric chloride is . . . . . | $HgCl_2$ |
| Mercuric nitrate is . . . . . | $Hg(NO_3)_2$ |
| Mercurous oxide is . . . . . | $Hg_2O$ |
| Mercurous chloride is . . . . | $Hg_2Cl_2$ |
| Mercurous nitrate is . . . . | $Hg_2(NO_3)_2$. |

## CLASS-ROOM EXERCISE. LVI.

1. Additional study of mercury.
   - (*a*) Occurrence of free and combined mercury.
   - (*b*) Industrial preparation.
   - (*c*) Impurities, and how removed.
   - (*d*) Properties.
   - (*e*) Uses.
   - (*f*) Significance of the common name.
   - (*g*) History.
2. Mercuric oxide.
   - (*a*) Preparation of the varieties.
   - (*b*) Names.
   - (*c*) Historical significance.
3. Mercurous compounds.
   - (*a*) Preparation, properties, uses, formulas, and common name of mercurous chloride.
   - (*b*) Other mercurous compounds.
   - (*c*) Valence of mercury in mercurous compounds.
4. Mercuric compounds.
   - (*a*) Preparation, properties, uses, formula, and common name of mercuric chloride.
   - (*b*) Other mercuric compounds.
   - (*c*) Valence of mercury in mercuric compounds.
5. Miscellaneous.
   - (*a*) Explain the statement, "the vapor of mercury and of zinc consists of monatomic molecules."
   - (*b*) Discuss the relation of zinc to mercury, and of both metals to magnesium and calcium in the light of the Periodic Law.
   - (*c*) Literal meaning of word "hydrargyrum."
6. Amalgams.
   - (*a*) Definition of an amalgam.
   - (*b*) Preparation. properties, and uses of sodium amalgam.
   - (*c*) What is amalgamated zinc, and for what is it used ?
   - (*d*) For what do dentists use amalgams ?
   - (*e*) How is gold often extracted from its ores ?

## PROBLEMS. XXXVI.

1. If 195 cc. of oxygen at 37° C. and 620 mm. are obtained by heating 2.7 gm. of mercuric oxide, what is the atomic weight of mercury ?

2. How much mercury is formed by heating 400 gm. of pure cinnabar ? (Equation is $HgS + O_2 = Hg + SO_2$.)

3. If 70 gm. of mercuric oxide are heated, what volume of oxygen at 21° C. and 740 mm. is evolved ?

4. If 177.1664 gm. of mercuric sulphide yield 152.745 gm. of mercury, what is the atomic weight of mercury ?

# CHAPTER XIX.

## ALUMINIUM — TIN — LEAD — CHROMIUM — MANGANESE — IRON — SEPARATION OF METALS.

**Aluminium,** owing to its extensive preparation by electrolytic methods, is a familiar metal. Alum has long been known.

### GENERAL PROPERTIES OF ALUMINIUM.

**Experiment 167.** — (*a*) Examine a piece of aluminium (sheet or wire), and observe its physical properties. Has it any "spring" like brass? Is it ductile, malleable, soft, hard, tough, brittle?

(*b*) Compare roughly the weight of a piece of aluminium with a piece of zinc, pasteboard, and glass having approximately the same volume.

If sufficient aluminium is available, determine the specific gravity. (See Exp. 150 (*b*).) Compare the result with the specific gravity of other metals.

It has been asserted that aluminium is not acted upon by acids and alkalies.

### ACTION OF ALUMINIUM WITH ACIDS AND ALKALIES.

**Experiment 168.** — (*a*) Add a small piece of aluminium to separate test tubes containing dilute sulphuric acid and dilute hydrochloric acid. Warm, if necessary. Describe the action. Test the gas evolved. What compound is formed in each case?

(*b*) Add a small piece of aluminium to a test tube half full of a dilute solution of sodium hydroxide, and boil. Test any gas evolved. If only a little gas is liberated, attach a simple delivery tube and collect the gas over water.

Other acids and alkalies act similarly; draw a general conclusion from this experiment. Is the aluminium examined in Exp. 167 tarnished? Scrape the surface, and examine again. Does aluminium differ from most metals in this respect?

**Aluminium Hydroxide** is an important compound of aluminium.

PREPARATION AND PROPERTIES OF ALUMINIUM HYDROXIDE.

**Experiment 169.** — (*a*) Add slowly a little potassium hydroxide or sodium hydroxide solution to a test tube half full of alum solution. The gelatinous precipitate is aluminium hydroxide. Now add an excess of the alkali to one half, and dilute hydrochloric acid to the other. Describe the results.

(*b*) Recall or devise another simple method of preparing aluminium hydroxide.

(*c*) Add a little solution of ammonium sulphide to a solution of alum. The precipitate is not a sulphide, but aluminium hydroxide, because aluminium forms no sulphide in the wet way.

(*d*) Add a little alum solution to a dilute solution of cochineal, then add ammonium hydroxide. The colored product is called carmine lake. It belongs to a class of dyes formed by the combination of a vegetable dye and a metallic hydroxide, usually aluminium hydroxide.

**Discussion of Experiment 169.** — When sodium hydroxide or potassium hydroxide is first added to a solution of a compound of aluminium, aluminium hydroxide is formed. The simplest equation for this reaction is —

$$\underset{\text{Aluminium Sulphate}}{Al_2(SO_4)_3} + 6\,KOH = 2\,\underset{\text{Aluminium Hydroxide}}{Al(OH)_3} + 3\,K_2SO_4.$$

An excess of alkali produces a chemical change represented by the equation: —

$$Al(OH)_3 + 3\,KOH = \underset{\text{Potassium Aluminate}}{Al(OK)_3} + 3\,H_2O$$

The chemical changes are similar to the changes of zinc compounds under the same conditions. (See Exp. 21.)

### TESTS FOR ALUMINIUM.

**Experiment 170.** — (*a*) What is a simple test for metallic aluminium?

(*b*) Recall or devise a test for combined aluminium. Verify it. How can aluminium compounds be distinguished from those of zinc?

(*c*) Heat a little aluminium sulphate or aluminium hydroxide on charcoal in the blowpipe flame. Cool and moisten with a drop of cobaltous nitrate solution. Heat again, and if the operation has been conducted properly, a blue residue will coat the charcoal. This color is characteristic of aluminium compounds. Compare this result with the action of other metallic compounds under similar circumstances.

**Alum** is by far the most useful compound of aluminium.

### PREPARATION AND PROPERTIES OF COMMON ALUM.

**Experiment 171.** — (*a*) Dissolve about 10 gm. of aluminium sulphate in the least possible amount of hot water. Dissolve 3 gm. of potassium sulphate in the same way. Mix the clear, hot, saturated solutions in a small shallow dish, and allow the solution to cool undisturbed. Crystals of potassium alum will be deposited. Remove the best ones; dry, and examine. Describe them, giving color, luster, size, and crystal form.

(*b*) Prove by actual tests that (1) they are a sulphate of potassium and of aluminium, and that (2) they contain water of crystallization.

### CLASS-ROOM EXERCISE. LVII.

1. Additional study of aluminium.
    - (*a*) Occurrence.
    - (*b*) Proportion in the earth's crust.
    - (*c*) Industrial preparation.
    - (*d*) Properties.
    - (*e*) Uses.
    - (*f*) History.
    - (*g*) Derivation of the words *aluminum* and *aluminium*.
2. Alums.
    - (*a*) General formula.
    - (*b*) Varieties.
    - (*c*) Properties of alums.
    - (*d*) Potassium alum.
    - (*e*) Uses.

3. Compounds.
   (*a*) Chemical name and formula of corundum, ruby, sapphire, cryolite, turquoise, bauxite, alumina, feldspar.

4. Miscellaneous.
   (*a*) Valence of aluminium in $Al_2O_3$ and $Al_2(SO_4)_3$.
   (*b*) What is burnt alum, a mordant, red liquor?
   (*c*) What elements are closely related to aluminium?
   (*d*) What is aluminium bronze? Uses.

5. Complete the following equations (note that one side is already complete): —
   (*a*) $Al + O = Al_2O_3$
   (*b*) $Al + H_2SO_4 = Al_2(SO_4)_3 + 3\,H_2$
   (*c*) $Al_2(SO_4)_3 + KOH = 2\,Al(OH)_3 + 3\,K_2SO_4$

### PROBLEMS. XXXVII.

1. How much aluminium can be obtained theoretically from 100 gm. of —
   (*a*) $Al_2O_3 . 3\,SiO_2$?
   (*b*) Cryolite, $AlNa_3F_6$?
   (*c*) Turquoise, $Al_2P_2O_8 . H_6Al_2O_6 . 2\,H_2O$?

2. How much aluminium can be obtained by the interaction of a kilogram of sodium aluminium chloride and sodium? (Equation is $AlCl_3 . NaCl + 3\,Na = Al + 4\,NaCl$.)

3. What volume of oxygen at 15° C. is needed to burn 5 gm. of aluminium to $Al_2O_3$?

4. Half a gram of aluminium when warmed with potassium hydroxide, liberated 660 cc. of hydrogen at 13° C. (over water). Calculate the atomic weight of aluminium. (Equation is $Al_2 + 2\,KOH + 2\,H_2O = 2\,KAlO_2 + 3\,H_2$.)

5. Review Problems 5, 6, 7, 8, 9 in Chapter IX, page 154.

**Tin** is a familiar metal owing to its extensive use as a protective coating for iron. Its compounds are not numerous, though one, stannous chloride, is widely used in dyeing.

### GENERAL PROPERTIES OF TIN.

**Experiment 172.** — (*a*) Examine a stick of tin or a piece of pure tinfoil, and state the most obvious physical properties. Bend a stick of tin, and note the crackling sound.

(*b*) If a stick of tin or a piece of block tin pipe is available, determine the specific gravity by either of the methods described in Exp. 150 (*b*). Compare the result with the specific gravity of other metals.

### ACTION OF TIN WITH ACIDS.

**Experiment 173.** — (*a*) Put a small piece of tin — about .5 gm. — in a test tube, cover with concentrated hydrochloric acid, add a little water, and heat — *in the hood.* Heat gently at first, and when action begins regulate the heat accordingly. Most of the tin disappears, soluble stannous chloride being formed. Save this solution for Exp. 174 (*c*).

(*b*) Treat a small piece of tin with concentrated nitric acid — *in the hood.* It is advisable to stand the test tube in the rack or in a bottle as soon as the action begins. The white, amorphous product is metastannic acid. How does the action of nitric acid on tin differ from and resemble its action on other metals, zinc, for example?

Stannous chloride has the formula $SnCl_2$, but the formula of metastannic acid varies with the temperature at which it is dried.

### TESTS FOR TIN.

**Experiment 174.** — (*a*) What is a simple test for metallic tin?

(*b*) Recall or repeat the action of tin when heated in a blowpipe flame.

(*c*) Add a few drops of mercuric chloride solution (*poison*) to a small portion of the stannous chloride solution prepared in Exp. 173 (*a*). The white precipitate is mercurous chloride. Add a little more stannous chloride solution and heat gently. The mercurous chloride is reduced finally to mercury, which appears as a grayish powder.

**Discussion of Experiment 174.** — Stannous compounds readily form stannic compounds, *i.e.* by an extension of the conception of oxidation and reduction, they reduce other compounds and become oxidized, though no oxygen may be involved. Thus the simplest equation for the interaction of stannous and mercuric chlorides is —

| $2\,HgCl_2$ | + | $SnCl_2$ | = | $Hg_2Cl_2$ | + | $SnCl_4$ |
|---|---|---|---|---|---|---|
| Mercuric Chloride | | Stannous Chloride | | Mercurous Chloride | | Stannic Chloride |

The mercuric chloride is reduced to mercurous chloride, while the stannous chloride is oxidized to stannic chloride. Reduction in its broadest sense means "going lower," and oxidation "going higher" in a series of compounds. Stannous chloride is often used as a reducing agent.

### DEPOSITION OF METALLIC TIN.

**Experiment 175.** — Put a strip of zinc in a slightly acid solution of stannous chloride. Examine after a short time, and the tin will be found adhering to the zinc as a grayish black deposit; sometimes it appears as minute scales. What becomes of the zinc?

### CLASS-ROOM EXERCISE. LVIII.

1. Additional study of tin.
   (*a*) Occurrence of combined tin.
   (*b*) Industrial preparation.
   (*c*) Properties.
   (*d*) Uses.
   (*e*) History.
2. Compounds.
   (*a*) Chemical name and formula of cassiterite, tin crystals, stream tin, tin dioxide, oxymuriate of tin.
3. Miscellaneous.
   (*a*) What is tin plate? "Tin"? Block tin?
   (*b*) Valence of tin in $SnCl_2$, $SnCl_4$, $SnO_2$, $SnS$.
4. Name the important alloys of tin.
5. Complete the equations: —

(*a*) $Sn + \quad = SnO_2$.

(*b*) $SnCl_2 + \quad = SnCl_4$.

(*c*) $SnCl_4 + \quad = SnCl_2 + 2\,HCl$.

### PROBLEMS. XXXVIII.

1. Calculate the percentage composition of—
   - (*a*) Tinstone, $SnO_2$.
   - (*b*) Stannous chloride, $SnCl_2$.
   - (*c*) Stannic chloride, $SnCl_4$.

2. If 100 gm. of tin are heated with nitric acid, and the stannic oxide formed on heating the product weighs 127.1 gm., what is the atomic weight of tin?

**Lead**, both free and combined, is well known, and has numerous applications. Its principal compounds are the oxides, sulphide, sulphate, chromate, nitrate, and carbonate.

### GENERAL PROPERTIES OF LEAD.

**Experiment 176.** — (*a*) Examine a piece of freshly cut lead and state its most obvious physical properties.

(*b*) Determine its specific gravity by either method described in Exp. 150 (*b*). Compare the result with the specific gravity of other metals.

### LABORATORY EXERCISE. XXXV.

1. What happens to lead when heated? When exposed to the air? Is it easily melted and tarnished?

2. What physical properties adapt it for its extensive use?

3. Draw a piece of lead across a sheet of white paper and describe the result. What is erroneously called "black lead"? Is there any lead in a lead pencil?

### TESTS FOR LEAD.

**Experiment 177.** — (*a*) Recall or repeat the reduction of lead oxide in the blowpipe flame.

(*b*) Recall or repeat the action of hydrogen sulphide with the solution of any lead compound.

(*c*) Add dilute hydrochloric acid to a little lead nitrate solution until precipitation ceases. Note the insolubility of the lead chloride which is formed. Warm gently as long as any decided change occurs. Describe the action. This is characteristic of lead chloride and permits its separation from the chloride of silver and of mercury (in the *-ous* condition).

(*d*) Add dilute sulphuric acid to a little lead nitrate solution until precipitation ceases. The precipitate is lead sulphate. Observe its properties. Is it soluble in hot water? Try it.

(*e*) Repeat (*d*), using potassium chromate or dichromate instead of sulphuric acid. The precipitate is lead chromate. Describe it, especially the color.

**Discussion of Experiment 177.** — The test in (*c*) is usually employed as the preliminary test, and it is confirmed by (*b*), (*d*), and (*e*), or any of these three.

Lead, like other metals, interacts with zinc.

### INTERACTION OF METALS — LEAD.

**Experiment 178.** — Repeat Exp. 162 (*b*), and describe the result.

**Soluble Lead Compounds** are often formed when lead is subjected to the action of water containing much gas, *e.g.* air and carbon dioxide.

### ACTION OF WATER ON LEAD.

**Experiment 179.** — Fill a bottle half full of water, pass carbon dioxide into it from the lungs, and shake vigorously. Put a piece of clean lead in the water, cork the bottle, and let the whole stand undisturbed for a day or two. Remove the lead and test separate small portions of the water for lead by Exp. 177 (*c*), (*d*), (*e*). If no lead is detected, let the action continue another day, or evaporate the liquid to a small bulk and test as above.

Ordinarily, drinking water has no action on lead, but if the solvent power is increased by the presence of an excess of gases or of organic matter containing sodium or potassium nitrate, enough lead compound may be taken up by the water to cause lead poisoning when the water is drunk.

**Oxides of Lead.** — There are several oxides of lead; the most important are lead monoxide, $PbO$, lead dioxide, $PbO_2$, and lead tetroxide, $Pb_3O_4$.

## PROPERTIES OF THE OXIDES OF LEAD.

**Experiment 180.** — (*a*) Examine the three oxides and tabulate their most obvious physical properties, stating the exact chemical name and formula and the popular name of each oxide.

(*b*) Repeat or recall the experiment in which lead was heated in the oxidizing flame, especially the color of the coating. What oxide of lead is thereby formed?

(*c*) Warm a little lead tetroxide with dilute nitric acid. The solid product is lead dioxide. Describe it.

## LABORATORY EXERCISE. XXXVI.

1. How might lead tetroxide be prepared?

2. If lead tetroxide is heated strongly, lead monoxide is formed. What does this fact reveal about the stability of lead tetroxide?

3. When lead dioxide and concentrated hydrochloric acid are mixed and heated, chlorine is evolved. Complete the equation —

$$PbO_2 \quad + \quad HCl \quad = \quad PbCl_2 \quad + \quad 2\,H_2O \quad +$$

How does this interaction resemble that of manganese dioxide and hydrochloric acid?

4. How may lead nitrate be formed? What happens when it is heated strongly? Prove that it is a nitrate and a lead compound.

5. Examine lead carbonate and state its most obvious properties. Prove that it is a carbonate and contains lead.

6. Examine a lump of galena and state its most obvious properties. Prove that it is lead sulphide. Complete the equation: —

$$PbS \quad + \quad O \quad = \qquad\quad + \quad SO_2.$$

## CLASS-ROOM EXERCISE. LIX.

1. Additional study of lead.

   (*a*) Occurrence.
   (*b*) Metallurgy.
   (*c*) Properties.
   (*d*) Uses.
   (*e*) History.

2. Oxides of lead. Give the preparation, properties, and uses of —

   (*a*) Lead monoxide.
   (*b*) Lead dioxide.
   (*c*) Lead tetroxide.

3. Lead carbonate.
   (*a*) Occurrence.
   (*b*) Industrial preparation — old and new methods.
   (*c*) Properties. (*d*) Uses.
4. Lead sulphide.
   (*a*) Properties of galena.
   (*b*) Properties of the artificial sulphide.
5. Compounds of lead.

Give the chemical name and formula of galena, cerussite, anglesite, sugar of lead, white lead, red lead, litharge, lead peroxide, chrome yellow.

6. Miscellaneous.
   (*a*) Valence of lead in $PbO$, $PbO_2$, $PbS$, $PbCO_3$, $PbCl_2$, $PbCl_4$.
   (*b*) Illustrate the Periodic Law by the compounds of tin and lead.
7. Alloys of lead.
   (*a*) Approximate composition of hard and soft solder, and of pewter; uses of each.
   (*b*) Approximate composition of Britannia metal and of type metal.
   (*c*) Components of a fusible alloy. Uses.
   (*d*) What is shot ?

### PROBLEMS. XXXIX.

1. What per cent of lead is contained in —
   (*a*) Galena, $PbS$.
   (*b*) Cerussite, $PbCO_3$.
   (*c*) Anglesite, $PbSO_4$.
   (*d*) Lead acetate, $Pb(C_2H_3O_2)_2 \cdot 3\,H_2O$

2. How much litharge can be made from 40.5 gm. of lead ? (Equation is $Pb + O = PbO$.)

3. An analysis of lead monoxide showed that 100 gm. contained 7.1724 gm. of oxygen. Calculate the atomic weight of lead

4. If 5 gm. of lead chloride give 5.16 gm. of silver chloride, what is the atomic weight of lead ? (Equation is $PbCl_2 + 2\,AgNO_3 = 2\,AgCl + Pb(NO_3)_2$.)

5. If 100 gm. of lead form 159.9703 gm. of lead nitrate, what is the atomic weight of lead ? (Assume $Pb + 2\,HNO_3 = Pb(NO_3)_2 + H_2$.)

**Chromium** is an uncommon metal, but several of its compounds, especially potassium chromate, potassium dichromate, and chrome alum, are familiar, and have numerous industrial applications.

### TESTS FOR CHROMIUM.

**Experiment 181.** — (*a*) Prepare a borax bead (see page 231), touch it with a minute quantity of any chromium compound, *e.g.* chrome alum, and heat in both the oxidizing and reducing flame. The green color is a characteristic and delicate test for chromium.

(*b*) Mix equal small quantities of potassium carbonate, potassium nitrate, and powdered chrome alum, place the mixture on a platinum foil, and hold it with the forceps in the upper Bunsen flame so that the mixture will fuse. If a platinum foil is not available, use a porcelain crucible. A yellow mass, due to the presence of potassium chromate, results. If the color is not decided, dissolve the mass in water, add acetic acid, slowly at first, and boil to expel the carbon dioxide. Add a few drops of lead nitrate solution to a portion, and yellow lead chromate is precipitated. (Compare Exp. 177 (*e*).) If the precipitate is white, it is lead carbonate, and shows that not all the potassium carbonate was decomposed, as intended.

(*c*) Add lead nitrate solution to potassium dichromate solution. Name and describe the precipitate. Try the solubility of the precipitate in acetic acid, dilute nitric acid, and sodium hydroxide.

**Chromium in its Compounds** is sometimes metallic and sometimes non-metallic toward the other components. Thus, in chromium trioxide, $CrO_3$, the anhydride of the hypothetical chromic acid, $H_2CrO_4$, chromium acts as a non-metal, just like sulphur in sulphuric acid; hence in chromates the chromium is acidic, or non-metallic, in its chemical relations with the other elements. In chromic compounds, however, chromium acts as a metal. Thus chromium hydroxide, $Cr(OH)_3$, is analogous to aluminium hydroxide; chrome alum is potassium chromium sulphate, and is analogous to aluminium alum. The two classes pass into each other by appropriate operations.

**Chromates.**—The two important chromates are potassium chromate, $K_2CrO_4$, and potassium dichromate, $K_2Cr_2O_7$.

PROPERTIES OF CHROMATES.

**Experiment 182.**—(*a*) Examine crystals of potassium chromate and dichromate, and state their most obvious physical properties. From previously determined facts, predict their solubility or insolubility in water. Verify the prediction.

(*b*) Recall the properties of lead chromate. Complete the equation :—

$$\underset{}{K_2CrO_4} + Pb(NO_3)_2 = \underset{\substack{\text{Lead}\\\text{Chromate}}}{PbCrO_4} +$$

(*c*) Add a few drops of concentrated hydrochloric acid to a dilute solution of potassium chromate and observe the change of color. This is due to the presence of potassium dichromate, which was formed from the chromate. The equation is —

$$2\,K_2CrO_4 + 2\,HCl = \underset{\substack{\text{Potassium}\\\text{Dichromate}}}{K_2Cr_2O_7} + 2\,KCl + H_2O.$$

(*d*) Add potassium hydroxide solution to 10 cc. of potassium dichromate solution until the color just changes. The yellow color is due to the presence of potassium chromate, which was formed from the dichromate. Thus —

$$K_2Cr_2O_7 + 2\,KOH = 2\,K_2CrO_4 + H_2O.$$

(*e*) The chromates are oxidizing agents. Add a few drops of concentrated hydrochloric acid to powdered potassium chromate and dichromate in separate test tubes. Chlorine is evolved, owing to the oxidation of hydrochloric acid. The simplest equation in the case of the chromate is —

$$K_2CrO_4 + 8\,HCl = 3\,Cl + \underset{\substack{\text{Chromic}\\\text{Chloride}}}{CrCl_3} + 2\,KCl + 4\,H_2O.$$

**Definition.** — **Oxidation** sometimes means the withdrawal of hydrogen from a compound, as well as the addition of oxygen. If the hydrogen can be removed, it is ultimately oxidized to water, hence oxidation actually occurs.

**Chromic Compounds.** — The chromic chloride formed in Exp. 182 (*c*) illustrates the possibility of passing from a chromate to a chromic salt. Here the chromate was reduced. Conversely, Exp. 181 (*b*) illustrates the formation of a chromate by the oxidation of a chromic compound.

### REDUCTION OF CHROMATES TO CHROMIC COMPOUNDS.

**Experiment 183.** — Add to a few cubic centimeters of potassium dichromate solution a little concentrated hydrochloric acid and a few drops of alcohol. Warm gently. Two important changes occur. The chromate is reduced to chromic chloride which colors the solution green; the alcohol is oxidized to aldehyde, which is detected by its peculiar odor. (See Exp. 111 (*a*).) The equation is —

$$K_2Cr_2O_7 + 8\,HCl + 3\,\underset{\text{Alcohol}}{C_2H_6O} = 2\,CrCl_3 + 3\,\underset{\text{Aldehyde}}{C_2H_4O} + 2\,KCl + 7\,H_2O.$$

### PROPERTIES OF CHROME ALUM.

**Experiment 184.** — (*a*) Examine chrome alum and state its most obvious physical properties.

(*b*) Recrystallize a little and describe the crystals.

(*c*) Prove that chrome alum is a sulphate, and that it contains chromium and water of crystallization.

### PREPARATION AND PROPERTIES OF CHROMIC HYDROXIDE.

**Experiment 185.** — (*a*) Add a little sodium hydroxide solution to a solution of chrome alum. The precipitate is chromic hydroxide. Describe it. Add an excess of sodium hydroxide solution and shake. Describe the result Boil, and state the result.

(*b*) Add a little, and then an excess, of ammonium sulphide to a solution of chrome alum. Compare the result with that in (*a*). Does chromium form a sulphide? Do aluminium and zinc?

### LABORATORY EXERCISE. XXXVII.

1. Compare the action of sodium hydroxide on ordinary alum and on chrome alum.

2. How can aluminium hydroxide be distinguished from chromic hydroxide?

3. Read the discussion of Exp. 169.

### CLASS-ROOM EXERCISE. LX.

1. Additional study of chromium.
   - (*a*) Occurrence and source.
   - (*b*) Preparation.
   - (*c*) Properties.
   - (*d*) Uses.
2. Chromates.
   Give the preparation, properties, and sources of—
   - (*a*) Potassium chromate.
   - (*b*) Potassium dichromate.
3. Chromic compounds.
   - (*a*) How do these compounds differ essentially from chromates?
   - (*b*) Names and formulas of those experimentally studied.
   - (*c*) Preparation and properties of chrome alum.
4. Miscellaneous.
   - (*a*) Valence of chromium in $Cr(OH)_3$, $Cr_2(SO_4)_3$, $K_2CrO_4$, $PbCrO_4$, $K_2Cr_2O_7$.
   - (*b*) Chemical name and formula of chromite, crocoite, chrome yellow, Guignet's green, chrome green, chrome alum, yellow chromate of potash.

### PROBLEMS. XL.

1. Find the percentage composition of—
   - (*a*) Lead chromate, $PbCrO_4$.
   - (*b*) Chromic oxide, $Cr_2O_3$.
   - (*c*) Chrome ironstone, $Cr_2O_3 . FeO$.
   - (*d*) Potassium chlorochromate, $KClCrO_3$.
   - (*e*) Chromium monoxide, $CrO$.
   - (*f*) Chromium trioxide, $CrO_3$.

2. If 100 gm. of lead nitrate produce 97.576 gm. of lead chromate, what is the atomic weight of chromium? (Equation is $K_2CrO_4 + Pb(NO_3)_2 = PbCrO_4 + 2\,KNO_3$.)

3. If 100 gm. of chromous chloride, $CrCl_2$, yield 57.5 gm. of chlorine, what is the atomic weight of chromium?

4. If 36.865 gm. of chromic chloride form 100 gm. of silver chloride, what is the atomic weight of chromium? (Equation is $2\ CrCl_3 + 6\ AgNO_3 = 2\ Cr(NO_3)_3 + 6\ AgCl$.)

**Manganese,** like chromium, is a rare metal. Two of its compounds, manganese dioxide and potassium permanganate, have already been used.

### TESTS FOR MANGANESE.

**Experiment 186.** — (*a*) Subject a minute quantity of manganese dioxide to the borax bead test, and note the color of the bead after heating in each flame.

(*b*) Fuse on a platinum foil or in a crucible a little manganese dioxide mixed with potassium carbonate and potassium nitrate. (See Exp. 181 (*b*).) The green mass is a characteristic test for manganese. It is due to the presence of potassium manganate.

(*c*) Add ammonium sulphide to manganese sulphate or chloride solution. The flesh-colored precipitate is manganese sulphide. Divide it into two parts. Add hydrochloric acid to one and acetic acid to the other, then add an excess of ammonium hydroxide to each. Draw a conclusion regarding the solubility of manganese sulphide.

### LABORATORY EXERCISE. XXXVIII.

1. For what has manganese dioxide been used in the laboratory? Is it an oxidizing agent?

2. For what has potassium permanganate been used in the laboratory?

3. Describe potassium permanganate. What can be said of its solubility in water?

**Potassium Permanganate** is a powerful oxidizing agent.

### OXIDATION WITH POTASSIUM PERMANGANATE.

**Experiment 187.** — (*a*) Add a few drops of sulphuric acid to a weak solution of fresh ferrous sulphate; then add, drop by drop, a dilute solution of potassium permanganate. Its color is changed, owing to the loss of oxygen which changes the ferrous to the ferric sulphate; the decomposition of the permanganate also allows the formation of potassium and manganese sulphates.

(*b*) Pour a solution of potassium permanganate upon a piece of filter paper. Describe and explain the result.

### LABORATORY EXERCISE. XXXIX.

1. What is the formula of potassium permanganate? Does the formula give a clue to the oxidizing power?

2. Potassium permanganate solution is often used as a disinfectant. Upon what property is this application based?

### CLASS-ROOM EXERCISE. LXI.

1. Additional study of manganese.
   (*a*) Occurrence.
   (*b*) Preparation.
   (*c*) Properties.
   (*d*) Use.
2. Manganese dioxide.
   (*a*) Names.
   (*b*) Occurrence in native state.
   (*c*) Properties.
   (*d*) Uses.
3. Potassium permanganate.
   (*a*) Preparation.
   (*b*) Properties.
   (*c*) Uses.
4. Miscellaneous.
   (*a*) Valence of manganese in $MnO$, $MnO_2$, $Mn_2O_3$, $Mn_3O_4$, $KMnO_4$, $MnS$, $MnCl_2$.

### PROBLEMS. XLI.

1. If manganous sulphate, $MnSO_4$, yields 42.392 per cent of oxygen, what is the atomic weight of manganese?

2. Calculate how much manganese can be obtained from 100 gm. of—

   (*a*) Manganese dioxide, $MnO_2$.
   (*b*) Manganese trioxide, $Mn_2O_3$.
   (*c*) Manganese sulphide, $MnS$.
   (*d*) Manganese alum, $K_2Mn_2(SO_4)_4 \cdot 24\,H_2O$.

3. Calculate the weight of oxygen liberated from 100 gm. of potassium permanganate when heated with sulphuric acid. (Equation is $2\,KMnO_4 + 3\,H_2SO_4 = 5\,O + 2\,MnSO_4 + K_2SO_4 + 3\,H_2O$.)

4. How much manganese ore containing 85 per cent of manganese dioxide is needed to prepare 300 lb. of chlorine? (Equation is $MnO_2 + 4\,HCl = MnCl_2 + 2\,H_2O + Cl_2$.)

**Iron** and many of its compounds are familiar and have numerous applications.

### GENERAL PROPERTIES OF IRON.

**Experiment 188.** — (*a*) Examine cast iron, wrought iron, and steel, and state their most obvious physical properties. Try the action of a magnet on each. Drop a pinch of iron powder into the Bunsen flame. Hold a piece of fine iron wire in the Bunsen flame. Describe the results and draw conclusions.

(*b*) Determine the specific gravity of a piece of iron. Compare the result with the specific gravity of other metals.

### LABORATORY EXERCISE. XL.

1. In previous experiments iron has been (1) heated in air, (2) heated in steam, (3) treated with acids, and (4) heated with sulphur. Recall and record briefly the essential result of each experiment.

2. What is "iron by hydrogen," and "alcoholized iron" (or "iron by alcohol")?

3. What use has been made of ferrous sulphide in the laboratory.

**Ferrous and Ferric Compounds.** — Iron forms two series of compounds, the ferrous and ferric. They are analogous to cuprous and cupric compounds. The ferrous compounds in the presence of free acid pass into the corresponding ferric compound by the action of oxidizing agents, *e.g.* oxygen, nitric acid, potassium chlorate, chlorine, and bromine. Conversely the ferric compounds are reduced to the ferrous by reducing agents, *e.g.* hydrogen, hydrogen sulphide, sulphur dioxide, and stannous chloride. The passage from one state to the other occurs easily, especially from ferrous to ferric.

**Ferrous Compounds.** — The preparation, properties, and tests of ferrous compounds are shown by the

### BEHAVIOR OF FERROUS COMPOUNDS.

**Experiment 189.** — (*a*) Put a few grams (3 to 5) of iron filings in a test tube, add about 10 cc. of dilute hydrochloric acid, and warm gently. Ferrous chloride is formed (in solution). (1) Pour a little into a test tube one-third full of sodium hydroxide solution. The precipitate is ferrous hydroxide. Watch the changes in color. To what are the changes due? (2) Add a second portion to potassium ferricyanide solution. The precipitate is ferrous ferricyanide. Describe it. (3) Add a third portion to potassium thiocyanate solution. If ferric salts are absent, no change results. (4) Add a fourth portion to potassium ferrocyanide solution. The precipitate is ferrous ferrocyanide. Describe it.

The above tests, especially (2), serve to distinguish ferrous from ferric compounds, as will be readily seen by a study of the

### BEHAVIOR OF FERRIC COMPOUNDS.

**Experiment 190.** — To a little ferric chloride solution add (1) sodium hydroxide solution. The precipitate is ferric hydroxide. Describe it. Add to ferric chloride solution (2) a little solution of potassium ferricyanide. Compare the negative result with (2) in Exp. 189. Add as above (3) a little solution of potassium thiocyanate. The rich wine-red coloration is caused by the soluble ferric thiocyanate. This test distinguishes ferric from ferrous compounds. Add as above (4) a little solution of potassium ferrocyanide. The precipitate is ferric ferrocyanide. Describe it.

Tabulate the results of Exps. 189 and 190.

### REDUCTION OF FERRIC COMPOUNDS.

**Experiment 191.** – Put a piece of zinc in ferric chloride solution made slightly acid by hydrochloric acid. The nascent hydrogen reduces the ferric to ferrous chloride. After the operation has proceeded for about fifteen minutes, test a portion of the liquid for a ferrous and a ferric compound. If the tests are not conclusive, continue the reduction and test again. Describe the result.

### OXIDATION OF FERROUS COMPOUNDS.

**Experiment 192.** — (*a*) To a solution of fresh or freshly washed ferrous sulphate add a little hydrochloric acid, warm gently, and then add a few crystals of potassium chlorate. After heating a short time, test portions of the liquid for a ferric and a ferrous compound.

(*b*) Add 10 cc. of concentrated nitric acid, drop by drop, to a hot solution of ferrous sulphate to which a little sulphuric acid has been added, and boil. Test portions of the liquid for a ferric and a ferrous compound.

(*c*) Recall a third illustration of the oxidation of a ferrous to a ferric compound.

### LABORATORY EXERCISE. XLI.

1. Read the discussion of Exp. 174.

2. Examine ferrous sulphate, and state its physical properties. Test a crystal, which has been exposed to the air, for both ferric and ferrous compounds. Explain.

3. Which seems the more stable form, ferrous or ferric?

4. Examine specimens of hematite, limonite, and magnetite. Describe each. Draw the first two across a sheet of rough paper or a piece of ground glass, and describe the "streak" made by each. What is the formula of each (pure) compound? Significance of each name?

5. Examine iron pyrites and state its most obvious physical properties. It is iron disulphide. What is the formula? For what is it used?

6. Examine specimens of siderite. Describe them. Siderite is ferrous carbonate. What is the formula? How is it formed in the earth's crust? Try the action of warm hydrochloric acid on a little powdered siderite. How does iron get into clay banks and into the soil?

### CLASS-ROOM EXERCISE. LXII.

1. Additional study of iron.
   - (*a*) Occurrence of free and of combined iron.
   - (*b*) Ores of iron.
   - (*c*) Metallurgy of iron and steel.
   - (*d*) Properties of cast iron, wrought iron, and steel.
   - (*e*) Composition of the three common kinds of iron.
   - (*f*) What is passive iron?
   - (*g*) Uses of the various kinds.

2. Explain the general relation of ferrous to ferric compounds, and illustrate by equations the method of passing from one to the other.

3. Compounds of iron.

Give the chemical name and formula of red hematite, iron pyrites, magnetite, loadstone, copperas, green vitriol, iron liquor.

4. Miscellaneous.

(*a*) What elements are related to iron? What metals are closely related to iron, and how is this relation emphasized?

(*b*) Valence of iron in $Fe_2O_3$, $Fe_3O_4$, $FeCO_3$, $Fe(OH)_3$, $FeO$, $FeSO_4$, $Fe_2(SO_4)_3$, $FeS$, $FeS_2$.

(*c*) The group CN, known as cyanogen, has the valence one. What is the valence of the iron atoms in ferrous ferricyanide, $Fe_3(Fe(CN)_6)_2$? In ferric ferrocyanide $Fe_4(Fe(CN)_6)_3$?

(*d*) What is ink? Common bluing? Galvanized iron? Iron rust?

### PROBLEMS. XLII.

1. Berzelius found that 1.586 gm. of iron formed 2.265 gm. of ferric oxide. Calculate the atomic weight of iron. (Equation is $2\,Fe + 3\,O = Fe_2O_3$.)

2. A gram of iron liberated 389.74 cc. of hydrogen at 14° C. and 820 mm. What is the equivalent of iron?

3. What volume of oxygen is used in the oxidation of 100 gm. of ferrous oxide to ferric oxide? (Equation is $2\,FeO + O = Fe_2O_3$.)

4. What volume of hydrogen at 13° C. and 780 mm. is needed to reduce 63 gm. of ferric oxide to metallic iron? (Equation is $Fe_2O_3 + 6\,H = 3\,H_2O + 2\,Fe$.)

**Separation of Metals.** — Many of the foregoing experiments illustrate the fundamental principles of qualitative analysis. The experiments, however, emphasize only the detection of elements or groups, *e.g.* zinc and the sulphate group when testing zinc sulphate. In many instances the various tests would be interfered with by the presence of other substances. Hence it is necessary to separate the essential compounds before applying tests. It is customary to speak of the separation of metals, though usually a compound of the metal is meant. Thus, if a mixture of lead chloride and silver chloride is boiled with water, the lead chloride dissolves, and filtration separates the lead chloride from the silver chloride, or, loosely, the lead from the silver. Once separated, the special tests may be used to detect the various metals. Some separations are simple, but others are complex. A common method of separation is illustrated in Exp. 135, since sulphides are usually solids and behave in various definite ways.

The following experiments illustrate simple separations, but are in no respects a substitute for a course in qualitative analysis. Such a course is beyond the scope of this book.

### SEPARATION OF LEAD AND SILVER.

**Experiment 193.** — Mix 5 cc. each of lead nitrate and silver nitrate solutions, and add dilute hydrochloric acid drop by drop, until precipitation ceases. Allow the mixed precipitates of lead chloride and silver chloride to settle, decant the supernatant liquid down a glass rod, add a little water to the precipitate, and boil. Filter. Test portions of the filtrate for lead (see Exp. 177 (*d*) and (*e*)). Test the precipitate for silver by (*a*) fusing a portion in the blowpipe flame, or (*b*) determining the solubility in ammonium hydroxide, or (*c*) exposing it to the sunlight (a change to purple indicates silver chloride).

### SEPARATION OF LEAD, SILVER, AND MERCURY (-OUS).

**Experiment 194.** — Mix 5 cc. each of lead nitrate, silver nitrate, and mercurous nitrate solutions, and add dilute hydrochloric acid drop by drop, until precipitation ceases. Separate and test the lead as in Exp. 193. Pierce a hole in the point of the filter paper with a glass rod, and wash the mixed precipitates of silver and mercurous chlorides into a test tube with dilute ammonium hydroxide. Warm gently and shake. Filter, and test the filtrate for silver as in Exp. 193. The black residue (see Exp. 166 (*a*)) is a sufficient test for mercury. Its presence may be confirmed thus: Dissolve the black precipitate in a *very* little *aqua regia*, dilute with water, and add a clean copper wire; remove the wire in a few minutes, wipe gently, and mercury will be seen on the wire as a bright silvery coating.

### SEPARATION OF SILVER AND COPPER, OR ANALYSIS OF A SILVER COIN.

**Experiment 195.** — Repeat the first part of Exp. 154 (*b*). Test the precipitate for silver, and the filtrate for copper (see Exp. 151 (*b*) and (*c*)).

### SEPARATION OF COPPER AND ZINC, OR ANALYSIS OF BRASS.

**Experiment 196.** — Dissolve a few grams of fine brass wire in a little dilute nitric acid, evaporate nearly to dryness, dissolve the residue in water, add 10 or 15 cc. of hydrochloric acid, and pass hydrogen sulphide gas into the solution for about twenty minutes, or add considerable hydrogen sulphide water. Filter a little, and add hydrogen sulphide water to see if precipitation is complete; if not, proceed as before, but if complete, continue the filtration.

The filtrate contains the zinc as zinc chloride; the precipitate is copper sulphide. Test the filtrate for zinc (see Exp. 21 (*a*)). Dissolve the copper sulphide in warm dilute nitric acid, and filter, if the solution is not clear. Test the filtrate for copper (see Exp. 151 (*b*) and (*c*)).

### SEPARATION OF LEAD AND TIN, OR ANALYSIS OF SOLDER.

**Experiment 197.** — Dissolve a gram of solder filings in as small a quantity of hot *aqua regia* as possible, evaporate nearly to dryness, dissolve the residue in water, add 10 to 15 cc. of hydrochloric acid, and precipitate the metals as sulphides as in Exp. 196. Filter, wash with hot water, pierce a hole in the filter paper, and wash the precipitate into a test tube with yellow ammonium sulphide. Add more ammonium sulphide, and shake. Filter.

The filtrate contains the tin as ammonium sulphostannate; add to it dilute hydrochloric acid, and yellow stannic sulphide appears.

The precipitate is lead sulphide. Dissolve it in hot dilute nitric acid, filter, and test the filtrate for lead.

### SEPARATION OF COPPER, IRON, AND SODIUM.

**Experiment 198.** — Repeat Exp. 135

### SEPARATION OF ZINC AND IRON.

**Experiment 199.** — Mix 5 cc. each of zinc sulphate and ferric chloride solutions, add an excess of sodium hydroxide drop by drop, and shake vigorously. Filter a little, and add sodium hydroxide to see if precipitation is complete; if not, proceed as before, but if complete, continue the filtration. Test the filtrate for zinc. Scrape the precipitate from the paper into a dish, dissolve in dilute hydrochloric acid, and test portions for ferric iron (see Exp. 190).

### SEPARATION OF ALUMINIUM AND IRON.

**Experiment 200.** — Mix 5 cc. each of alum (common) and ferric chloride solutions, precipitate, and separate as in Exp. 199. Test the filtrate for aluminium (see Exp. 169 (*a*)) and the precipitate for iron as in Exp. 190.

### SEPARATION OF ZINC AND ALUMINIUM.

**Experiment 201.** — Mix 5 cc. each of zinc sulphate and alum (common) solutions. Precipitate and redissolve the hydroxides of zinc and aluminium with sodium hydroxide solution. Pass hydrogen sulphide gas into the solution, or add considerable hydrogen sulphide water. Filter. The precipitate is zinc sulphide, which may be further tested, if desired, though the formation of zinc sulphide under these conditions indicates the presence of zinc. The aluminium is in the filtrate as sodium aluminate. Add to the solution of sodium aluminate hydrochloric acid to acid reaction, and boil; filter, if necessary, and test the filtrate for aluminium.

### CLASS-ROOM EXERCISE. LXIII

1. Devise and describe a method for the separation and detection of the metals in a mixture of —

(*a*) Manganese sulphate and chrome alum.
(*b*) Silver nitrate and alum.
(*c*) Copper and aluminium (aluminium bronze).
(*d*) Lead sulphide and silver sulphide.

# APPENDIXES.

# APPENDIX A.

## MANIPULATION — WEIGHING AND MEASURING.

1. **Hard Glass Tubing** is infusible in ordinary flames, and is used in all operations which require intense heat. It is cut in the same manner as soft tubing, though the scratch must be deeper. It can be bent by using the hottest possible flat flame of a Bunsen burner. Ignition tubes and combustion tubes are made of hard glass tubing and unless they are heated and cooled slowly they will surely crack. Constant care should be taken in using such glass. It is advisable to heat the whole tube gently at first, before directing the flame upon any particular part. Combustion tubes may be safely heated by either (1) winding a piece of fine wire gauze (iron or copper) around the part of the tube to be heated, or (2) winding a piece of copper wire (No. 20) once or twice around each end of the tube so that pieces of the wire project like spokes of a wheel. See Figs. 45 and 59. The latter device has been repeatedly used with success in all the experiments in this book demanding such a precaution.

### MAKING IGNITION TUBES.

**Experiment 6.** — Ignition tubes have thick walls and are made of hard glass. Read again the precautions to be observed in heating hard glass.

Select a piece of hard tubing about 20 centimeters long and of the desired internal diameter (from 7 to 10 millimeters is a convenient size). Warm it by holding it over the flat Bunsen flame, and gradually lower the tube into the flame so that the middle is in the hottest part. Rotate it slowly and evenly. When soft, remove it from the flame, and pull it apart a short distance, as shown in Fig. 84. Then heat *A* so that the flame will melt that part along the dotted line; as soon as it yields to a gentle pull, remove from the flame, and draw it out a little more. Continue this operation until *A* and *B* are separated. Do not pull the two portions apart too quickly, or the glass will be thin at the end. The part *A* may have a little thread or bit of glass on the end. Heat this end and *pinch* it off with the forceps; if large, it may be pulled off, but this operation is liable to remove too much glass and make the end thin. If the closed end is not symmetrical, heat it, and then blow gently into the tube, heating and blowing until the desired shape is produced. Heat *B* at the narrower part and proceed as with *A*.

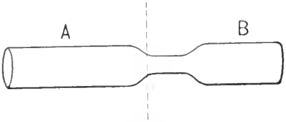

FIG. 84. — Ignition tube, partially made.

If a blast lamp is available, better results may be obtained by the same procedure. Hard glass tubes, after heating, should always be coated with soot from the yellow flame and thus allowed to cool slowly.

**2. Heating.** — Flasks, beakers, retorts, and all glass vessels should never be heated when empty, nor over a direct or free flame even if they contain something, unless the directions so indicate. They should be placed on a piece of iron or brass wire gauze (Fig. 85) supported by a tripod or the ring of an iron stand, and heated gradually from beneath.

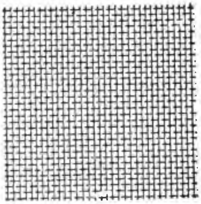

FIG. 85. — A piece of wire gauze.

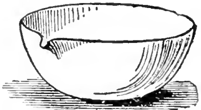

FIG. 86. — Porcelain dish.

Porcelain dishes (Fig. 86) should be heated with even more precaution than large glass vessels. They should always be placed on a piece of gauze, or similar protection, and heated and cooled gradually. They should never be laid on a cold surface when they are hot, but on a block of unfinished wood, a piece of asbestos board, or a straw ring. Porcelain crucibles (Fig. 87), however, owing to their thin walls and small size are usually supported on a triangle (Fig. 88), or a pronged tripod (Fig. 21), and heated safely with a low, free flame. The covers of crucibles often snap, and should be heated with more precaution than the crucibles themselves, especially if the cover has been weighed, since the loss of a weighed cover means a repetition of the experiment.

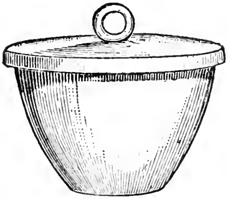

FIG. 87.—Covered porcelain crucible.

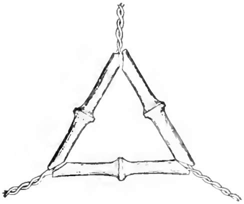

FIG. 88.—Triangle.

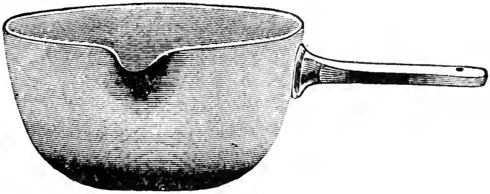

FIG. 89.—Casserole.

An exceedingly convenient and inexpensive porcelain vessel for general use is a casserole (Fig. 89). It is a deep, thin-walled dish with a handle and a flat bottom. It is made in several sizes, and the one containing 125 cubic centimeters is best adapted to ordinary use. It may be heated over a free flame, if partly filled with liquid and kept in motion, though it is safer to employ a piece of gauze.

**3. Evaporation** is the slow conversion of a liquid into vapor. It may be accomplished in several ways. A free flame is best suited for rapid evaporation, but it should never be used for concentrated solutions or semi-fluid masses, as loss or accident may occur by spattering, decomposition, or breakage. As the concentration increases, it is advisable to transfer the vessel to a water bath (Fig. 90), which is intended to accomplish slow evaporation. The containing vessel, which is copper or iron, has a top consisting of movable concentric rings to adapt it to dishes of various sizes. It is filled two-thirds full of water, supported on a tripod or ring of an iron stand, and heated underneath, the vessel meanwhile having been so placed on the bath that it is almost entirely in the steam. The temperature of the solution to be evaporated is thus slowly raised to nearly 100° C. The water in the bath must never be allowed to evaporate entirely or "run dry."

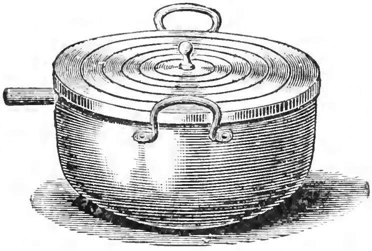
FIG. 90. — Water bath.

Equable heat at a higher temperature than 100° C. may be conveniently obtained by using a sand bath This is simply a shallow iron pan (Fig. 91) filled with dry, clean sand and supported on the ring of an iron stand or on a tripod. The dish to be heated is partly imbedded in the sand, and heat is supplied by a burner under the bath. A small tin pan serves the purpose, though iron pans of all sizes may be obtained from dealers.

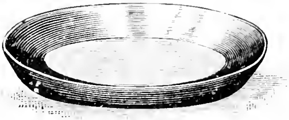
FIG. 91. — Sand bath pan.

Evaporation is also performed by placing the vessel containing the solution on a piece of asbestos board about 15 centimeters square. The asbestos is supported in the same way as the sand bath. An air bath is sometimes used for evaporation (Fig. 92). It is also called a drying oven, for it is really nothing but a small copper oven supported on four legs. A hole in the top contains a cork carrying a thermometer which allows the temperature of the bath to be noted without opening the door. An oil stove and its accompanying baking oven make a convenient air bath.

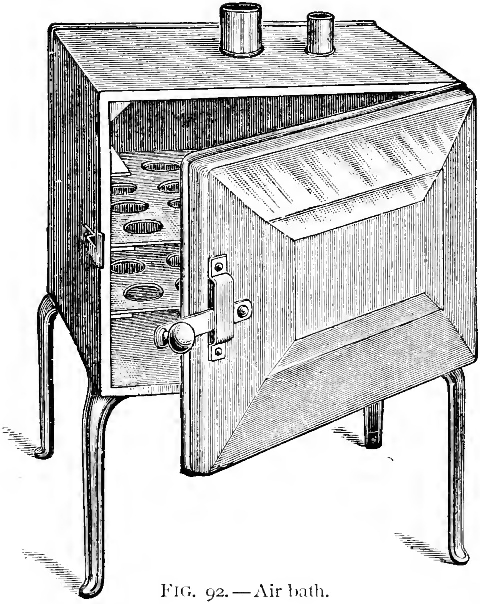
FIG. 92. — Air bath.

4. **Filtration.** — The folded paper should rest firmly on the walls of the funnel so that the weight of the added liquid

will not tear the filter paper. If moistened with water or the liquid to be filtered after being fitted to the funnel, the paper will keep its place more securely and filter more rapidly. The paper should never project above the edge of funnel. The stem of the funnel should be near the wall of the vessel and just above the surface of the liquid to prevent spattering. It is better, also, to pour the liquid to be filtered down a rod upon the side of the paper, as shown in Fig. 93, otherwise the liquid may run down the outside of the containing vessel, as shown in Fig. 94. If a pump is used for rapid filtration, the apex of the filter paper should be protected by a platinum cone, by a small cone of parchment paper pricked with holes, or by a small square of cheese cloth folded into the point of the filter paper.

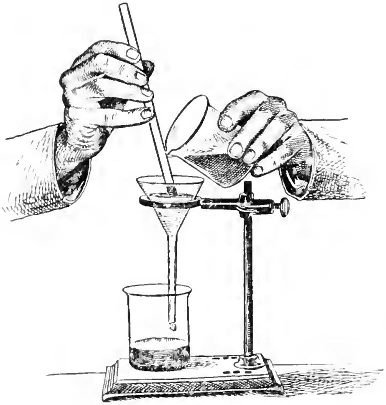

FIG. 93.—Filtering—correct.

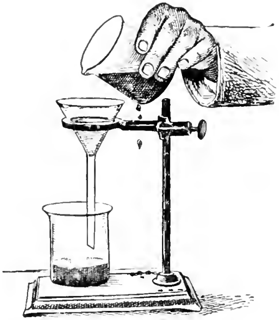

FIG. 94.—Filtering—incorrect.

**5. Stoppers, Corks, Joints, and Safety Tubes.** — Rubber stoppers with one and with two holes are recommended. They are decidedly superior to corks. If corks are used, they should be free from cracks and be softened by rolling or pressing. A convenient way is to wrap the cork in paper to protect it from dirt, and then roll it under the foot on the floor. Holes of any size are made in corks by a cork borer. A set of cork borers is shown in Fig. 95. The stout wire which accompanies each set of cork borers serves as a handle when passed through the holes in the cap of the borer, and also as a piston to remove the cylinder of cork which often remains in the borer after the operation.

FIG. 95. — Set of cork borers.

To bore a hole in a cork proceed as follows: Select a cork free from cracks or channels and use a borer which is one size smaller than the desired hole. Hold the cork between the thumb and forefinger, as in Fig. 96, press the larger end against a firm but soft board, and slowly push the borer by a rotary movement through the cork, taking care to keep the borer perpendicular to the cork. If the hole is too small, enlarge it with a round file. If corks are used instead of rubber stoppers, the apparatus should always be tested before use by blowing into it, stopping of course all legitimate outlets. A poor cork often means a failure, to say nothing of wasted time.

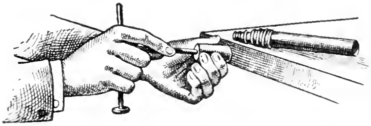

FIG. 96. — Boring a hole in a cork.

Rubber stoppers to be used in quantitative experiments should be boiled in dilute sodium hydroxide solution, rinsed with water, then boiled in dilute hydrochloric acid, and finally washed with water. This operation removes all particles of matter which might adhere to tubes and thereby introduce into the final result a needless error.

Glass tubes are joined by short pieces of rubber tubing called rubber connectors. Such joints should be gas-tight. Before experiments begin leaks should be discovered by testing the apparatus at the necessary points. Do not stop leaks by wax, vaseline, or any temporary makeshift. Select pieces of apparatus which fit. Rubber connectors may be tightened by tying them to the glass tube with a waxed thread. Joints are often made tighter by having on the end of the glass tube a flange, made by heating the end of the tube and pressing it while hot upon a hard surface.

Safety tubes are recommended in place of the ordinary straight thistle tube. See Fig. 30. They should fit perfectly and always have enough liquid in the lower bend to prevent a backward escape of gas. If the liquid to be introduced will not run down, loosen the stopper slightly. If the straight thistle tube is used, it must dip into the liquid in the flask or bottle.

**6. To Cut off the Bottom of a Bottle.**—Select a bottle with walls of uniform thickness. Tie a piece of cotton string loosely around the bottle at the point where it is to be cut. The string should be just tight enough to stay in place and the knot very small. Pour a little kerosene on the string, turning the bottle slowly at the same time, until the string is saturated. The kerosene should follow the string and in no case be allowed to run down the side of the bottle. Light the kerosene at one point and rotate the bottle slowly so that the flame will heat the glass beneath the string. When the circle is complete, quickly plunge the bottle into a pail of water, and it will crack evenly at the desired point. The edges should be smoothed with emery paper. Both parts of the bottle may be utilized in many experiments. The ordinary five-pint acid bottle, or lithia bottle, is well suited to this method.

**7. To Insert a Glass Tube into Rubber Tubing.** — Cut the rubber tubing at an angle, as shown in Fig. 97, moisten the smoothed end of the glass tube with water, place the end of the glass tube in the angular shaped cavity so that both tubes are at about a right angle, and then slip the *rubber* tube slowly up and over the end of the glass tube. If the glass tube is large or the rubber tube stiff, the rubber tube must be held firmly between the thumb and forefinger to keep it from slipping off until it is securely adjusted.

FIG. 97. — Rubber tube cut at an angle (exact size).

**8. To Fit a Glass Tube to a Stopper.** — First round the edges of the tube in the flame. When it is cold, moisten the end with a little water, glycerine, or vaseline (preferably the first), grasp it firmly about an inch from the end, hold the stopper between the thumb and forefinger of the other hand, and work the tube into the hole by a gradual rotary motion. Proceed in the same manner, if the tube is to be pushed through the stopper. *Never* point the tube toward the palm of the hand which holds the stopper. *Never* grasp a safety tube or any bent tube at the bend when inserting it into a stopper — it may break.

**9. To Dry the Inside of a Tube or Bottle.** — Moisten the inside with alcohol, and then with a bellows force in air through a glass tube reaching to the bottom of the vessel; hold the bottle mouth downward and withdraw the glass tube as the alcohol evaporates. Never stand a wet bottle on a hot support, for a drop of water may be jostled down upon the heated bottom and shatter the bottle. The inside will dry if the bottle is exposed to the sun, but the operation is tedious.

**10. To Clean the Inside of a Bottle.** — Partly fill the bottle with water, drop in wads of soft paper, shot, or sand, and shake the bottle vigorously. This device is also applicable to the large tube of a condenser, and to large flasks.

**11. To Introduce a Powder into a Tube.** — First fold a narrow strip of smooth paper so that it will slip into the tube easily. Place the powder at one end of the troughlike holder, and slowly push the paper into the tube, as far as necessary, moving the tube rather than the paper. This operation is shown in Fig. 98. Rotate the tube or turn the paper, and the powder will be deposited at the desired point (Fig. 99). Carefully withdraw the paper. Introduce all powders in this way, whether the tube is open or closed, large or small.

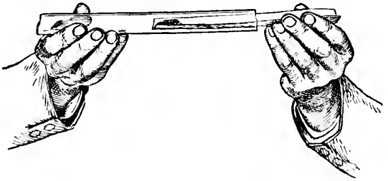

Fig. 98. — Introducing a powder into a tube — first stage.

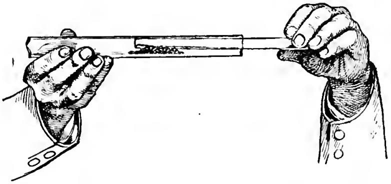

Fig. 99. — Introducing a powder into a tube — second stage.

**12. Pouring Liquids and Transferring Solids.** — Liquids may be poured from a vessel without spilling, by moistening a glass rod with the liquid and then pouring it down the rod as is shown in Fig. 100. The angle at which the rod is held varies with circumstances. This is a convenient way to pour a liquid from a vessel containing a solid without disturbing the solid. Solids should never be poured directly from a large bottle into a test tube, retort, or similar vessel. A convenient method is as follows: Rotate the bottle slowly so that the solid will roll out in small quantities; catch this solid on a narrow strip of paper folded along the middle, and slide the solid from the paper into the desired vessel. The last part of the operation is shown in Fig. 101. Liquids can often be poured from a bottle by holding the bottle as shown in Fig. 102. Notice that the stopper and bottle are held in the same hand. This is accomplished by holding the palm of the hand upward and removing the

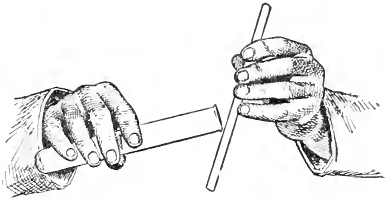

FIG. 100. — Pouring a liquid down a rod.

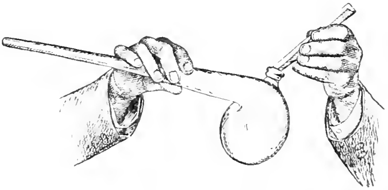

FIG. 101. — Pouring a solid into a vessel with a small opening.

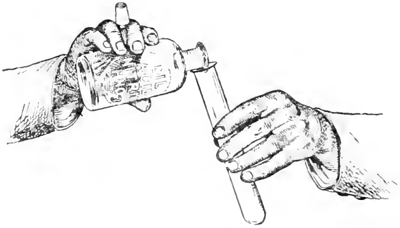

FIG. 102. — The way in which a glass stopper should be held while a liquid is being poured from a bottle.

stopper by grasping it between the fingers before the bottle is lifted. All stoppers should be removed this way, when possible, and never be laid down because the impurities adhering to the stopper may run down into the bottle and contaminate the solution. Never return any reagent to a bottle; if too much has been taken, throw it away. Stoppers and bottles will not become mixed if each bottle and stopper has a similar label.

**13. To Make a Platinum Tip.** — The steps are shown in Fig. 103. *A* is a piece of glass tubing about 12 centimeters long and 5 millimeters in internal diameter. Heat this in the middle in the Bunsen flame and draw it out into the shape shown in *B*; draw it slowly at first, then rapidly. Cut *B* at the point indicated by the dotted line. Roll a piece of thin platinum foil, about 25 millimeters square into a conical tube, as shown in *D*; begin at one corner and roll it around the small end of a round file, gradually shaping the platinum tube so that it will slip into the larger end of *C* but not out of the smaller end. Heat *E* where the dotted line is drawn, and the glass and platinum will firmly unite. The completed tip is like *E*, though it may vary

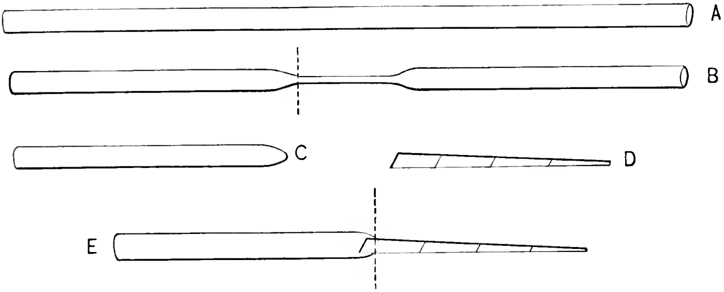

FIG. 103. — Stages in making a platinum tip.

with circumstances. It is used in burning hydrogen and other gases and is attached to the generator by a rubber connector. This tip is more expensive than the customary form, but it is safe and durable. A short piece of pipestem or capillary tubing may be used instead of the platinum tip, but these are not so satisfactory.

**14. To Seal a Platinum Wire into a Glass Rod.** — Rotate one end of a piece of glass rod, about 10 centimeters long, in the flame until it softens. At the same time grasp a piece of platinum wire from 5 to 7 centimeters long firmly in the forceps about 1 centimeter from the end and hold it in the flame. When the rod is soft enough, gently push the hot wire into the rod. Cool the rod gradually by rotating it in the flame. The completed wire is shown in Fig. 104. If a glass tube is used, instead of a rod, it should be drawn out to a very small diameter (see Fig. 103, *B*) before inserting the platinum wire, but in other respects the two operations are practically identical.

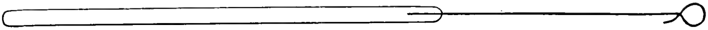

FIG. 104. — Platinum test wire.

**15. Collecting Gases.** — Gases are usually collected over water by means of a pneumatic trough, a common form of which is shown in Figs. 23 and 25. The vessel to be filled with gas is first filled with water, covered with a piece of filter paper, inverted, and placed mouth downward on the shelf of the trough, which is previously filled with water just above the shelf. The paper is then removed and the vessel placed over the hole in the shelf of the trough. Glass plates instead of filter paper may be used to cover the bottles

Figure 105 is a conventional sketch of a cylinder in position in a pneumatic trough. It is represented partly filled with gas which comes through a delivery tube and bubbles up through the water into the cylinder. As the gas rises, water is forced down out of the bottle into the trough. Bottles are usually used in place of cylinders, and the delivery tubes have various shapes. All gases insoluble in water may be collected over water in this way. Some heavy gases, like chlorine, hydrochloric acid, and sulphur dioxide, are collected by allowing the gas to flow downward into an empty bottle and displace the air in the bottle, *i.e.* by *downward displacement*. See Fig. 52. Ammonia gas, being a light gas, is collected by allowing the gas to flow upward into a bottle, *i.e.* by *upward displacement*. See Fig. 58.

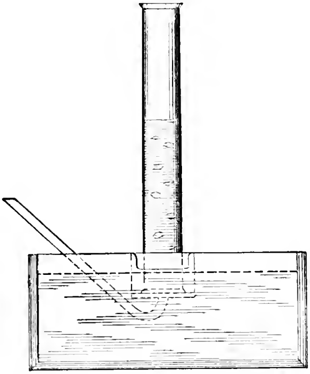

FIG. 105. — Cylinder partly filled with gas in position in a pneumatic trough.

**16. A Gas Holder.** — A small gas holder is constructed as follows: A bottle, as large as available, is provided with a two-hole rubber stopper through which pass two tubes, each bent at a right angle; one tube, *B*, reaches to the bottom, the other, *A*, is just even with the stopper. Attach about 20 centimeters of rubber tubing to each glass tube and provide each rubber tube with a Hofmann screw (Fig. 106) near the end of the glass tube. All joints must be air tight. The holder is filled as follows: Fill the bottle and tubes with water, close *A* with the screw about 5 centimeters from the end, and put the outer end of *B* in a vessel of water. When the apparatus which generates or contains the gas to be introduced is ready, pinch *B* with the thumb and forefinger, loosen the screw on *A*, connect the delivery tube with *A*, and immediately remove the pressure from *B*. The gas will flow in through *A* and the water out through *B*. The initial pressure of the entering gas must be strong enough to start the flow through *B*, and once started the operation proceeds smoothly. When the holder is full, pinch *A*, instantly slip off the delivery tube, and compress *A* firmly with the screw a few centimeters from the end. Attach the end of *B* to the lower opening of an aspirator bottle or to a large funnel supported above the holder to force out the gas. *B* is loosely clamped or left open, but if clamped, the flow from the reservoir to the holder is easily controlled. The gas is driven from the holder by loosening the screw attached to *A*. Do not attach *B* to the faucet, to force out the gas, as air from the water pipe is apt to pass into the holder along with the water.

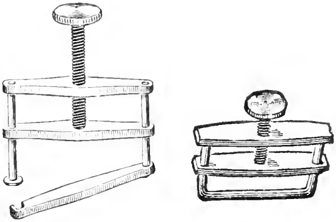

FIG. 106. — Hofmann screws.

**17. Aspirators.** — It is often necessary to draw (or force) a gas through a tube. This may be done by a filter pump, or by an aspirator. The former may be obtained from the dealer, the latter is easily constructed. A common form is shown in Fig. 107. A five-pint acid bottle is provided with a two-hole rubber stopper carrying two open tubes. The shorter is bent at a single right angle, it extends just within the bottle, and is called the inlet tube.

The longer tube is bent at a double right angle, the inner arm of which extends nearly to the bottom of the bottle, while the other arm terminates at any convenient point below the bottom of the bottle. A Hofmann screw, attached to a piece of rubber tubing, *A*, near the end of this outer tube, serves to regulate the flow of water. If more convenient, the rubber tube may replace the greater portion of the outer arm of the outlet tube. An aspirator bottle may be obtained from the dealer, but one made as above serves the purpose. To draw — or aspirate — a gas, fill the bottle with water and insert the stopper with its tubes, attach the shorter arm to the apparatus, loosen the screw, start the water flowing, and regulate the flow by the screw.

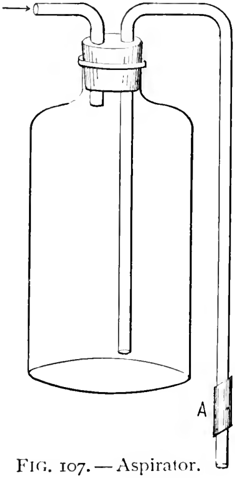

FIG. 107. — Aspirator.

18. **The Metric System** originated in France during the French Revolution, and its use has since been required or legalized in most civilized countries. The fundamental unit of the system is the meter, which is approximately equal to the ten-millionth part of the distance from the equator to the north pole. This distance was ascertained by actual measurement of an arc of a meridian passing through Barcelona in Spain and Dunkirk in France. The legal equivalent of the meter in the United States is 39.37 inches.

The superior advantage of the metric system is its decimal character, which allows rapid transformations from volume into weight and *vice versa*, and between denominations of the same unit. Each unit has multiples and submultiples, which are designated by prefixes attached to the particular unit. The prefixes denoting multiples are *deca-*, *hecto-*, and *kilo-*, equivalent respectively to 10, 100, and 1000. The submultiple prefixes are *deci-*, *centi-*, and *milli-*, which correspond respectively to .1, .01, and .001.

The meter — the unit of length — is seldom used in chemistry, but the centimeter is often employed to express the length of a tube, or the linear dimensions of a large vessel. The height of the barometer is stated in centimeters or millimeters, and the length of the smallest pieces of apparatus is often expressed in the latter denomination.

The unit of weight is the gram, which, like the meter, has its multiples and submultiples. The gram itself is actually derived from the kilogram, which is the weight in a vacuum of a cubic decimeter, or 1000 cubic centimeters, of pure water at its maximum density. Hence *one gram is the weight of one cubic centimeter of water*. The weights of small masses are expressed in terms of the gram. If an object weighs, for example, 2 grams, 2 centigrams, and 5 milligrams, the weight is written 2.025 grams, though the two small weights may read 20 and 5 milligrams. Two milligrams is more often written as .002 gram than as 2 milligrams, though both forms are used. The decimal form of the fraction is always used in the metric system. Thus, 4 decigrams is *not* written $\frac{4}{10}$ grams, but .4 gram, or often, 0.4 gm.

The unit of volume is the liter, which is used for both dry and liquid measure. It is equal to the capacity of the vessel containing the standard kilogram. It therefore contains 1000 cubic centimeters, *i.e.* it is a cubic decimeter. A liter of water weighs 1000 grams. Volume is usually expressed in cubic centimeters, and most graduated volumetric apparatus is described by this denomination. Thus a liter flask is often marked 1000 cubic centimeters; burettes are made to deliver 50 cubic centimeters, or 100 cubic centimeters, and pipettes from 1 to 100 cubic centimeters. The following tabular view will make clear the relations between the units, their multiples, and submultiples:—

| LENGTH. | WEIGHT. | VOLUME. | NOTATION. |
|---|---|---|---|
| Kilometer | Kilogram | Kiloliter | 1000. |
| Hectometer | Hectogram | Hectoliter | 100. |
| Decameter | Decagram | Decaliter | 10. |
| METER | GRAM | LITER | 1. |
| Decimeter | Decigram | Deciliter | 0.1 |
| Centimeter | Centigram | Centiliter | 0.01 |
| Millimeter | Milligram | Milliliter | 0.001 |

From this table it is evident that 10 milligrams equal 1 centigram, 10 centigrams equal 1 decigram, 10 decigrams equal 1 gram, and so on. The numerical significance of the prefixes is also apparent.

The passage from the English to the metric system may be accomplished by utilizing the following:—

### TABLE OF TRANSFORMATION.

| TO CHANGE | MULTIPLY BY |
|---|---|
| Inches to centimeters . . . . . . . . . . . . . . . . . | 2.54 |
| Centimeters to inches . . . . . . . . . . . . . . . . . | .3937 |
| Cubic inches to cubic centimeters . . . . . . . . . . | 16.387 |
| Cubic centimeters to cubic inches . . . . . . . . . . | .061 |
| Ounces to grams . . . . . . . . . . . . . . . . . . | 28.35 |
| Grams to ounces . . . . . . . . . . . . . . . . . . | 0.0353 |

**19. To counterpoise** a balance proceed as follows: Clean the pans with soft paper or cheese cloth, and allow them to swing freely to ascertain the exact condition of the balance. If the pointer does not make equal excursions on each side of the middle or zero point, then add to the lighter side weight sufficient to restore equilibrium. Do not wait for the pointer to come to rest, but estimate distances to the right and left. Balls of paper, bits of wire or of match, may be added to the proper pan of the horn pan balance to secure equilibrium. Often a piece of wire bent in the form of a spring is attached to the beam, if the balance is persistently "off." Weight is added to the lighter pan of the trip scales by properly adjusting the screws near the pointer.

The horn pan balance may be protected from drafts by hanging the balance from the top of a box shaped like an ordinary balance case. The box is open in front, but the rear is covered with cheese cloth, held firmly in place by half-round moulding. The box and its enclosed balance are shown in Fig. 108.

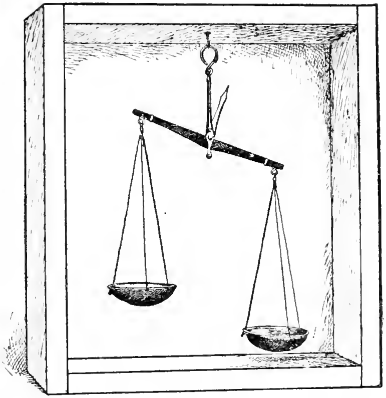

FIG. 108. — Horn pan balance in a case.

**20. Weights** from the kilogram to the gram are made of iron or brass, and the smaller weights of platinum or some other durable metal. They may be obtained in sets possessing any degree of accuracy. A set from 50 gm. to

1 cg. is large enough for the exact work in this book. See Fig. 109.

**21. Weighing.** — The object to be weighed is placed in the center of the left-hand pan of the counterpoised balance, and in the other pan a weight is placed, which is assumed to be the approximate weight of the object. If the weight is too heavy, return it to the box and place on the pan a weight or several weights smaller in amount. Proceed thus, adding and removing weights, until the balance is judged to be in equilibrium; then allow the pointer to swing freely several times, and note the distances. If they are unequal, adjust the weights accordingly; if they are equal, the balance is in equilibrium and the sum of the weights is the weight of the object. Record the weight immediately in the proper place in a note-book — *not on a scrap of paper*. The most satisfactory method of determining the correct sum of the individual weights is as follows: Add the weights missing from the box and record their sum; then add the weights on the balance and compare with the amount recorded; finally, as the weights are replaced, beginning with the largest, add again and check the first result. This operation takes

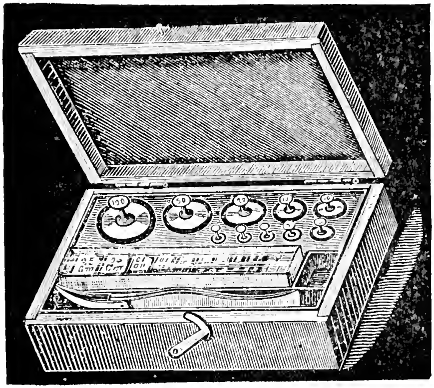

FIG. 109. — Set of weights.

time, but assuredly less time than that consumed by a repetition of the weighing or of an experiment.

Certain precautions must be observed in weighing.

(1) Substances should not be weighed on the bare pan, but on a piece of smooth paper creased on the sides or in the middle, or in some counterpoised vessel, *e.g.* a watch glass, beaker, crucible, tube, or flask.

(2) Never touch the weights with the fingers. Use clean forceps.

(3) Arrange the weights symmetrically about the center of the pan. This is especially necessary with large weights on the trip scales.

(4) Never attempt to weigh the exact amount specified, *e.g.* 1.49 gm., but weigh accurately an approximate amount, *e.g.* 1.47 gm., or 1.50 gm., unless, of course, the directions state exactly 1.49 gm. The expression "from 2.1 to 2.9 gm." means any weight between these two, but the weight, whatever it is, must be exactly known; the two amounts, 2.1 and 2.9 gm., are simply limits adapted to the experiment. So also the expression, "weigh about 2.4 gm.," means an *exact* amount which approximates 2.4 gm.

(5) If the same object is to be weighed more than once, it should be weighed, if possible, with the same balance and weights. If the balances are carefully counterpoised, this precaution is, however, not absolutely necessary.

(6) The balance and weights must always be left in a usable condition — and they will always be so found.

**22. Measuring Liquids and Gases.** — Liquids are measured in graduated cylinders, or graduates, tubes, burettes, and pipettes. Vessels of various capacities are used, depending upon the volume of liquid to be measured. A liter graduate is shown in Fig. 110. A graduate holding 100 cc. is best adapted to general use. Small volumes are more conveniently measured in a burette, or in a pipette (Fig. 40). Burettes and pipettes also hold various volumes, though the burettes most often used hold 50 cc., and pipettes 10 cc., 25 cc., and occasionally 50 cc. The method of using a burette is explained in the experiment on neutralization (Exp. 53). The pipette is used when definite small volumes are wanted, such as 10 cc., or from 1 to 5 cc. The pointed end is dipped into the liquid, which is then gently sucked up into the tube to a point just above the mark on the stem, and the top of the tube is quickly closed with the forefinger. The relative position of the hand and pipette at this stage is shown in Fig. 40. If the pressure of the finger is lessened, the liquid will slowly fall to the mark on the stem, and farther if desired. Pipettes hold the indicated volume between this mark and the extreme lower end. Occasionally flasks holding a liter or its fractions are used to obtain exact volumes.

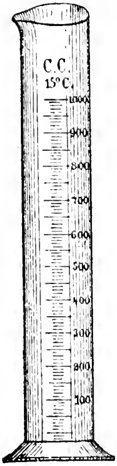

FIG. 110. — Liter graduate.

Gases are measured roughly by collecting them over water in a vessel the capacity of which is known or subsequently determined by filling the space occupied by the gas with water from a graduate or burette. In accurate experiments, the gas is collected in a graduated tube called a gas measuring tube (see Figs. 32 and 48), or in a graduated tube having two platinum wires sealed into the walls near the closed end, and called a eudiometer (see Figs. 111 and 44). The latter is used only in those cases in which electric sparks are to be passed through the contents of the tube. Often the capacity of a plain tube may be found by a burette and used instead of the more expensive graduated tube.

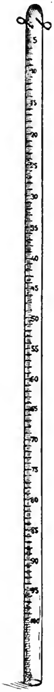

FIG. 111.—Eudiometer.

Certain essential precautions must be observed in reading volumes of liquids or gases.

(1) Notice the exact value of a unit space on the graduated scale, since single spaces often vary in value with vessels of different capacities. Thus one space may, and often does, equal 0.1 cc., 0.2 cc., 1 cc., 2 cc., and even 5 cc. on different vessels.

(2) The surface of most liquids in a tube, especially in a small tube, is not horizontal but concave. This curved surface is called the *meniscus*. It is customary to call the lowest point of the meniscus the correct height of the column of water, or depth of the column of gas, as the case may be. Thus in Fig. 112 the correct reading of the meniscus is 11.7 cc., though it may be regarded erroneously as more or less, if the eye were at II., or III., instead of at I., the correct position. The eye must be in a line tangent to the lowest point of the meniscus in order to determine the correct reading. The reading 11.7 cc. is the depth of the gas column in Fig. 112. If the height of the water column is desired, as is often the case, then 11.7 must be subtracted from the graduated capacity

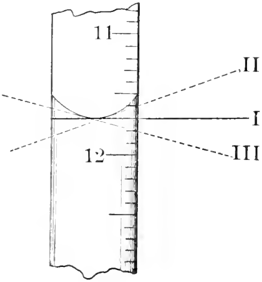

FIG. 112.—Meniscus. Correct reading is along the line I.

of the vessel, if the graduation is from the top downward. Usually the graduation of a cylinder or graduate begins at the bottom, and the height of the meniscus is then the actual height of the water column. Sometimes tubes are graduated in both directions, and the reading is then simple.

**23. A Thermometer** (Fig. 113) is an instrument for measuring temperatures. There are two kinds in use; the one in scientific use is the Centigrade, the one in general use in this country is the Fahrenheit. The change in temperature is measured by expansion and contraction of the mercury, or popularly speaking, by "rise and fall of the thermometer." The abbreviation for centigrade is C. and for Fahrenheit is F.

The point where the mercury stands when held for a time in boiling water is called the boiling point of water, and the point where the mercury stands when the thermometer is held in melting ice is called the freezing point of water.

On the centigrade thermometer the boiling point is 100, and the freezing point is zero; on the Fahrenheit thermometer the boiling point is 212, and the freezing point is 32. The space between these two points is divided into equal portions called degrees. The abbreviation for degrees is °, *e.g.* 212°. Between these two points on the centigrade scale, as the space is called, there are 100

Fig. 113.— Thermometer.

equal portions; but on the Fahrenheit scale there are 180 equal portions. Hence 100 degrees centigrade equal 180 degrees Fahrenheit. Since

$$100° \text{ C.} = 180° \text{ F.}$$

$$\text{then} \quad 5° \text{ C.} = 9° \text{ F.}$$

$$\text{or} \quad 1° \text{ C.} = \tfrac{9}{5} \text{ of } 1° \text{ F.}$$

But as all readings are made from 0°, 32 must be added when the change is made *into* Fahrenheit, and subtracted when the change is made *from* Fahrenheit. Hence

$$\text{F.} = \tfrac{9}{5} \text{ C.} + 32$$

$$\text{C.} = \tfrac{5}{9} (\text{F.} - 32)$$

Notice that in the second formula 32 is to be subtracted before multiplying by $\frac{5}{9}$.

Observe the following precautions in using a thermometer: —

(1) A thermometer should never be thrust suddenly into a hot liquid, but held in the vapor until the temperature registered is nearly that of the liquid.

(2) Always allow the thermometer to stand in a liquid long enough to assume the temperature of the liquid, before reading the thermometer.

(3) Verify all readings by a second or independent observation.

(4) Record all readings in the note-book as soon as made.

**24. The Barometer** is used to measure the pressure of the atmosphere.

The normal height of the barometer is 760 mm. It is assumed in all problems in this book that the pressure is 760 mm., unless otherwise stated.

Precautions (3), (4) under thermometers must also be observed in reading the barometer.

# APPENDIX B.

## I. GAS LAWS.

**1. Law of Charles.** — It has been found that all gases under constant pressure expand or contract uniformly for the same change of temperature. This law, known as the *Law of Charles*, may be stated more explicitly thus: —

*A given volume of gas under constant pressure expands or contracts $\frac{1}{273}$ of its volume at zero degrees centigrade for every degree through which it is heated or cooled.*

This law means that —

273 cc. at 0° become $273 + 1 = 274$ cc. at 1° C.

273 cc. at 0° become $273 + 2 = 275$ cc. at 2° C.

273 cc. at 0° become $273 + 3 = 276$ cc. at 3° C.

or 273 volumes at $t°$ become $273 + t$ volumes at $t°$ C.

Let $V =$ the volume at 0° C.

and $V' =$ the volume at $t°$ C.

Then $V : V' :: 273 : 273 + t$.

This proportion may be solved for either $V$ or $V'$; thus —

$$V = \frac{V' \times 273}{273 + t}. \tag{1}$$

$$V' = \frac{V(273 + t)}{273}. \tag{2}$$

From (1) we can easily calculate the volume a gas would occupy if at 0° C. For example, suppose a volume of gas at 17° C. measures 245 cc., its volume at 0° C. would be found thus:—

$$V = \frac{245 \times 273}{273 + 17} = 230.6 \text{ cc.}$$

And from (2) the volume at any temperature may be found, if the volume at 0° C. is known. Thus if the volume is 230.6 cc. at 0° C., the volume at 17° C. would be found by substituting the proper values in (2). Thus—

$$V' = \frac{230.6\,(273 + 17)}{273} = 245 \text{ cc.}$$

**2. Correction for Temperature.** — Since the volume varies with the temperature, gas volumes to be comparable must be at the same temperature. This situation is not always possible, hence it is customary to reduce the observed volume by formula (1) to the volume it would occupy if it were to exist at 0° C., which is the normal or standard temperature. The operation is called "reducing to standard temperature," or "correcting for temperature." The significance of the law and the correction for temperature will be clearer after the solution of some typical problems.

### PROBLEMS. III.

*In the following problems the pressure is to be regarded as constant.*

1. Find the volume which 173 cc. of gas at 12° C. would occupy at 0° C.

2. If a gas volume measures 466 cc. at 14° C., what would it measure at 0° C.?

3. What would be the volume at 0° C. of a volume of gas which measures 706 cc. at 15° C.?

4. A certain volume of a gas measures 25 cc. at 27° C. What would be its volume at 0° C.?

5. Two gases are equal in volume; one is at −20° C., the other is at +20° C. What would each volume be at 0° C.?

6. A gas volume measures 3 l. at 0° C. At what temperature would its volume be 4 l.?

7. A volume of gas at 13° C. measures 100 cc. What would be its volume at −130° C.? At −13° C.? At +130° C.?

8. If 1000 cc. of gas are heated from 0° C. to 39° C, what is the new volume?

9. If 1000 cc. of gas are heated from −39° C. to 52° C., what is the new volume?

10. Suppose 10 cc. of gas are measured at 15° C. What volume will the gas occupy at 150° C.?

11. If 500 cc. of nitrogen at 27° C. are cooled to −5° C., what is the new volume?

12. A volume of gas measured 120 cc. at 15° C. What will be the volumes at 30° C., 0° C., and −10° C.?

13. If 743 cc. of gas are heated from 47° C. to 83° C., what is the new volume?

14. One thousand cubic centimeters of air at 13° C. will occupy what volume at 65° C.?

15. If 155 l. of air are cooled from 150° C. to 0° C., what will be the new volume?

16. A liter of oxygen is heated from 14° C. to 42° C. Find the new volume.

17. A volume of carbon dioxide measures 1500 cc. at 0° C. What will be its volume at each of the following temperatures: (*a*) 15° C., (*b*) 50° C., (*c*) 100° C., (*d*) 300° C.? At what temperature will it measure a liter?

18. If a volume of hydrogen measures 100 cc. at 100° C., what will be the volume at −100° C.?

19. Twenty-five volumes of air measured at −10° C. are heated until they occupy 30 volumes. To what temperature must they have been heated?

**3. Absolute Temperature.** — Since a gas contracts $\frac{1}{273}$ of its volume at zero C. for every degree through which it is cooled, then at −273° C. the gas would have no volume!

This condition is not possible, and as a matter of fact as gases approach this point (which has not yet been reached), they deviate from the law. However, for convenience the point −273° C. is regarded as a starting or zero point, and is called absolute zero. *Absolute* temperature is reckoned from this point, just as centigrade temperature is reckoned from zero on the centigrade scale. Degrees on the absolute scale are found by adding 273 to the readings of the centigrade thermometer. Thus—

273° absolute = 0° C.
274° absolute = + 1° C.
272° absolute = − 1° C.

Hence the Law of Charles is often stated thus:—

*The volume of a gas under constant pressure varies directly as its absolute temperature.*

**4. Law of Boyle.**—The effect of pressure on gases was early discovered, but it was not until about 1662 that Boyle announced his Law:—

*The volume of a gas at a constant temperature varies inversely as the pressure.*

This law means that 200 cc. of gas under a pressure of 125 mm. become—

100 cc. under a pressure of 250 mm.
50 cc. under a pressure of 500 mm.
25 cc. under a pressure of 1000 mm.

or,

$$200 \times 125 = 25 \times 1000.$$

If

$$200 \text{ cc.} = V$$
$$125 \text{ mm.} = P$$
$$25 \text{ cc.} = V'$$
$$1000 \text{ mm.} = P',$$

then

$$V \times P = V' \times P'$$

or,

$$V : V' :: P' : P$$

This proportion may be solved for $V$ or $V'$; thus—

$$V = \frac{V'P'}{P}$$

$$V' = \frac{VP}{P'}$$

By agreement among chemists, $P$ is equal to 760 mm. Hence the above equations become respectively:—

$$V = \frac{V'P'}{760} \qquad (3)$$

$$V' = \frac{V \times 760}{P'} \qquad (4)$$

Suppose a gas at 740 mm. measures 420 cc., its volume at 760 mm. would be found thus:—

$$V = \frac{420 \times 740}{760} = 408.9$$

Formula (4) is not often used in actual work, but it serves to extend the conception of the general application of the law.

**5. Correction for Pressure.**—In order to compare gas volumes they must be measured when under the same pressure. This, however, is impracticable, or, at least, inconvenient. Hence it is customary, as in the case of the temperature correction, to measure each volume under the existing pressure, which is read directly from the barometer, and then to reduce the observed volume to the volume it would occupy if at 760 mm., which is the standard or normal pressure. This operation is called "reducing to standard pressure," or "correcting for pressure."

### PROBLEMS. IV.

*In the following problems the temperature is to be regarded as constant.*

1. The volume of a gas at 740 mm. is 200 cc. What is the volume at 760 mm.?

2. If a gas volume at 780 mm. is 25 cc., what would the volume be at 760 mm.?

3. A gas holder contained 4.5 l. of oxygen when the barometer stood at 755 mm. What would the volume be at 762 mm.?

4. A tube contained 137 cc. of air when the barometer stood at 766 mm. The next day the barometer reading was 757 mm. What was the new volume?

5. A gas measures 1000 cc. at 770 mm. What is its volume at 530 mm.?

6. A volume of hydrogen measures 467 cc. at 756 mm. What would its volume be at the normal pressure?

7. A volume of gas was measured and found to be 195 cc. when the barometer stood at 740 mm. If the volume should be increased to 200 cc., what would be the barometer reading?

8. A gas volume was 467 cc. at 756 mm. What would the volume have been at 29.92 in. by the barometer?

### CLASS-ROOM EXERCISE. II.

1. Historical.
    (*a*) Essential facts in the life of Charles.
    (*b*) Essential facts in the life of Boyle.
    (*c*) Mariotte's share in the discovery of Boyle's Law.
2. Experimental.
    (*a*) Proof of Boyle's Law.
    (*b*) Proof of Charles's Law.

**6. Aqueous Tension.** — When different gases are confined in the same vessel, each gas exerts its pressure against the atmospheric pressure. Hence, when a gas is collected over water, as many gases are, it is saturated with water vapor and does not bear the full pressure of the atmosphere, since part is borne by the water vapor. Hence the volume of a gas is increased by the pressure of water vapor. The pressure of water vapor, or aqueous tension, as it is usually called, has been calculated in terms of millimeters of mercury, and may be subtracted directly from the barometer reading. Hence the pressure under which a dry gas is to be regarded as existing is really $P' - a$, $a$ being the symbol commonly employed to designate aqueous tension.

The amount of water vapor present in a volume of saturated gas varies with the temperature, hence the pressure exerted by this confined water vapor in its effort to escape varies with the temperature.

The conception and significance of the tension of aqueous vapor is not always grasped by students. The fact may be illustrated as follows: —

Take a glass tube about 90 cm. long and closed at one end, and dry and warm it. Fill it with dry, warm mercury and invert in a mortar containing the same liquid. The mercury will fall a little in the tube. Introduce a little water by means of a small pipette or by a medicine dropper, and the mercury will soon fall slightly, owing to the pressure of the water vapor — tension of aqueous vapor — in the space above the mercury. Warm the upper end of the tube with the hand or a lighted candle, and the mercury falls still lower. If it were possible to boil water in the tube, the pressure of the vapor would be equal to the pressure of the atmosphere at that instant, since water boils when it just overcomes atmospheric pressure. The correction is often larger than might be suspected and must be made in all accurate work.

Gases are assumed to be saturated when measured over water, but their volumes are compared when dry. Hence a "correction for aqueous tension" as well as for temperature and pressure must be made before the volumes are comparable.

The following is a —

TABLE OF THE TENSION OF AQUEOUS VAPOR (REGNAULT).

| *t.* | mm. | *t.* | mm. | *t.* | mm. | *t.* | mm. |
|---|---|---|---|---|---|---|---|
| 10 | 9.17 | 16 | 13.54 | 22 | 19.66 | 28 | 28.10 |
| ·5 | 9.47 | ·5 | 13.97 | ·5 | 20.27 | ·5 | 28.93 |
| 11 | 9.79 | 17 | 14.42 | 23 | 20.89 | 29 | 29.78 |
| ·5 | 10.12 | ·5 | 14.88 | ·5 | 21.53 | ·5 | 30.65 |
| 12 | 10.46 | 18 | 15.36 | 24 | 22.18 | 30 | 31.55 |
| ·5 | 10.80 | ·5 | 15.85 | ·5 | 22.86 | ·5 | 32.46 |
| 13 | 11.16 | 19 | 16.35 | 25 | 23.55 | 31 | 33.41 |
| ·5 | 11.53 | ·5 | 16.86 | ·5 | 24.26 | ·5 | 34.37 |
| 14 | 11.91 | 20 | 17.39 | 26 | 24.99 | 32 | 35.36 |
| ·5 | 12.30 | ·5 | 17.94 | ·5 | 25.74 | ·5 | 36.37 |
| 15 | 12.70 | 21 | 18.50 | 27 | 26.51 | 33 | 37.41 |
| ·5 | 13.11 | ·5 | 19.07 | ·5 | 27.29 | ·5 | 38.47 |

CLASS-ROOM EXERCISE. III.

1. Would all liquids have the same vapor tension at a given temperature?

2. Why should a gas, the volume of which is to be measured over water, be *saturated* before its volume can be accurately determined?

3. Does the amount of water vapor alter the pressure for a given temperature?

**7. Formula for the Reduction of Gas Volumes to Standard Conditions.** — Since gas volumes to be comparable must be corrected for temperature, pressure, and aqueous tension (*i.e.* reduced to standard conditions), it is convenient to make the corrections simultaneously by the formula:—

$$V = \frac{V'(P' - a)}{760(1 + .00366\,t)} \qquad (5)$$

In this formula—

$V$ = the corrected volume.

$V'$ = the observed volume.

$P'$ = the observed pressure.

$t$ = the observed temperature.

$a$ = the aqueous tension at $t°$ C.

"Corrected" means "under standard conditions," and "observed" means "under any other conditions."

This formula is so important, its derivation should be mastered.

(*a*) **Temperature Correction.**—Since all gases expand $\frac{1}{273}$ of their volume at zero for each rise of a degree centigrade, then

$$V' = V + \left(\frac{t}{273} \times V\right)$$

But $\frac{t}{273} = .00366 \times t$. Hence the above equation becomes—

$$V' = V + (.00366 \times t) \times V$$

$$= V \times (1 + .00366\,t)$$

$$\therefore V = \frac{V'}{(1 + .00366\,t)}$$

(*b*) **Pressure Correction.**—Since volumes of gases are inversely proportional to pressure, the volume relations of the same gas under two different pressures may be expressed by the proportion —

$$V : V' :: P' : P.$$

Therefore $$VP = V'P'$$

or $$V = \frac{V'P'}{P}$$

and since $P = 760$, $$\therefore V = \frac{V'P'}{760}$$

(*c*) **Combined Corrections for Temperature and Pressure.** — To reduce a gas volume to 0° and 760 mm. it must first be multiplied by $\frac{1}{(1 + .00366\,t)}$ and then by $\frac{P'}{760}$

Hence $$V = V' \times \frac{1}{(1 + .00366\,t)} \times \frac{P'}{760}$$

$$= \frac{V'P'}{760\,(1 + .00366\,t)} \qquad (6)$$

Formula (6) is used when there is *no* correction for aqueous tension.

(*d*) **Combined Corrections.** — Since the correction for aqueous tension is made by subtracting from the observed pressure the aqueous tension in millimeters of mercury (found in the Table in § 6), formula (6) becomes -

$$V = \frac{V'(P' - a)}{760\,(1 + .00366\,t)}$$

An example will make the formula more intelligible. A student actually found that a mixture of potassium chlorate and manganese dioxide yielded 2192.4 cc. of oxygen when the barometer read 757.8 mm. and the thermometer 21.5° C. The aqueous tension was taken as 19.097. Substituting these values, the formula becomes —

$$V = \frac{2192.4\,(757.8 - 19.097)}{760\,(1 + .00366 \times 21.5)}$$

$$= \frac{2192.4 \times 738.703}{760\,(1 + .07869)} = 1980.6+ \text{ cc.}$$

Therefore, 1980.6 cc. is the volume which would be occupied at standard conditions by 2192.4 cc. at the observed temperature and pressure.

In this solution notice (1) that the aqueous tension is subtracted from the barometer reading and *that remainder* is multiplied by the observed gas volume, and (2) that the temperature is multiplied by .00366 and 1 is added to *that result* before multiplying by 760.

## PROBLEMS V.

*In these problems the gas is to be regarded as dry, unless otherwise stated.*

1. Reduce to standard conditions the following: —

(*a*) 147 cc. measured at 570 mm. and 136.5° C.
(*b*) 320 cc. measured at 950 mm. and 91° C.
(*c*) 480 cc. measured at 380 mm. and 68.25° C.
(*d*) 25 cc. measured at 780 mm. and 27° C.
(*e*) 14 cc. measured at 763 mm. and 11° C.
(*f*) 18 cc. measured at 742 mm. and 14° C.
(*g*) 0.18 cc. measured at 764 mm. and 20° C.
(*h*) 375 cc. measured at 745 mm. and 17° C.

2. A vessel 10 cm. long, 5 cm. wide, and 3.5 cm. deep, is filled with gas at 770 mm. and 100° C. What is its volume under standard conditions?

3. A gas under standard conditions has its pressure doubled and its temperature raised until the volume is the same as the original volume. What is the final temperature?

4. A gas volume measures 12 liters at 0° C. and 760 mm. What would its volume be at 800 mm. and 17° C.?

5. A gas measures 150 cc. at 10° C. and 500 mm. What would be its volume at 16.4° C. and 520 mm.?

6. If 852 cc. of hydrogen are measured at 11° C. and 760 mm., what would be the volume at 27° C. and 900 mm.?

7. A certain volume of hydrogen measures 250 liters at 745 mm. and −15° C. What would be its volume at 20° C. and 765 mm.?

8. What volume under standard conditions would be occupied by 200 cc. of gas at 14° C. and 756 mm.?

9. Reduce to standard conditions 1328 cc. of gas saturated with water vapor and measured under the following conditions: —

Pressure, 765 mm.
Temperature, 18° C.
Aqueous tension, 15.357 mm.

10. Correct 99.5 cc. of air for pressure, temperature, and aqueous tension from the following data: —

Pressure, 756.3 mm.
Temperature, 20.5° C.
Aqueous tension, 17.943 mm.

11. What would be the corrected volume of 77.7 cc. of saturated air, if measured at 17.5° C. when the barometer stood at 755.5 mm., allowing 14.889 mm. for aqueous tension?

12. Reduce to standard conditions in the following cases the gas volume measured under the given conditions: —

(*a*) 97.2 cc.
20.3° C.
756 mm.
17.65 *a.*

(*b*) 77 cc.
17.5° C.
755 mm.
14.89 *a.*

(*c*) 81.2 cc.
746.8 mm.
19.5° C.
16.87 *a.*

(*d*) 100 cc.
755.3 mm.
18.5° C.
15.85 *a.*

(*e*) 100 cc.
756 mm
24° C.
22.18 *a.*

(*f*) 99.8 cc.
753.2 mm.
20.3° C.
17.65 *a.*

(*g*) 79 cc.
754 mm.
19.5° C.
16.35 *a*.

(*h*) 98.8 cc.
756.5 mm.
20.5° C.
17.94 *a*.

(*i*) 78.6 cc.
758.4 mm.
22.5° C.
20.29 *a*.

(*j*) 97.8 cc.
756.3 mm.
19.5° C.
16.87 *a*.

(*k*) 77 cc.
758.4 mm.
19.5° C.
16.87 *a*.

### PROBLEMS. VI. (REVIEW.)

1. A certain volume of air measures 150 cc. when the barometer stands at 760 mm. On the following day its volume had decreased 1.52 cc. Calculate the change in the height of the barometer.

2. 146 cc. of air at 10° C. and 730 mm. are measured over water. What would the dry air measure under standard conditions?

3. A certain quantity of nitrogen measures 155 cc. at 10° C., and under a pressure of 530 mm. What will the volume become at 18.7° C. and under a pressure of 590 mm.?

4. A certain weight of air measures a liter at 0° C. How much will the air expand if heated to 100° C.?

5. A liter of air at 39° C. is cooled to – 26° C. Find the new volume.

6. 134 l. of air are heated from – 30° C. to 60° C. Find the new volume.

7. 50 cc. of air at 10° C. occupy what volume at 24° C.?

8. 100 cc. of air at 12° C. are heated until they occupy 145 cc. Find the new temperature.

9. 700 cc. of air at 77° C. and 1000 mm. measure what at 163° C. and 872 mm.?

10. A quantity of air which measures 230 l. at 14° C. and 740 mm. will measure what at 0° C. and 760 mm.?

## II. MANIPULATION OF GASES.

Certain facts, simple in themselves but easily overlooked, render work with gases somewhat difficult.

(1) Most gases are colorless and their escape is not easily discovered.

(2) They are difficult to dry and purify.

(3) Their volume changes quickly and considerably with changes of temperature and pressure.

(4) They cannot be weighed directly without delicate apparatus and equable surroundings, hence their weight must be indirectly determined.

(5) Several exact readings must be made.

Nevertheless, gases present a most fruitful field of study, and if certain precautions are heeded, fairly accurate results are possible. Inasmuch as many of the exact experiments in this book deal with gases, the difficulties above mentioned warrant further discussion.

(1) Since most gases are collected over water, the whole system should be watched — height of water, bottle, or tube, and space occupied by the gas. The water level is an indicator of the gas level, and the eye should lead the mind to regard this water level as the essential indicator of the conditions. Again, all apparatus used in gas experiments should be rigorously tested before the experiment begins, or, better still, as the arrangement of the apparatus proceeds. If the gas is to be under increased or diminished pressure at any time, the connections should be perfect. Stoppers *must* fit and tubing *must* be the exact size. Do not use wax to stop leaky joints. Never use vaseline unless it is prescribed in the directions.

(2) Each gas has its particular drying or purifying agents, and these agents vary in efficiency. Water vapor is the most common ingredient to be removed. It is customary to pass a gas through a bottle or U-tube containing the purifying agent. The two substances used for the removal of water vapor are concentrated sulphuric acid and calcium chloride. Either may be put in a U-tube. The sulphuric acid drier is more efficient if both limbs are

nearly but loosely filled with glass beads or fragments of glass tubing and then enough acid poured into each limb to drench the glass but not seal the bend. This device exposes a large surface of acid, permits a steady current of gas, and relieves the pressure on the generator, since the gas does not have to overcome the pressure of the heavy acid in the bend of the U-tube. If the last precaution is unnecessary, then a Drechsel drying bottle (Fig. 114) is a convenient substitute for the U-tube. The Drechsel bottle may be replaced by the apparatus shown in Fig. 56. If calcium chloride is used, it should be the fused variety and be free from powder, as the latter hardens when moistened and clogs the tube. Often the U-tube may be successfully replaced by a calcium chloride tube or jar, or even an ordinary glass tube. All drying apparatus must be protected from the air by tight corks or glass plugs (Fig. 47). Ammonia is dried by passing it over lime or soda lime.

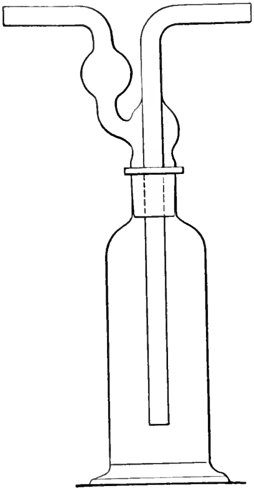

FIG. 114.—Drechsel drying (or washing) bottle.

(3) and (4) Since gases cannot be easily weighed, their weight is calculated from their volume. If the volume changes and the change in conditions is ignored or not revealed by the thermometer or barometer, the weight will apparently change, while the mass is constant. The heat of the hand or even of the body will often alter a gas volume, yet the thermometer shows no appreciable change. Hence, *tubes containing gas to be measured must never be*

*touched with the hand.* If it is necessary to move them, do so by the clamp, or with a slip of paper wound around the tube, or a piece of straw matting, or a test tube holder. The temperature of a gas is taken either by hanging a thermometer beside the gas tube, or finding the temperature of the water over which the gas stands. The latter way is more reliable. Nevertheless, errors lurk in this method. The gas must stand over the water *at least fifteen minutes* before the temperature is taken so that the whole system may be at the same temperature. It is convenient to measure gases over the water in tall jars which have been standing in the laboratory for several days before use. The thermometers may remain in the jars all the time and the tubes containing the gas may stand in the water over night, readings being taken at the close of the session or on the following morning before the regular laboratory period. Again, the reading of the gas volume and the temperature must be made simultaneously, and the thermometer must be read while the bulb is in the water. A lens is often necessary to read the exact height of the mercury, especially if the reading is estimated to the fraction of a degree.

A barometer reading taken at the beginning of the laboratory period and posted in plain sight is sufficiently accurate for several hours, unless violent atmospheric changes are in operation. *All readings should be verified,* either by the teacher or a student, and any disagreement remedied before the calculations are made.

### CLASS-ROOM EXERCISE. IV.

1. Define law, hypothesis, and theory, as the terms are used in science. Of what value is each?
2. Why should gas volumes be reduced to standard conditions?

# APPENDIX C.

## I. GLOSSARY OF CHEMICAL TERMS.

A WORKING familiarity with the language of Chemistry is indispensable to intelligent reading, speaking, and note-taking. Many of the following definitions are only suggestive; all should be interpreted by the larger and more recent text-books.

**Acid,** adj. Sour; reddens blue litmus. n. A substance which neutralizes a base, decomposes a carbonate, and reddens blue litmus.

**Acidic,** adj. Same as Acid (adj.).

**Affinity,** n. See Attraction.

**Agitate,** v. To shake; to mix by violent shaking.

**Alkali,** n. Popularly a substance which neutralizes an acid, and turns red litmus blue. A strong hydroxide. A caustic substance. A narrow use of the term Base.

**Alkaline,** adj. Having the properties of an alkali. Basic.

**Allotrope,** n. An allotropic form.

**Allotrŏp′-ic,** adj. Illustrating Allotropy.

**Allotropism, Allŏt′-ropy,** n. Variation in the physical properties assumed by some elements, especially sulphur, phosphorus, and carbon.

**Alloy′,** n. A compound or mixture of metals having metallic properties, *e.g.* brass.

**Amalgam,** n. Alloy in which mercury is a constituent, *e.g.* sodium amalgam.

**Amorphous,** adj. Non-crystalline, *e.g.* amorphous sulphur. Often applied to powders, as amorphous carbon.

**Analysis,** n. A separation; a partial or complete decomposition. The opposite of synthesis. The determination by a chemical operation of the presence of a substance, as qualitative analysis, or of the amount of a substance, as quantitative, gravimetric, or volumetric analysis.

**Anhydride,** n. An oxide which unites with water to form an acid, *e.g.* sulphurous anhydride, carbonic anhydride. Often called an acid anhydride.

**Anhydrous,** adj. Applied to a compound from which water of crystallization has been removed, *e.g.* anhydrous copper sulphate. Any "water-free" substance, as anhydrous alcohol or ether.

**Anion,** n. A negative or acidic ion.

**Anneal**, v. To render less brittle by slow cooling.

**Aqua**, n. Latin for water. *Aqua regia* means royal water.

**Aqueous**, adj. Relating to water, as aqueous vapor, aqueous tension, or an aqueous solution as distinguished from an alcoholic solution.

**Aspirate**, v. To draw or force air or any other gas through an apparatus.

**Aspirator**, n. A filter pump. A bottle to force or draw a gas through an apparatus.

**Atmospheric**, adj. Relating to the atmosphere.

**Attraction**, n. Affinity. Force which causes and maintains combination of elements.

**Bare**, adj. Same as free in "free flame." Bare flask is one not protected by a gauze or a similar covering.

**Base**, n. A substance which neutralizes an acid. A loose general term including metallic oxides and hydroxides. See Alkali and Acid.

**Basic**, adj. A synonym of alkaline. The opposite of Acid (adj.) and Acidic. Mono-, di-, tribasic, etc., indicate the number of replaceable hydrogen atoms in a molecule of an acid. Applied to a class of salts.

**Bi-**. A prefix meaning two. Same as Di-.

**Binary**, adj. Applied to a compound of two elements.

**Bump**, v. Loose term meaning sudden or violent boiling.

**Cation**, n. A positive or metallic ion.

**Caustic**, adj. Burning, corroding, *e.g.* caustic potash, lunar caustic (silver nitrate).

**Chemical**, adj. Relating to chemistry, *e.g.* chemical change. Implies deep-seated action.

n. Chemical compound; reagent; any substance used in the laboratory.

**Combustible**, adj. Inflammable, burns with flame, ignites easily.

**Combustion**, n. Chemical action accompanied by light and heat. Strictly, a union with certain elements, *e.g.* oxygen; broadly, a burning.

**Commercial**, adj. Not chemically pure; as ordinarily furnished by dealer.

**Compound**, n. Product of the combination of two or more elements in definite proportions. Often means the opposite of mixture.

**Concentrated**, adj. Strong. Opposite of Dilute.

**C. P.** Abbreviation of "chemically pure." Free from impurities.

**Crystal**, n. A solid, bounded by geometrically arranged surfaces and deposited usually from a vapor or saturated solution.

**Crystalline**, adj. Consisting of crystals or resembling a crystal. Opposite of Amorphous.

**Crystallize**, v. To change into or to form crystals.

**Crystallized**, adj. Having the form of crystals. Not powdered, or dissolved, or amorphous.

**Decant**, v. To pour a supernatant liquid from a vessel without disturbing the sediment.

**Decantation**, n. Act of decanting, as "to wash by decantation."

**Decompose**, v. To separate into simpler parts.

**Decomposition**, n. Act of decomposing.

**Deflagrate**, v. To burn suddenly.

**Deflagrating**, adj. Applied to a spoon in which deflagration occurs. Sometimes applied to a mixture, or to a single substance.

**Deflagration**, n. Rapid and sudden burning.

**Deliquesce**, v. To absorb moisture rapidly from the air.

**Deliquescence**, n. Liquefaction or dissolving due to absorption of moisture from the air.

**Deliquescent**, adj. Applied to substances which deliquesce.

**Diffusion**, n. Mixing, intermingling, especially of gases.

**Di-**. A prefix. See Bi-.

**Digest′**, v. To soften by heat and moisture; to heat in a closed vessel preparatory to chemical operations.

**Dilute**, adj. Weak. See Concentrated. v. To weaken, usually by addition of water.

**Displacement**, n. Collection of a gas by allowing it to fall or rise into a vessel. The former method is downward displacement, the latter is upward displacement. Substitution of one element for another in a compound; replacement.

**Distil**, **v.** To evaporate a liquid and recover it by subsequent condensation.

**Distillate**, n. The condensed vapor.

**Distillation**, n. Act of distilling. Destructive distillation is often applied to the preparation of coal gas, or to the manufacture of various liquids from wood.

**Ductile**, adj. May be drawn into wire.

**Dyad**, n. Element with valence of two. A bivalent element.

**Ebullition**, n. Boiling or bubbling of a liquid due to escaping gas.

**Effervescence**, n. The bubbling of a gas through a liquid. The result of chemical action, and usually occurring without the application of heat. *Not boiling. E.g.* the escape of carbon dioxide from the decomposition of a carbonate by an acid.

**Effloresce**, v. To lose easily water of crystallization.

**Efflorescence**, n. A turning to powder or loss of luster due to escape of water of crystallization at the ordinary temperature.

**Efflorescent**, adj. Applied to substances which effloresce.

**Electrode**, n. Pole or terminal of a wire conducting an electric current.

**Element**, n. A substance which has not yet been decomposed by any means known to science.

**Equivalence**, n. See Valence.

**Evaporate**, v. To convert into vapor. To heat until all liquid has been turned into vapor; *e.g.* "evaporate to dryness," means to heat until no liquid remains.

**Evolve**, v. See Liberate.

**Evolution,** n. Liberation; escape as a result of chemical action, as evolution of oxygen from potassium chlorate.

**Excess,** n. A large quantity. More than a given quantity. "A slight excess" is a little more than is necessary to produce certain results.

**Extract,** v. To remove the essential part of a mixture, usually by dissolving it in hot water, or alcohol, or a special solvent, and then filtering.

**Filter,** v. To separate a solid from a liquid; to purify.

n. Loosely applied to the filter paper, or sometimes to the whole apparatus by which filtration is accomplished.

**Filtrate,** n. The liquid which passes through the paper. Opposite of Precipitate (n.).

**Fixed,** adj. Not volatile.

**Formula,** n. A group of symbols expressing the composition of a compound.

**Formulæ,** **Formulas,** n. Plural of formula.

**Free,** adj. Not opposed. Free flame touches the object. Free channel for gases has no obstruction.

**Fuse,** v. To melt.

**Fusible,** adj. Capable of being melted.

**Gas,** n. The aeriform state of matter. Popularly, a mixture of gases, *e.g.* illuminating gas. See Vapor.

**Gelatinous,** adj. Like jelly. Applied to precipitates like aluminium hydroxide.

**Generate,** v. See Liberate and Evolve.

**Generator,** n. Apparatus in which gases are generated.

**Gram,** n. Unit of weight in the Metric System. Equal in weight to 1000 mg. and to 1 cc. of water at 4° C.

**Granulated,** adj. In small pieces varying in size from a single crystal of sugar to a hazel nut. Applied to zinc usually used to generate hydrogen.

**Gravimetric,** adj. By weight.

**Halogen,** n. and adj. A group of elements consisting of chlorine, bromine, iodine, and sometimes fluorine. Halogen means a "sea salt producer." It is applied to the members of this group because their sodium salts resemble sea salt. Chlorides, bromides, etc., are sometimes called halogen salts, or halides.

**Hard,** adj. Applied to water which contains salts of calcium or magnesium, or both.

**Hydrate,** n. Used inaccurately as a synonym of hydroxide.

**Hydrated,** adj. Combined with water, opposite of dehydrated and anhydrous. "Opal is hydrated silica."

**Hydro-.** Prefix meaning (*a*) an acid containing no oxygen, or (*b*) a compound containing hydrogen, as in the word *hydrocarbon.*

**Hydroxyl,** n. The radical OH.

**Hygroscopic,** adj. Applied to bodies which readily absorb moisture from the atmosphere, *e.g.* potassium carbonate.

**Hypo-.** A prefix meaning below. Applied to a compound in which the distinctive element has a low, and usually the lowest, valence, *e.g.* hypochlorous acid.

**-ide.** Ending of binary compounds.

**Inflammable,** adj. Applied to a vapor or gas which burns readily. The vapor of such liquids as carbon bisulphide, ether, alcohol, naphtha, and benzine is inflammable.

**Infusible,** adj. Not fusible; not easily melted.

**Ignite,** v. To set on fire, to burn, to kindle.

**Ignition,** n. The act of igniting. Used with tube, *e.g.* ignition tube, a tube which will not melt at a high temperature.

**Inorganic,** adj. Opposite of organic; relating to mineral substances, as inorganic or mineral acids. Applied to all compounds except the complex derivatives of carbon.

**Insoluble,** adj. Not soluble.

**Interact,** v. To react, to undergo chemical action, to act.

**Interaction,** n. Reaction. Especially that chemical action in which there is double decomposition. The reaction consisting of both decomposition and combination.

**Ion,** n. An electrically charged atom or atomic group. See Chapter IX. See Anion and Cation.

**Ionization,** n. The act of breaking up into ions.

**Liberate,** v. To set free, to evolve, *e.g.* to liberate hydrogen from an acid.

**Limb,** n. An arm or portion of a tube, *e.g.* limb of a U-tube.

**Liter,** n. Unit of volume in the Metric System. A cubic decimeter, equivalent to 1000 cc., a kilogram, 2.2 lb., or 1.06 quarts of water.

**Malleable,** adj. May be rolled into sheets. Compare Germ. *hämmerbar.*

**Manipulation,** n. Experimenting. "Handling" apparatus. The manual part of an experiment.

**Metal,** n. A substance, generally an element, possessing characteristic properties. They are usually hard, their smooth surfaces reflect light, they are good conductors of heat and electricity. All, except mercury, are solid at the ordinary temperature. A base-forming element. An alloy.

**Molecular,** adj. Applying to a molecule, as molecular weight. A weight proportional or equal to the molecular weight, as molecular proportions.

**Molecule,** n. A group of atoms chemically combined. The smallest particle of a compound possessing the properties of the compound.

**Molten,** adj. Melted, fluid or semi-fluid, thick, viscous. Applied to a hot mass.

**Monad,** n. Element with a valence of one. A univalent element.

**Monatomic,** adj. One atom to the molecule.

**Mono-.** A prefix meaning one. Same as Uni-.

**Nascent**, adj. Applied to an element the instant it is evolved or released from a compound. An active state.

**Neck**, n. The slender part of a vessel, *e.g.* the long stem of a retort, the part of a bottle or flask to which a stopper is fitted.

**Neutral**, adj. Indifferent to litmus, as, "water is neutral." Neither acid nor alkaline, as, "sodium chloride is neutral."

**Neutralization**, n. The interaction of an acid and base resulting in the formation of a salt.

**Neutralize**, v. To make neutral.

**Nomenclature**, n. The language of chemistry. A system of terms and terminology. The naming of elements and compounds.

**Non-metal**, n. An element not a metal. An acid-forming element, such as sulphur or nitrogen.

**Normal**, adj. Standard. Normal Pressure = 760 mm. Normal Temperature = 0° C. Normal volume is one measured at or reduced to these conditions. Salt containing no replaceable hydrogen, such as $K_2SO_4$.

**Notation**, n. The written language of chemistry, especially the symbols and formulas.

**Organic**, adj. Applied to natural or artificial products from animal or vegetable life. More specifically, complex compounds of carbon. Opposite of Inorganic.

**Oxidation**, n. Act of causing a combination with oxygen.

**Oxide**, n. Compound of oxygen with one other element.

**Oxidize**, v. To change to an oxide. To add oxygen. To increase the valence. To withdraw hydrogen.

**Oxidizing Agent.** A compound yielding oxygen easily, as nitric acid or potassium chlorate.

**Per-.** A prefix meaning beyond. Originally denoting the highest valence, as persulphuric acid — the acid in which sulphur has the highest valence. Often denoting that an element has a higher valence than in similar compounds.

**Physical**, adj. Relating to Physics, *e.g.* physical change. Implies a change of properties which can be readily seen or detected by physical means.

**Pneumatic**, adj. Relating to gases, *e.g.* pneumatic trough, a vessel for collecting gases. The Pneumatic Period of chemistry was characterized by the discovery of several gases.

**Precipitate**, v. To produce an insoluble solid by mixing solutions.

n. The solid so produced. The solid retained by the paper in filtering. Literally, a substance thrown down, because the insoluble solid often falls to the bottom of the vessel.

**Proto-.** } A prefix meaning first or
**Prot-.** } lowest. Denoting the first of a series of compounds, as protochloride, protoxide. Sometimes used synonymously with Mono.

**Qualitative**, adj. Referring only to the kind of matter.

**Quantitative**, adj. Referring to exact weight or volume.

**Radical,** n. A group of different atoms acting as a single element in a series of compounds and incapable of independent existence, as $NH_4$, $OH$, $CN$.

**React,** v. To act chemically, to interact.

**Reaction,** n. A chemical change, usually applied to a complete change. Behavior toward litmus, as acid reaction. Loosely used as synonymous with equation.

**Reagent,** n. A mixture or compound, often in solution, used to produce a desired chemical change, *e.g.* a group reagent. Often applied to any "chemical."

**Receiver,** n. Any vessel which catches or receives a distillate.

**Reduce,** v. To remove oxygen. To lower the valence. To add hydrogen. See Oxidize.

**Reducing Agent.** A compound abstracting oxygen easily, such as hydrogen or carbon.

**Reduction,** n. Act of reducing, or removing oxygen.

**Replace,** v. See Substitute.

**Residual,** adj. Whatever remains, as "residual gas."

**Residue,** n. A solid left after filtration or evaporation. A final product. A remainder.

**Salt,** n. A compound formed by the replacement of the hydrogen of an acid by the metal of a base. Popular name of sodium chloride.

**Saturated,** adj. Containing a maximum amount under given conditions. A saturated solution remains unchanged on the addition of more solid. State in which all chemical attraction is satisfied. A compound to which nothing can be chemically added, such as water.

**Seal,** v. To close. To seal a tube is to melt the open end (or ends) together. To seal a flask is to close it — usually air tight.

**Sesqui-.** A prefix meaning one and one-half. Applied to compounds containing three atoms of one element and two of another, as $Fe_2O_3$.

**Soluble,** adj. Applied to all substances which dissolve (usually in water). Opposite of Insoluble.

**Solute,** n. The solid which the liquid dissolves.

**Solvent,** n. The liquid in which the solid dissolves.

**Stable,** adj. Not easily decomposed.

**Standard.** See Normal.

**Sublimate,** n. A substance sublimed.

**Sublimation,** n. Act of distilling a solid. Usually, converting a solid into a vapor which immediately solidifies in a purer state higher up in the containing vessel.

**Substitute,** v. To replace one element or group by another.

**Substitution,** n. Replacement in a compound.

**Sulpho-.** Prefix denoting a compound of sulphur.

**Supernatant,** adj. Applied to a liquid standing over a solid, especially a precipitate. See Decant.

**Supersaturated,** adj. Applied to a solution containing more solid than it would contain if an undissolved portion were present.

**Suspension,** n. Used in the phrase "in suspension" to describe a solid distributed in fine particles through a liquid.

**Symbol,** n. The abbreviation of the name of an element.

**Synthesis,** n. Putting together. Opposite of Analysis.

**Tetra-.** A prefix meaning four.

**Tetrad,** n. Element with the valence of four. A tetravalent or quadrivalent element.

**Thio-.** Same as Sulpho-.

**Treat,** v. To add to, to mix with, to allow to react with; *e.g.* "treat with hydrochloric acid" means to add hydrochloric acid.

**Tri-.** A prefix meaning three.

**Triad,** n. Element with the valence of three. A trivalent element.

**Tubulated,** adj. Provided with a tubulure. Applied to a glass-stoppered vessel, *e.g.* a retort.

**Tubulure,** n. A short tube-like opening in the bulb of a retort or side of a bottle.

**Uni-.** Same as Mono-.

**Unstable,** adj. Easily decomposed.

**Valence,** n. The replacing or combining power of an element in terms of hydrogen.

**Vapor,** n. Gaseous form of a substance normally liquid or solid. Any gas near its condensing point. Popularly, moisture in the air. Loosely, any cloud-like substance.

**Volatile,** adj. Applied to all substances which easily change into a vapor.

**Volumetric,** adj. By volume. See Gravimetric.

**-yl.** Suffix of some radicals, *e.g.* hydroxyl.

## II. BIBLIOGRAPHY.

SUCCESSFUL experimental work is enlivened and fixed in mind by intelligent use of the best books. A chemical library should be in every laboratory; pupils should have free access to books, and should be encouraged to consult them frequently at all stages of the work. The following list contains only a portion of the books which the author regards as useful to both teacher and pupil in elementary work. The starred (*) titles indicate books which are intended for the teacher's use. In most cases the book which is accurate and simple in statement, full but not ponderous, and new, is first in its own list; such a discrimination is, of course, largely personal. The prices in some cases are subject to discount. Special books, monographs, and pamphlets are noted under separate chapters in the Teacher's Supplement.

### I. DICTIONARIES.

*1. DICTIONARY OF CHEMISTRY, Watts, Morley, and Muir. Longmans, Green, & Co. 4 vols. $65.

*2. A DICTIONARY OF CHEMICAL SOLUBILITIES, Comey. The Macmillan Co. 515 pp. $5, net.

* 3. SCIENTIFIC AMERICAN CYCLOPEDIA OF RECEIPTS. Munn & Co., New York. $5. (Contains 12,500 useful receipts.)

* 4. CHEMIKER KALENDAR, Biedermann. Springer, Berlin (annually). $1. (Contains numerous tables.)

### II. LARGE DESCRIPTIVE.

1. TEXT-BOOK OF INORGANIC CHEMISTRY, Newth. Longmans, Green, & Co. 682 pp., 146 illus. $1.75.

* 2. TREATISE ON CHEMISTRY, Roscoe and Schorlemmer. D. Appleton & Co. 2 vols. $8. (Inorganic.)

3. CHEMISTRY (ADVANCED COURSE), Remsen. Henry Holt & Co. 850 pp. $2.80. (Includes experiments.)

* 4. PRINCIPLES OF CHEMISTRY, Mendeléeff. Edited by Greenaway. Longmans, Green, & Co. 2 vols. $10.

5. GENERAL INORGANIC CHEMISTRY, Freer. Allyn & Bacon, Boston. 559 pp. $3. (Includes experiments.)

### III. SMALL DESCRIPTIVE.

1. ELEMENTS OF CHEMISTRY, Shepard. D. C. Heath & Co. 366 pp. $1.40.

2. CHEMISTRY (BRIEFER COURSE), Remsen. Henry Holt & Co. 435 pp. $1.12.

3. ELEMENTARY MANUAL OF CHEMISTRY, Storer and Lindsay. American Book Co. 453 pp. $1.20.

4. ELEMENTS OF CHEMISTRY, Williams. Ginn & Co. 397 pp. $1.10.

### IV. THEORETICAL AND PHYSICAL.

1. PHYSICAL CHEMISTRY FOR BEGINNERS, Van Deventer. Translated by Boltwood. John Wiley & Sons, New York. 154 pp. $1.50.

* 2. OUTLINES OF GENERAL CHEMISTRY, Ostwald. Translated by Walker. The Macmillan Co. 386 pp. $3.50.

* 3. OUTLINES OF THEORETICAL CHEMISTRY, Meyer. Translated by Bedson and Williams. Longmans, Green, & Co. 232 pp. $2.50.

4. PRINCIPLES OF THEORETICAL CHEMISTRY, Remsen. Lea Bros., Philadelphia. 361 pp. $2.50.

* 5. OUTLINE OF THE THEORY OF SOLUTION, Morgan. John Wiley & Sons. $1.

* 6. ELEMENTS OF PHYSICAL CHEMISTRY, Morgan. John Wiley & Sons. 299 pp. $2.

* 7. SCIENTIFIC FOUNDATIONS OF ANALYTICAL CHEMISTRY, Ostwald. The Macmillan Co. 216 pp. $2.

*8. SOLUTIONS, Ostwald. Translated by Muir. Longmans, Green, & Co. 310 pp. $2.

9. CHEMICAL THEORY FOR BEGINNERS, Dobbin and Walker. The Macmillan Co. 236 pp. $.70.

10. SCIENTIFIC MEMOIRS, J. S. Ames, Editor. American Book Co.
    - (*a*) MODERN THEORY OF SOLUTION, Jones. 134 pp. $1.
    - (*b*) FUNDAMENTAL LAWS OF ELECTROLYTIC CONDUCTION, Goodwin. 96 pp. $1.

*11. ATOMIC THEORY, Wurtz. D. Appleton & Co. 344 pp. $1.50.

*12. INTRODUCTION TO PHYSICAL CHEMISTRY, Walker. The Macmillan Co. 332 pp. $3.

Books in this class are constantly appearing. Of the above, 2, 3, and 8 are excellent, though not the latest; 1 is good, though brief; 7 and 10 are special, but exact; 12 is the latest and from many standpoints the best for a first study of the whole subject.

## V. HISTORICAL.

1. HISTORY OF CHEMISTRY, Venable. D. C. Heath & Co. 172 pp. $1.

*2. HISTORY OF CHEMISTRY, von Meyer. Translated by M'Gowan. The Macmillan Co. 544 pp. $4.50.

*3. ESSAYS IN HISTORICAL CHEMISTRY, T. E. Thorpe. The Macmillan Co. 381 pp. $2.20. (Biographies of eleven famous chemists.)

4. GASES OF THE ATMOSPHERE, Ramsay. The Macmillan Co. 240 pp. $2. (Covers pneumatic period.)

5. ALEMBIC CLUB REPRINTS, Chemical Pub. Co., Easton, Pa. $.50 each.
    1. Experiments on Magnesia Alba.
    2. Foundations of the Atomic Theory.
    3. Experiments on Air.
    4. Foundations of the Molecular Theory.
    5. Extracts from Micrographia.
    6. Decomposition of the Fixed Alkalies.
    7. Discovery of Oxygen. Part 1.
    8. Discovery of Oxygen. Part 2.
    9. Elementary Nature of Chlorine.
    10. Researches on the Arseniates, Phosphates, and Modifications of Phosphoric Acid.
    11. Increase in Weight of Tin and Lead on Calcination.
    12. Liquefaction of Gases.
    13. Early History of Chlorine.
    14. Researches on Molecular Asymmetry.

6. CENTURY SCIENCE SERIES. The Macmillan Co. $1.25 each.
   1. Humphry Davy, T. E. Thorpe.
   2. John Dalton, Roscoe.
   3. Michael Faraday, Thompson.
   4. Justus von Liebig, Shenstone.

7. HEROES OF SCIENCE — CHEMISTS, Muir. E. & J. B. Young & Co., N.Y. 350 pp. $1.50. Popular but accurate.

8. HEROES OF SCIENCE — PHYSICISTS, Garnet. (Same as 7.)

* 9. HISTORY OF THE WARFARE OF SCIENCE WITH THEOLOGY, White. D. Appleton & Co. 2 vols. $5. Vol. 1, Chap. 12, contains a valuable account of the growth of chemistry.

10. FARADAY AS A DISCOVERER, Tyndall. D. Appleton & Co. 171 pp. $1.

## VI. ORGANIC.

1. ORGANIC CHEMISTRY, Perkin & Kipping. J. B. Lippincott Co., Philadelphia. 552 pp. (2 vols.). $2.

2. ORGANIC CHEMISTRY, Remsen. D. C. Heath & Co. 364 pp. $1.30.

3. LABORATORY MANUAL OF ORGANIC CHEMISTRY, Orndorff. D. C. Heath & Co. 78 experiments. $.40. Designed to accompany 2 (supra).

* 4. THE PRACTICAL METHODS OF ORGANIC CHEMISTRY, Gattermann. Translated by Shober. The Macmillan Co. 330 pp. $1.60.

* 5. ORGANIC CHEMISTRY, Richter. Translated by E. F. Smith. P. Blakiston's Son & Co. Vol. I., 625 pp. $3. Vol. II., 671 pp. $3.

## VII. MISCELLANEOUS.

1. OUTLINES OF INDUSTRIAL CHEMISTRY, F. H. Thorp. The Macmillan Co. 528 pp. $3.50.

2. METHODS OF GLASS BLOWING, Shenstone. Longmans, Green, & Co. 96 pp. $.50.

* 3. PHYSICAL PROPERTIES OF GASES, Kimball. Houghton, Mifflin, & Co. 238 pp. $1.25.

4. MINERALS AND HOW TO STUDY THEM, E. S. Dana. John Wiley & Sons. 368 pp. $1.25.

* 5. QUESTIONS ON CHEMISTRY, Jones. The Macmillan Co. $.75. (About 1500 unanswered questions.)

6. FERMENTATION, Schützenberger. D. Appleton & Co. 331 pp. $1.50.

7. CHEMISTRY OF DAILY LIFE, Lassar-Cohn. J. B. Lippincott Co. $1.75.

8. Aluminum and Aluminum Alloys. The Pittsburg (Pa.) Reducing Co. 266 pp. $1.50.

*9. Development of the Periodic Law, Venable. Chemical Pub. Co. 321 pp. $2.50.

10. Chemistry of Common Life, Johnston. D. Appleton & Co. 592 pp. $2.

*11. Manual of Chemical Technology, Wagner. D. Appleton & Co. 568 pp. $7.50.

12. The New Chemistry, Cooke. D. Appleton & Co. 393 pp. $2.

13. The Chemistry of Cookery, Williams. D. Appleton & Co. $1.50.

14. Story of a Piece of Coal, Martin. D. Appleton & Co. 165 pp. $.40.

15. Coal and the Coal Mines, Green. Houghton, Mifflin, & Co. 240 pp. $.75.

16. Inorganic Chemical Preparations, F. H. Thorp. Ginn & Co. 238 pp. $1.50.

17. Laboratory Practice, Cooke. D. Appleton & Co. 192 pp. $1.

18. Liquefaction of Gases, Hardin. The Macmillan Co. 244 pp. $1.50.

## III. APPARATUS AND CHEMICALS.

Various specific directions regarding laboratory supplies are given in the Teacher's Supplement, in the list of Reagents, and in the directions for the experiments. The subjoined lists are inserted simply to assist in the preparation of an order. Quantities and prices have been intentionally omitted, since it is believed that both dealer and purchaser will make more felicitous arrangements, if both are unrestricted by a list unavoidably incomplete. No teacher can judge for another without knowing the working conditions, and no author has a right to insist that a dealer shall always furnish an article for a fixed sum. The market price fluctuates, and prices vary with quantity and quality as well as with the method of ordering. The author, at his own suggestion, has lodged with the L. E. Knott Apparatus Co., 16 Ashburton Place, Boston, Mass., information regarding the quantities of apparatus and chemicals used by his own classes. It is hoped that teachers will avail themselves of the opportunity to correspond with both author and dealer in the preparation of order lists.

### I. INDIVIDUAL APPARATUS.

This list includes the apparatus constantly used by a single student, who should be provided with each piece, and be held responsible for its return in usable condition. The apparatus, as itemized below, should not cost more than $6.75.

12 Test tubes, 6 x $\frac{3}{4}$.
2 Test tubes, 8 x 1.
1 Test tube holder.
1 Test tube rack.
1 Test tube brush.
1 Bunsen burner.
3 Feet rubber tubing, $\frac{1}{4}$ in. int. diam.
1 Wing-top burner.
1 Blowpipe.
1 Blowpipe tube.
1 Horn spoon, 5 in.
100 Filter papers, 4 in.
5 Bottles, salt mouth, 250 cc. (8 oz.).
1 Funnel, $2\frac{1}{2}$ in.
1 Safety tube.
1 Flask, 500 cc.
1 Two-hole rubber stopper, to fit above.
1 One-hole rubber stopper, to fit above.
1 Evaporating dish, 3 in.
1 Crucible with cover, No. 0.
1 Crucible block.
1 Casserole, 125 cc.
1 Pair iron forceps.
1 Triangular file.
1 Mortar and pestle, 3 in.
1 Ignition tube, 4 in. × $\frac{1}{4}$ in.
1 Deflagrating spoon.
1 Towel.

## II. SPECIAL APPARATUS.

This list includes the apparatus used occasionally, but of such a nature that it should be given only when demanded, and returned as soon as used. The only items costing over $2.00 in this list are the burettes (if with glass stop-cock) and the platinum crucible. The latter need not cost over $8.00 (see Exp. 140). In some cases cheaper forms than those called for may answer, but in the author's judgment it is imprudent to economize on permanent equipment.

- Beads.
- Beakers, 2.
- Bottle, Drechsel.
- Burettes.
- Calcium chloride tube.
- Chlorine decomposition tube.
- Condenser.
- Crucible, Hessian.
- Dish, lead.
  - glass (shallow).
- Eudiometer, 100 cc.
- Flask, Erlenmeyer, 250 cc and 125 cc.
- Funnel, dropping
- Gas tube, 100 cc. and 50 cc.
- Marchard tube.
- Pinchcock, Mohr.
  - Hofmann
- Pipe, iron.
- Pipette.
- Platinum crucible.
  - foil.
  - wire.
- Thermometer.
- Thistle tube top.
- Trap, bulb (for Exp. 92).
  - steam.
- U-tube.

### III. GENERAL APPARATUS.

This list includes all apparatus intended for general use. Items costing \$10 or over are marked with a *.

*Air bath.
*    pump.
Asbestos boards, 6 x 6 in.
Aspirator bottle.
Balance, horn pan.
Barometer.
Battery.
Blast lamp.
Bottles, wide mouth, 2 oz. and 4 oz.
Candles.
Chimneys, student lamp.
*Combustion furnace, 10 burner.
Corks, assorted.
Cork borers, set.
Cylinders, graduated (set) 7.
Dowel rod.
Electrolysis apparatus, $10\frac{1}{2}$ in.
Emery paper.
File, round.
Filter pump.
*Gas holder.
Gauze, iron, 6 x 6 in.
Glass plates, 6 x 6 in.
    tubing, soft.
    tubing, hard.
    rod.
Graduates. See Cylinders.
Induction coil.
Iron stands, with 3 rings and 2 clamps.
Jars, battery, 6 x 8 in.
    tall, 3 x 15 in.
    waste.
Labels (Dennison, 223, 205, 219, 201).
Magnets.
Magnifying glass.
Meter stick.
Pneumatic troughs.
Rule, 30 cm.
Sand bath pans, 4 in.
Scales, trip.
Scissors, 6 in.
Shears, 12 in.
* Spectroscope.
Stoppers, rubber, with 1 and 2 holes.
Tapers, box.
Thread, cotton.
Triangle, porcelain.
Tripods.
Tubing, rubber (ordinary and pressure), $\frac{1}{4}$, $\frac{3}{16}$, $\frac{1}{8}$ in. (int. diam.).
Water baths, 5 in.
Weights (for horn pan and trip scales).
Wire cutter.
Wood (splints).
Wooden blocks, 6 x 6 x 1 in.
    6 x 6 x $\frac{3}{4}$ in.
    4 x 4 x $\frac{1}{2}$ in.
Woulff bottle, 2-neck.

### IV. CHEMICALS.

This list includes all chemicals needed for this book.

Acid, acetic.
    citric.
    hydrochloric
    nitric.
Acid, oxalic.
    sulphuric.
    tartaric.

Alcohol, ethyl.
    methyl.
Alum, chrome.
    potassium.
Aluminium, metal.
    sulphate.
Ammonium, chloride.
    hydroxide.
    nitrate.
    oxalate.
    sulphide.
Antimony, metal.
Arsenious oxide.
Asbestos, shredded.

Barium chloride.
    hydroxide.
Beeswax.
Bleaching powder.
Borax (powd.).
Brass wire, No. 24.

Calcite (lump).
Calcium carbide.
    carbonate (marble).
    chloride.
    fluoride.
    oxide (lime).
    sulphate.
Carbon bisulphide.
Charcoal, animal (powd.).
    lump.
    wood (powd.).
Coal, soft.
Cobalt nitrate.
Cochineal.
Coin (silver).
Copper borings.
    nitrate.
    oxide (granulated).
    sheet.
    sulphate (cryst.).
    sulphate (dehydrated).
    wire.
Cotton (absorbent).
Cream of tartar.

Ether.

Galena.
Gelatine.
Glass wool.
Glycerine.
Gold leaf (book).

Hematite.

Indigo
Iodine.
Iron, by hydrogen.
    chloride (*ic*).
    filings.
    pyrites.
    sulphate (*ous*).
    sulphide (*ous*).

Kerosene.

Lead acetate.
    carbonate.
    dioxide (peroxide).
    nitrate.
    monoxide (litharge).
    sheet.
    tea.
    tetroxide.
Limonite.
Litmus, cubes.
    paper.

Magnesium carbonate (magnesite).
    oxide.
    powder.
    ribbon.
    sulphate.
Manganese dioxide.
    sulphate.

- Mercury.
- Mercuric chloride.
  - nitrate.
  - oxide.
  - sulphide (cinnabar).
- Mercurous nitrate.
- Mustard.

- Paraffine.
- Phenolphthalein.
- Phosphorus, yellow (worm-shaped).
- Picture cord (iron).
- Potassium, metal.
  - bromide.
  - carbonate.
  - chlorate (cryst.).
  - chlorate (powd.).
  - chloride.
  - chromate.
  - dichromate.
  - ferricyanide.
  - ferrocyanide.
  - hydroxide.
  - iodide.
  - nitrate.
  - permanganate.
  - sulphate.
  - sulphocyanide.
- Sand.
- Selenite (gypsum, cryst.).
- Shellac.
- Siderite.
- Silver nitrate.
- Soap.
- Solder.
- Soda lime.
- Sodium, metal.
  - acetate.
  - bicarbonate.
  - carbonate.
  - chloride.
  - hydroxide.
  - hyposulphite (thiosulphate).
  - nitrate.
  - nitrite.
  - phosphate (disodium phosphate).
  - silicate.
  - sulphate.
- Stannous chloride (tin crystals).
- Starch.
- Sugar, cane.
  - grape.
- Sulphur, flowers.
  - roll.

- Tartar emetic.
- Tin, granulated.
- Turpentine.

- Vaseline.
- Vinegar.

- Water, distilled.
- Wood ashes.

- Zinc, dust.
  - granulated.
  - oxide.
  - pure.
  - sheet.
  - sulphate.

## IV. REAGENTS.

THE following apparatus is helpful in the preparation of reagents: —

1 Porcelain dish, 14 in.
1 Funnel, 7 in.
1 Graduate, 1000 cc.
1 Graduate, 500 cc.
1 Iron dish (frying dish).
1 Tin dish (8 in. sauce pan).

1 Mortar, 7 in.
1 Set of hydrometers.
Glass rod (1 cm. in diam.).
Glass wool.
Asbestos, shredded.
Filter paper, 12 in.
Filter pump, flask and Witt plate.
Lithia bottles.
Glass stoppered bottles (5 pint acid).

The preparation of large quantities of a solution is hastened by using a dissolver. One is described in the *American Chemical Journal*, Vol. 22, No. 5 (November, 1899), p. 407.

The following list of reagents includes all those required for this book. A few others have been inserted for completeness. Reagents not mentioned are either unimportant or are fully explained in the experiments requiring their use. If a reagent is not mentioned in this list, consult the index.

**Acetic Acid,** $C_2H_4O_2$. — (1) Use the commercial acid (sp. gr. 1.04).

(2) Add 1000 cc. of glacial acetic acid to 2500 cc. of water.

**Alcohol,** $C_2H_6O$. — Use the commercial liquid (sp. gr. .815; contains 95 per cent of alcohol). See Appendix C, Table X.

**Alum.** — *Solid.* Use the commercial crystals.

*Solution.* Dissolve 100 gm. of the commercial alum in 1000 cc. of water. This proportion answers for any alum.

**Aluminium Chloride,** $AlCl_3$. — (1) Dissolve 100 gm. of the commercial salt in 1000 cc. of water.

(2) Precipitate aluminium hydroxide by adding ammonium hydroxide to a solution of alum. Wash the precipitate free from acid, and add a slight excess of concentrated hydrochloric acid.

**Ammonium Carbonate,** $(NH_4)_2CO_3$. — Dissolve 250 gm. of the commercial salt in 1000 cc. of water, *without heating*, and add 100 cc. of commercial ammonium hydroxide. Filter. Keep tightly stoppered.

**Ammonium Chloride,** $NH_4Cl$. — *Solid.* Use the commercial salt.

*Solution.* Dissolve 100 gm. of the commercial salt in 1000 cc. of water. Filter.

**Ammonium Hydroxide,** $NH_4OH$. — *Concentrated.* Use the commercial liquid (26° or sp. gr. = .96).

*Dilute.* Add three volumes of water to one volume of commercial ammonium hydroxide.

**Ammonium Molybdate,** $(NH_4)_2MoO_4$. — (1) Add 100 gm. of molybdenum oxide ($MoO_3$) to 400 cc. of commercial ammonium hydroxide, and stir until dissolved. Cool, and *add all at once* 1000 cc. of nitric acid (sp. gr. 1.2), using no heat.

(2) Dissolve 150 gm. of the salt in 1000 cc. of water and *pour this solution* into 1000 cc. of nitric acid (1 vol. acid to 1 vol. water).

(3) Dissolve 60 gm. of the dry salt in 400 cc. of commercial ammonium hydroxide, add 400 cc. of

water, and then add slowly 500 cc. of commercial nitric acid.

**Ammonium Oxalate,** $(NH_4)_2C_2O_4$. — Dissolve 40 gm. of the crystallized commercial salt in 1000 cc. of water. Filter. Should be made as needed.

**Ammonium Sulphide,** $(NH_4)_2S$. — (1) Add 500 cc. of the commercial solution to 500 cc. of water.

(2) Pass purified hydrogen sulphide gas into 600 cc. of *dilute* ammonium hydroxide until the solution is saturated, and then add 400 cc. of dilute ammonium hydroxide. The saturation point may be found either by the intense odor of the solution, or by the fact that the saturated solution gives no precipitate when heated gently with magnesium sulphate. Ammonium sulphide when exposed to the air soon forms the *yellow ammonium sulphide*. The latter may be prepared by dissolving a little flowers of sulphur in ammonium sulphide.

**Ammonium Sulphide** (Yellow), $(NH_4)_2S_x$. — Saturate 1500 cc. of ammonium hydroxide (sp. gr. .90) with hydrogen sulphide gas. Add 1100 cc. of ammonium hydroxide (sp. gr. .90) and 2500 cc. of water. Dissolve in the above from 50 to 75 gm. of flowers of sulphur. Filter.

**Ammonium Sulphocyanide** (or **Thiocyanate**), $NH_4SCN$. — Dissolve 10 gm. of the commercial salt in 1000 cc. of water.

**Barium Chloride,** $BaCl_2$. — *Solid.* Use the crystallized commercial salt.

*Solution.* Dissolve 50 gm. of the commercial salt in 1000 cc. of water free from sulphuric acid and sulphates. Filter if not clear.

**Barium Hydroxide,** $Ba(OH)_2$. — Dissolve 50 gm. of the commercial compound in 1000 cc. of hot water free from carbon dioxide and carbonates. Let the solution stand a day, and then filter if not clear. Keep tightly stoppered.

**Bromine Water.** — Shake 500 cc. of water with a few drops of bromine. Keep tightly stoppered and stand in the hood, where the bottle cannot be accidentally upset. A saturated solution contains about 40 gm. to the liter.

**Calcium Chloride,** $CaCl_2$. — *Solid.* For ordinary work use the commercial salt; for quantitative work use lumps of the fused variety. Keep tightly stoppered, or preserve in a quick-sealing jar.

*Solution.* Dissolve 100 gm. of the commercial salt in 1000 cc. of water. Let the solution stand a day, and then filter if not clear.

**Calcium Sulphate,** $CaSO_4$. — *Solid.* For water of crystallization use crystallized gypsum (selenite). For other work use the powdered variety.

*Solution.* Make a saturated solution by allowing 1000 cc. of cold water to stand over 10 gm. of the powdered salt. Shake often, and then siphon off the clear supernatant liquid.

**Chlorine Water.** — (1) Pass chlorine

into water until the solution smells strongly of the gas.

(2) Slowly add concentrated hydrochloric acid to a few grams of crystallized potassium chlorate, and then add 500 cc. of water. Keep tightly stoppered and in a dark place. This solution does not keep well, and should be made as needed.

**Chrome Alum.** — See Alum.

**Cobalt Nitrate,** $Co(NO_3)_2$. — Dissolve 10 gm. of the commercial salt in 1000 cc. of water. Filter if not clear.

**Cobalt Chloride,** $CoCl_2$. — Dissolve 50 gm. of the commercial salt in 1000 cc. of water. Filter if not clear.

**Cochineal.** — (1) Grind a few pieces of cochineal, or a little powder, in water, and dilute with water to the desired tint.

(2) Digest 5 gm. of cochineal powder with 50 cc. of alcohol (25 per cent). (See Table X., Appendix C.)

**Copper Sulphate,** $CuSO_4$. — Dissolve 100 gm. of the crystallized commercial salt in 1000 cc. of hot water.

**Disodium Phosphate,** $HNa_2PO_4$. — Dissolve 100 gm. of the crystallized commercial salt in 1000 cc. of water.

**Ferric Chloride,** $FeCl_3$. — (1) Dissolve 50 gm. of the commercial salt in 1000 cc. of water.

(2) Boil ferrous sulphate solution with nitric acid, add ammonium hydroxide, filter, wash the precipitated ferric hydroxide free from acid and alkali, and dissolve in concentrated hydrochloric acid.

**Ferrous Sulphate,** $FeSO_4$. — Dissolve 100 gm. of the clean, fresh, or freshly washed, commercial salt in 1000 cc. of cold water. The solution should be made as needed. If necessary to preserve it, keep iron wire in the bottle, and add occasionally (once a day) a drop or two of concentrated sulphuric acid.

**Fusion Mixture.** — Mix four parts of anhydrous sodium carbonate and five parts of dry potassium carbonate.

**Hydrochloric Acid,** $HCl$. — *Concentrated.* Use the commercial acid (sp. gr. 1.2).

*Dilute.* Add one volume of commercial acid to four volumes of water.

**Hydrogen Peroxide,** $H_2O_2$. — Use the commercial solution. Keep in a cool place, and occasionally loosen the cork for an instant.

**Indigo.** — *Solid.* Use the commercial substance.

*Solution.* (1) Dissolve a little commercial indigo paste in 1000 cc. of water.

(2) Slowly add 10 gm. of powdered commercial indigo to 25 cc. of concentrated sulphuric acid. Let the mixture stand a day, then add it slowly, with constant stirring, to 1000 cc. of water.

**Iodine,** I. — *Solid.* Use the commercial crystals.

*Solution.* — (1) Grind to complete solution in a mortar 10 cc. of water, 12 gm. of commercial iodine, and 20 gm. of potassium iodide, and then add the solution to 1000 cc. of water.

(2) Use the commercial solution.

**Lead Acetate,** $Pb(C_2H_3O_2)_2$. — Dissolve 100 gm. of the commercial salt in 1000 cc. of water. Let the solution stand a day, then filter if not clear.

**Lead Nitrate,** $Pb(NO_3)_2$. — Dissolve 100 gm. of the commercial salt in 1000 cc. of water. Let the solution stand until clear, then decant or filter.

**Lime Water,** $Ca(OH)_2$.— Slake quicklime in an iron or tin dish, add water enough to make "milk of lime," pour into a stoppered bottle, shake occasionally for several days, then let stand, and siphon off the clear supernatant liquid. Several portions of water may be added to the same solid.

**Litmus.** — *Solid.* Use the commercial blocks.

*Solution.* (1) Grind a few blocks of litmus, with a little water, in a mortar, and add the paste to enough water to produce the desired color.

(2) Grind 10 gm. of litmus blocks, with hot distilled water, in a mortar, and, after the mass has stood for fifteen minutes, pour off the liquid. Extract once or twice more with hot distilled water, dilute the liquid with about 150 cc. of distilled water, and keep in a *loosely* stoppered bottle. A few drops of chloroform occasionally added will prevent the formation of mould.

**Magnesium Sulphate,** $MgSO_4$. — Dissolve 100 gm. of the commercial salt in 1000 cc. of water.

**Magnesium Chloride,** $MgCl_2$. — Dissolve 100 gm. of the commercial salt in 1000 cc. of water.

**Manganous Chloride,** $MnCl_2$. — Dissolve 100 gm. of the commercial salt in 1000 cc. of water.

**Mercuric Chloride,** $HgCl_2$. — Dissolve 50 gm. of the commercial salt in 1000 cc. of warm water. POISON.

**Mercurous Nitrate,** $Hg_2(NO_3)_2$. — (1) Dissolve 50 gm. of the crystallized salt in 1000 cc. of distilled water, add 75 cc. of concentrated nitric acid and a little mercury.

(2) Dissolve metallic mercury moderately warm concentrated nitric acid, taking care to leave a little mercury undissolved.

**Nessler's Reagent.** — Dissolve 3.5 gm. of potassium iodide in 10 cc. of water; dissolve 1.6 gm. of mercuric chloride in 30 cc. of water. Add the latter to the former gradually and with constant stirring until the precipitate ceases to be redissolved. Then add a solution of 6.5 gm. of potassium hydroxide in 60 cc. of water. Filter. Keep tightly stoppered.

**Nitric Acid,** $HNO_3$. — *Concentrated.* Use the commercial acid (sp. gr. 1.42).

*Dilute.* Add one volume of commercial acid to four volumes of water.

**Phenolphthalein.** — (1) Dissolve 1 gm. of the solid in 100 cc. of alcohol (50 per cent.) (See Table X., Appendix C.)

(2) Dissolve 1 gm. of the solid in 100 cc. of alcohol (95 per cent). Dilute with distilled water until a precipitate forms; then add just enough alcohol to redissolve the precipitate.

(3) Use the commercial solution.

**Platinic Chloride,** $PtCl_4$. — (1) Buy a solution of desired or known strength.

(2) Dissolve scrap platinum in *aqua regia*, evaporate to pasty dryness, and dissolve the residue in distilled water.

**Potassium Bromide,** $KBr$. — Dissolve 50 gm. of the commercial crystals in 1000 cc. of water.

**Potassium Carbonate,** $K_2CO_3$. — *Solid.* Use the dry commercial salt, which should always be kept tightly stoppered, or preserved in a quick-sealing jar.

*Solution.* Dissolve 100 gm. of the dry commercial salt in 1000 cc. of water.

**Potassium Chloride,** $KCl$. — Dissolve 50 gm. of the commercial salt in 1000 cc. of water. Filter if not clear.

**Potassium Chromate,** $K_2CrO_4$. — Dissolve 100 gm. of the commercial salt in 1000 cc. of water. Filter if not clear.

**Potassium Cyanide,** $KCN$. — Dissolve 100 gm. of the commercial salt in 1000 cc. of water. This solution is a *violent poison*, and should be prepared in small quantities, plainly labelled, and kept where it cannot be promiscuously handled.

**Potassium Dichromate** (or **Bichromate**), $K_2Cr_2O_7$. — Dissolve 50 gm. of the powdered commercial salt in 1000 cc. of water.

**Potassium Ferricyanide,** $K_3Fe(CN)_6$. — Dissolve 75 gm. of the commercial salt in 1000 cc. of water. Keep tightly stoppered and in a dark place.

**Potassium Ferrocyanide,** $K_4Fe(CN)_6$. — Dissolve 75 gm. of the commercial salt in 1000 cc. of water.

**Potassium Hydroxide,** $KOH$. — *Solid.* Use the stick form.

*Solution.* Dissolve 100 gm. of the commercial lump potash in 1000 cc. of water. Use an iron dish, and stir constantly. If the solution is not clear, filter through glass wool or shredded asbestos. Keep tightly stoppered. For absorption of carbon dioxide in quantitative work, use a solution containing 500 gm. to the liter (sp. gr. = 1.27 approximately).

**Potassium Iodide,** $KI$. — *Solid.* Use the commercial crystals.

*Solution.* Dissolve 50 gm. of the commercial salt in 1000 cc. of water.

**Potassium Nitrate,** $KNO_3$. — Dissolve 100 gm. of the commercial salt in 1000 cc. of water. Filter.

**Potassium Permanganate,** $KMnO_4$. — Dissolve 5 gm. of the commercial crystals in 1000 cc. of water.

**Potassium Sulphate,** $K_2SO_4$. — Dissolve 100 gm. of the commercial salt in 1000 cc. of water.

**Potassium Sulphocyanide** (or **Thiocyanate**), $KSCN$. — Dissolve 10 gm. of the commercial salt in 1000 cc. of water.

**Silver Nitrate**, $AgNO_3$. — Dissolve 50 gm. of the commercial crystals in 1000 cc. of distilled water. Let the solution stand a day, and if not clear, filter through glass wool, or better, siphon off the clear supernatant liquid. This solution should be kept free from dust and organic matter.

**Soap Bubble Solution.** — Dissolve 10 gm. of thin shavings of castile soap in 100 cc. of distilled water. Shake until solution is complete, and filter if not clear. Add 50 cc. of glycerine.

**Sodium Amalgam.** — See Exp. 63.

**Sodium Carbonate**, $Na_2CO_3$. — *Solid.* Use the commercial powder or crystals as directions require.

*Solution.* Dissolve 100 gm. of the crystallized commercial compound in 1000 cc. of water. A saturated solution is made by dissolving about 430 gm. in 1000 cc. of water.

**Sodium Chloride**, $NaCl$. — Dissolve 100 gm. of the commercial salt in 1000 cc. of water. Filter.

**Sodium Hydroxide**, $NaOH$. — *Solid.* Use the stick form.

*Solution.* Dissolve 100 gm. of the commercial lump caustic soda in 1000 cc. of water. Use an iron dish, and add the solid slowly to the liquid with constant stirring. Filter through glass wool or shredded asbestos. Keep tightly stoppered.

**Sodium Sulphite**, $Na_2SO_3$. — Dissolve 200 gm. of the crystallized commercial salt in 1000 cc. of water.

**Stannous Chloride**, $SnCl_2$. — (1) Use the commercial solution.

(2) Dissolve granulated tin in an excess of hot concentrated hydrochloric acid until the evolution of hydrogen ceases. Add 5 volumes of water to the solution.

(3) Dissolve 500 gm. of the commercial salt in 1000 cc. of hot concentrated hydrochloric acid. Keep a few pieces of tin in the solution.

(4) Dissolve 112 gm. of the commercial salt in 200 cc. of hydrochloric acid (1 to 1) and add 800 cc. of water. Keep a few pieces of tin in the solution.

Stannous chloride solution does not keep unchanged, and should be made as needed.

**Starch Paste.** — Grind 10 gm. of starch to a paste with a little cold water in a mortar. Add the paste to 200 cc. of water. Boil. Use when cold.

**Sulphuric Acid**, $H_2SO_4$. — *Concentrated.* Use the commercial acid (sp. gr. 1.84).

*Dilute.* Slowly pour one volume of commercial acid into four volumes of water. Stir constantly during the operation. Cool and filter, if the commercial acid is not colorless.

**Tartar Emetic**, $K(SbO)C_4H_4O_6$. — Dissolve 100 gm. of the commercial salt in 1000 cc. of water.

**Zinc Chloride**, $ZnCl_2$. — Dissolve 50 gm. of the commercial salt in 1000 cc. of water.

**Zinc Sulphate**, $ZnSO_4$. — Dissolve 100 gm. of the commercial salt in 1000 cc. of water. Filter if the solution is not clear.

## V. TABLES.

### TABLE I.—IMPORTANT ELEMENTS AND THEIR ATOMIC WEIGHTS.

| NAME. | SYMBOL. | ATOMIC WEIGHT. | | | | |
|---|---|---|---|---|---|---|
| | | Clarke. | | Richards. | German. | Approximate. |
| | | H = 1.000 | O = 16 | O = 16 | O = 16 | O = 16 |
| Aluminium | Al | 26.9 | 27.1 | 27.1 | 27.1 | 27 |
| Antimony | Sb | 119.5 | 120.4 | 120.0 | 120 | 120 |
| Arsenic | As | 74.45 | 75 | 75.0 | 75 | 75 |
| Barium | Ba | 136.4 | 137.4 | 137.43 | 137.4 | 137 |
| Bismuth | Bi | 206.5 | 208.1 | 208 | 208.5 | 208 |
| Boron | B | 10.9 | 11 | 11.0 | 11 | 11 |
| Bromine | Br | 79.34 | 79.95 | 79.955 | 79.96 | 80 |
| Cadmium | Cd | 111.55 | 112.4 | 112.3 | 112 | 112 |
| Calcium | Ca | 39.8 | 40.1 | 40.1 | 40 | 40 |
| Carbon | C | 11.9 | 12 | 12.001 | 12 | 12 |
| Chlorine | Cl | 35.18 | 35.45 | 35.455 | 35.45 | 35.5 |
| Chromium | Cr | 51.7 | 52.1 | 52.14 | 52.1 | 52 |
| Cobalt | Co | 58.55 | 59 | 59.00 | 59 | 59 |
| Copper | Cu | 63.1 | 63.6 | 63.60 | 63.6 | 63.5 |
| Fluorine | F | 18.9 | 19.05 | 19.05 | 19 | 19 |
| Gold | Au | 195.7 | 197.2 | 197.3 | 197.2 | 197 |
| Hydrogen | H | 1 | 1.008 | 1.0075 | 1.01 | 1 |
| Iodine | I | 125.89 | 126.85 | 126.85 | 126.85 | 127 |
| Iron | Fe | 55.6 | 56 | 55.9 | 56 | 56 |
| Lead | Pb | 205.36 | 206.92 | 206.92 | 206.9 | 207 |
| Lithium | Li | 6.97 | 7.03 | 7.03 | 7.03 | 7 |
| Magnesium | Mg | 24.1 | 24.3 | 24.36 | 24.36 | 24 |
| Manganese | Mn | 54.6 | 55 | 55.02 | 55 | 55 |
| Mercury | Hg | 198.50 | 200 | 200.0 | 200.3 | 200 |
| Nickel | Ni | 58.25 | 58.70 | 58.70 | 58.7 | 58.5 |

## TABLE I.—IMPORTANT ELEMENTS AND THEIR ATOMIC WEIGHTS—(*Continued*).

| Name. | Symbol. | Atomic Weight. | | | | |
|---|---|---|---|---|---|---|
| | | Clarke. | | Richards | German. | Approximate. |
| | | H = 1.000 | O = 16 | O = 16 | O = 16 | O = 16 |
| Nitrogen . . . | N | 13.93 | 14.04 | 14.045 | 14.04 | 14 |
| Oxygen . . . | O | 15.88 | 16 | 16.000 | 16 | 16 |
| Phosphorus . . | P | 30.75 | 31 | 31.0 | 31 | 31 |
| Platinum . . . | Pt | 193.4 | 194.9 | 195.2 | 194.8 | 195 |
| Potassium . . | K | 38.82 | 39.11 | 39.14 | 39.15 | 39 |
| Selenium . . . | Se | 78.6 | 79.2 | 79.2 | 79.1 | 79 |
| Silicon . . . | Si | 28.2 | 28.4 | 28.4 | 28.4 | 28 |
| Silver . . . . | Ag | 107.11 | 107.92 | 107.93 | 107.93 | 108 |
| Sodium . . . | Na | 22.88 | 23.05 | 23.05 | 23.05 | 23 |
| Strontium . . | Sr | 86.95 | 87.60 | 87.68 | 87.6 | 87.5 |
| Sulphur . . . | S | 31.83 | 32.07 | 32.065 | 32.06 | 32 |
| Tellurium . . | Te | 126.5 | 127.5? | 127.5? | 127 | 127 |
| Tin . . . . . | Sn | 118.1 | 119 | 119.0 | 118.5 | 119 |
| Zinc . . . . | Zn | 64.9 | 65.4 | 65.40 | 65.4 | 65 |

The above list is taken from the *Journal of the American Chemical Society*, Vol. XXII., No. 2 (February, 1900). Richards's list was revised by a private communication in the final proof. The **approximate** *list is chosen to facilitate calculation, and should be used in solving all problems in this book.*

## TABLE II.—COLORS OF COATINGS ON CHARCOAL.

| Metal. | Hot. | Cold. |
|---|---|---|
| Lead | Lemon-yellow | Lemon-yellow |
| Tin | Yellow (faint) | White |
| Zinc | Yellow | White |

## TABLE III.—COLORS OF BORAX BEADS.

| Metal. | Oxidizing Flame. | | Reducing Flame. | |
|---|---|---|---|---|
| | Hot. | Cold. | Hot. | Cold. |
| Chromium | Reddish-Yellow | Yellowish-Green | Green | Green |
| Cobalt | Blue | Blue | Blue | Blue |
| Copper | Green | Greenish-Blue | Colorless | Red |
| Manganese | Violet | Violet | Colorless | Colorless |

## TABLE IV.—COLORS OF RESIDUES MOISTENED WITH COBALT NITRATE.

| Metal. | Color (best seen when cold). |
|---|---|
| Aluminium | Blue |
| Magnesium | Pink |
| Zinc | Green |

## TABLE V.—COLORS OF FLAMES.

| Metal. | Color of Bunsen Flame. |
|---|---|
| Barium | Green |
| Calcium | Yellowish-red |
| Potassium | Violet |
| Sodium | Yellow |
| Strontium | Crimson |

## TABLE VI.—CONVERSION OF THERMOMETRIC READINGS.

| Fahrenheit to Centigrade. | Centigrade to Fahrenheit. |
|---|---|
| $C. = \frac{F. - 32}{1.8}$ | $F. = (C. \times 1.8) + 32$ |
| or | or |
| $C. = \frac{5\,(F. - 32)}{9}$ | $F. = \frac{9\,C.}{5} + 32$ |

## TABLE VII.—SPECIFIC GRAVITY AND MELTING POINT (APPROXIMATE) OF METALS.

| Metal. | Specific Gravity. | Melting Point. | |
|---|---|---|---|
| | | Centigrade. | Fahrenheit. |
| Aluminium | 2.6 | 625 | 1157 |
| Copper | 8.9 | 1054 | 1929 |
| Iron (wrought) | 7.8 | 1600 | 2912 |
| Lead | 11.35 | 326 | 619 |
| Mercury | 13.59 | − 38.5 | − 37.3 |
| Platinum | 21.5 | 1775 | 3227 |
| Silver | 10.53 | 954 | 1749 |
| Tin | 7.3 | 232.7 | 450.9 |
| Zinc | 6.86–7.2 | 433 | 811.4 |

## TABLE VIII.—SOLUBILITY OF SALTS IN WATER.

| Salt. | Grams soluble in 100 gm. at | |
|---|---|---|
| | 20° C | 100° C. |
| Copper sulphate (cryst.) | 42.31 | 203.32 |
| Potassium chlorate | 7.2 | 59.5 |
| Potassium chloride | 35 | 57 |
| Potassium dichromate | 13 | 102 |
| Potassium nitrate | 31.7 | 246 |
| Potassium sulphate | 10.6 | 26 |
| Sodium chloride | 36 | 39.7 |
| Sodium sulphate (anhyd.) | 20 | 43 |
| Sodium sulphate (cryst.) | 58.35 | 212.47 |

## TABLE IX.—WEIGHT (IN GRAMS) OF A LITER OF DRY GASES AT 0° C. AND 760 MM.

| | | | |
|---|---|---|---|
| Air | 1.293 | Hydrogen | .0896 |
| Ammonia | .77 | Hydrogen sulphide | 1.542 |
| Carbon dioxide | 1.977 | Oxygen | 1.429 |
| Carbon monoxide | 1.251 | Nitrogen | 1.256 |
| Chlorine | 3.18 | Steam | .806 |
| Hydrochloric acid | 1.61 | Sulphur dioxide | 2.869 |

## TABLE X. — FORMULA FOR PREPARING AN ALCOHOLIC SOLUTION OF DESIRED STRENGTH.

$$x = \frac{ac}{b}$$

$x$ = vol. required.
$a$ = % of solution used.
$b$ = % of solution desired.
$c$ = vol. of stock solution used.

## TABLE XI. — COMPOSITION OF TYPICAL COALS.

| Kind. | C. | H. | O. | N. |
|---|---|---|---|---|
| Wood . . . . . . . . . . . | 50 | 6 | 43 | 1 |
| Peat . . . . . . . . . . . | 59 | 7 | 34 | 0 |
| Lignite . . . . . . . . . . | 69 | 5.2 | 25 | .8 |
| Bituminous . . . . . . . . | 82 | 5 | 12.2 | .8 |
| Anthracite . . . . . . . . | 95 | 2.5 | 2.5 | trace |

## TABLE XII. — PERIODIC ARRANGEMENT OF THE ELEMENTS, ACCORDING TO MENDELÉEFF.

Revised by F. W. Clarke. On the Basis of O = 16.

| Series. | Group I. $R_2O$. | Group II. RO. | Group III. $R_2O_3$. | Group IV. $RH_4$, $RO_2$. |
|---|---|---|---|---|
| 1 | H = 1.008 | — | — | — |
| 2 | Li = 7.03 | Gl = 9.1 | B = 11 | G = 12 |
| 3 | Na = 23.05 | Mg = 24.3 | Al = 27.1 | Si = 28.4 |
| 4 | K = 39.11 | Ca = 40 | Sc = 44.1 | Ti = 48.1 |
| 5 | (Cu = 63.6) | Zn = 65.4 | Ga = 69.9 | Gi = 72.5 |
| 6 | Rb = 85.4 | Sr = 87.6 | Yt = 89 | Zr = 90.4 |
| 7 | (Ag = 107.92) | Cd = 112.4 | In = 113.8 | Sn = 119 |
| 8 | Cs = 132.9 | Ba = 137.4 | La = 138.6 | Ce = 139.4 |
| 9 | (—) | — | — | — |
| 10 | — | — | Er = 166.3 | — |
| 11 | (Au = 197.2) | Hg = 200 | Tl = 204.1 | Pb = 206.9 |
| 12 | — | — | — | Th = 232.6 |

TABLE XII. — (*Continued*).

| SERIES. | GROUP V. $RH_3$, $R_2O_5$. | GROUP VI. $RH_2$, $RO_3$. | GROUP VII. $RH$, $R_2O_7$. | GROUP VIII. $RO_4$. | |
|---|---|---|---|---|---|
| 1 | —— | —— | —— | | |
| 2 | N = 14.04 | O = 16 | F = 19 | | |
| 3 | P = 31 | S = 32.07 | Cl = 35.45 | | |
| 4 | V = 51.4 | Cr = 52.1 | Mn = 55 | Fe = 56 | Ni = 58.07 |
| | | | | Co = 59 | Cu = 63.6 |
| 5 | As = 75 | Se = 79.2 | Br = 79.95 | | |
| 6 | Cb = 93.7 | Mo = 96 | —— = 100. | Ru=101.7 | Rh=103 |
| | | | | Pd=106.4 | Ag=107.92 |
| 7 | Sb = 120.4 | Te = 127.5 | I = 126.85 | | |
| 8 | Pr = 140.5 | Nd = 143.6. | | — — | — — |
| 9 | —— | —— | —— | | |
| 10 | Ta = 182.8 | W = 184.8 | —— | Os = 191 | Ir = 193.1 |
| | | | | Pt = 194.9 | Au = 197.2 |
| 11 | Bi = 208.1 | —— | —— | | |
| 12 | —— | U = 239.6 | —— | — — | — — |

NOTE. — This classification does not include argon, helium, other recently announced constituents of the atmosphere, and the rare earth metals.

## XIII. — EMERGENCY SET.

A box or cabinet provided with the following articles should be kept in a convenient place: —

Absorbent cotton, ½ lb.
Linen rolls, 1 in. and 2 in.
Court-plaster, thin and thick, cut in strips.
Emulsion of lime water and oil, 500 cc.
Sodium bicarbonate.
Vaseline, 1 lb. box.
Mortar or shallow dish.
Smelling salts.
Ammonium hydroxide.
Sand.
Blanket.

Cuts and burns should receive prompt attention. Fires may be extinguished by wet sand, if caused by burning liquids, or by a damp towel or blanket, if due to burning clothing. Acids and alkalies should be removed from the flesh with warm water, and the injured spot immediately covered with a paste made by mixing sodium bicarbonate and the emulsion of lime water and oil; the paste may be held in place by cotton and a linen bandage.

## VI. SUGGESTIONS FOR A SHORT COURSE.

The experiments incorporated in this book constitute a liberal course in chemistry. In view of the fact, however, that conditions, often temporary or unavoidable, may prevent the pupil from performing all the experiments, this course may be shortened without detriment to its integrity, if the following selection and apportionment of experiments be made: —

### A. PUPIL'S EXPERIMENTS.

1,
2,
3,
4,
8,
9,
10,
11 (*a*)
(or 11 (*b*)),
14,
18,
24,
25,
27,
28,
29,
32,
33,
34,
35,
36,
37,
46,
47,
48,
49,
50,
51,
52,
56,
58,
59,
64,
69,
70,
71,
72,
73,
76,
77,
81,
82,
83,
84,
86,
88,
90 (*a*)
and (*b*),
102,
107,
112,
114,
116,
117,
129,
133,
138,
139,
150,
151,
152,
160,
161 (*a*),
162,
167,
168,
176,
177 (*b*),
(*c*),
(*d*),
(*e*),
188,
189,
190,
191,
192.

### B. TEACHER'S EXPERIMENTS.

12,
13,
17,
19,
20,
21,
23,
30,
38 (or 39),
42,
57,
60,
61,
62,
65,
75,
98,
99,
108 (or 109),
124,
125,
128,
130,
131,
134,
136,
179.

The course provided by the above experiments is short, representative, and continuous. Suggestions regarding these and other experiments may be found in the TEACHER'S SUPPLEMENT, which will be sent to teachers on application to the publishers. The author takes advantage of a second opportunity to invite correspondence concerning the adaptation of the book to meet special demands.

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