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SCHOOL

CHEMISTRY.

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PREFACE.

This small book is the outcome of five years' experience in teaching teachers in the Summer School of the University of North Carolina. The material, which is not new to chemists, has been adapted to the needs of the schools in the Southern States, with which the author is most familiar. Science now plays an important part in the child's education. A large portion of the ninety per cent. of high school graduates, who do not secure a college or university training, never learn a chemical principle or fact except by accident. The course presented is not intended to replace college work, but seeks that class of people mentioned, who not only wish, but need, to know more of Nature's economy. By reference to the cost of apparatus (p. 152)—a serious stumbling block hitherto to the introduction of the subject into the curricula of high schools, when any one of the excellent elementary treatises were used—it will be seen that the laboratory equipment necessary is reduced to the minimum of cost. By a perusal of the experiments given, it will be seen that much store is laid by home-made apparatus.

While a preliminary course in Elementary Physics is desirable, such is not deemed necessary to the successful use of this book. Some good text-book on Physics, as Gage's, should always be on the teacher's table for reference. Richter's "Inorganic Chemistry" and Venable and Howe's "Inorganic Chemistry according to the Periodic Law," may be profitably used for reference, if the pupil desires fuller information on any point. Many words are purposely introduced into the text to broaden the vocabulary of the pupil. Hence a good dictionary should be close at hand.

For heating purposes, alcohol lamps may be used where gas is not handy.

The course laid out is planned to cover forty weeks, when one hour per week is devoted to it. Thirty-seven lessons are given; the next two periods should be used for blackboard review, using the Table of Contents as a guide; the fortieth period is to be used for examination. Experience has taught that the hour may be most profitably spent by the following apportionment of time:—Fifteen minutes to questioning on last lesson and comments by teacher on experiments already performed, attention being called to incorrect conclusions drawn; fifteen minutes to the reading aloud by the pupils of the lesson of the day; thirty minutes to the performance of experiments.

Absolute cleanliness of the laboratory tables and neat note books should be insisted upon. Bunsen said one should keep his laboratory so neat that he might work there in evening clothes without the least fear of soiling them. A good plan is to spread a sheet of white paper under all material with which one is working. The observations made from the experiments should be accurately described by the pupil in note books, clearness of diction being emphasized. The results of the experiments in Lesson I are given to be used by pupils as a type.

A diligent and interested teacher, with the aid of succeeding classes, can make a valuable collection of many of the substances studied. Where possible, the teacher could opportunely omit the experiments and take the class on an excursion through concerns dependent upon chemical principles—gas and dye houses, fertilizer works, for examples. (See Lesson XXVII.)

In Lesson XXXV, it is suggested that the teacher read to the class from Prof. King's "The Soil" (Macmillan & Co.), pp. 76–106, instead of laboratory work. Read slowly, and have pupils write what they remember.

To Dr. F. P. Venable, Professor of Chemistry, and Professor M. C. S. Noble, Professor of Pedagogy in the University of North Carolina, and Dr. W. L. Dudley, Professor of Chemistry in Vanderbilt University, thanks are due for reading the manuscript and making many suggestions of value. Mr. E. W. Myers, Engineer to the N. C. Geological Survey, kindly made the drawings from which the illustrations were made.

University of North Carolina. C. B.

TABLE OF CONTENTS.

TO BE USED FOR BLACKBOARD REVIEW.

SCHOOL CHEMISTRY.

LESSON I.

WATER.

1. Importance.—Did you ever notice that, when a man starts to build a house, one of the first things he does is to find the nearest source of water supply? Do you know that most of the prosperous towns or cities are located upon the banks of some stream, or lake, or arm of the sea? Did you ever observe that many of the country houses are near a spring, creek, or lake? Ask one of the aldermen of your place how much money the town spends for water. How long do you think you could go without drinking any water?

2. Physical States.—"Oh! water is such a simple thing," you exclaim. If you heat water sufficiently, you know it boils; it turns into steam; it seems to disappear; it evaporates. If you hold a cold plate over the top of a boiling kettle, you notice that the hot steam condenses and you get some of the water back again. How often you have seen the icicles hanging from the roof of the house! If you place a lot of them in a tin cup and put it upon the stove, you see them melt down to water; and leaving the cup there long enough, the water boils and passes off as vapor. We thus see the solid changed into a liquid, and then into a gas. These are the three physical states of matter—SOLID, LIQUID and GASEOUS. Almost all substances can exist in these three interchangeable states.

3. Matter.—But what is MATTER? Anything that occupies space and has weight is matter. Steam or any other gas has weight; so has a cannon ball or a potato; and as each "takes up room," it is a form of matter.

4. Divisibility of Matter. Molecules.—If with a sharp knife we cut a potato in half; then halve one of these two parts; in turn halve one of the quarters; then split one of the eighths, eventually the piece will be so small that you cannot hold it in order to cut it further. Shall we therefore assume that it is impossible to further subdivide this small particle because we have not the proper instruments to work with? No; but let us imagine ourselves with the very finest tools. We could continue the subdivision until the particle became very, very small. The ancient Greek philosophers were divided into two great schools of thought. One school maintained that the division could continue until nothing remained; *i. e.*, that matter is infinitely divisible. The other claimed that a point would be reached beyond which no subdivision would be possible; because, following the former reasoning, matter would be composed of an infinite amount of nothing. The generally accepted theory to-day is that matter is composed of very minute particles, which by themselves are not appreciable to the senses, yet maintain the properties or characteristics of the substance. These small particles are called MOLECULES. At a distance we see only a pile of bricks; we do not see the individual bricks. A molecule corresponds to a single brick in the pile. It is the smallest particle of matter which retains its properties, a part of the whole, and, if broken up, it ceases to be a brick, or molecule.

5. Conservation of Matter.—When you heated the water you did not destroy it. You allow it to freeze by standing out in the cold over night; you obtain the liquid again when the ice thaws. We might say, therefore, that it must be a substance that we cannot break up by heat or cold; *i. e.*, elementary, although we can get it into very fine particles. Snowflakes are only tiny frozen particles of water. They are so small, too, that they can sift in, down or up, through very small cracks. We can neither taste nor feel the tiniest snowflake, yet one of them changes easily into liquid water. These little crystals SEEM to be molecules of water. They are not, however. Molecules are even smaller, and the beautiful snow crystals are aggregations of numbers of molecules.

6. Omnipresence of Water.—Water is present in almost all substances we know. It is in the air, as evidenced by the rains and dew. It is in many rocks, although we do not see it, and in the soil. Plants and animals are largely composed of water, with a small proportion of mineral matter. It is omnipresent, and a most difficult problem with a chemist is to get a substance dry—absolutely free from water.

EXPERIMENTS.

I. Allow some ice to melt in a tin cup; note the temperature of the ice with a thermometer, and also the water produced by the melting of the ice. Heat the cup until the water boils, and note the temperature of the boiling water and the steam. Hold a dry plate over the boiling water, and incline it so that the condensed steam may be caught in another cup. (Can you make the water get hotter by more vigorous and longer heating?)

II. In a small glass tube, closed at one end, heat a little soil that appears dry; in another, heat a crystal of laundry soda or alum; in a third tube, heat gently some green twigs.

III. Cleanse two empty tin cans. Make a hole in the bottom of each with a nail. Fill one of them with clean sand; the other with any kind of soil. Place them upon glasses and pour in a dipper of water, and let stand until the next lesson.

A typical page from a note book:

.... (Date.)

1. The temperature of ice is zero; that of boiling water 100°. I observe that the solid ice changes to liquid on applying heat, and that the temperature of this liquid remains near zero till all the ice is melted. If the heating be continued, the liquid changes to a gas. When the heated vapor comes into contact with the cold surface of the plate, I observe that water is obtained. I would conclude that the three physical states depend upon the addition or subtraction of heat; in the former case, the changes are from the solid to the liquid, and then to the gaseous state; in the latter, the converse is true.

2. In all three of these experiments, I proved the presence of water in substances that appeared dry. As the materials used are very diverse in character, water must be widespread in its occurrence.

3. I observe that water passed through the two cans packed respectively with sand and dirt. That which passed through the sand was crystal in appearance; that which came through the soil acquired color and odor different from the original water. I conclude that if water passes through this small amount of soil, much of our rains must penetrate the earth; and, when the water makes its appearance again in springs or wells, it may have acquired an odor and color. This must depend upon the kind of soil through which it has passed.

.... (Name).

LESSON II.

WATER (Continued).

7. Molecular Vibrations.—From an experiment in our last lesson we learned that ice melts at zero, and water boils at 100° C. There is a natural law by which, when a substance is heated it expands. This is explained by supposing the molecules to be in continual vibratory motion. Heat augments this vibratory motion of the molecules, and increases the range of these movements; expansion is the result.

8. Cohesion *versus* **Vibratory Motion.**—Solids hold their shape through the influence of a powerful force called COHESION, which binds the molecules together. The vibratory motion of the molecules tends to overcome this cohesion. In solids the latter is stronger, but in liquids, that binding force is largely overcome, and the molecules have greater freedom of motion. In water, for instance, the molecules move so easily over one another that the shape of the liquid depends upon the confining walls of the containing vessel. If the vibration of the molecules be even greater, the entire cohesive force may be eventually overcome. Such is the case in gases, where instead of attraction, there seems to be an actual repellent force acting between the molecules.

9. Melting and Boiling Points.—These molecular vibrations in a solid are increased by heat until a point is reached when the substance fuses or assumes the liquid state. This is called the MELTING POINT. On further addition of heat to the liquid, the movements of the molecules become more and more active, resulting in greater expansion. A time comes in the heating when the molecules rush off into space with some violence; the liquid boils. The temperature at which this takes place is called the BOILING POINT. Many molecules of a liquid in their restless vibration pass off below the boilng point, presenting to us the phenomenon of EVAPORATION. For this reason, bottles of perfume have to be kept stoppered when not in use.

10. Solution.—If we place a substance in a liquid and the substance disappears—that is, unites with the liquid to form a homogeneous fluid—it is said to dissolve, and a solution results. We all know that water will dissolve sugar and salt. It dissolves many substances, but there are many substances it does not dissolve. A gas, when dissolved, is commonly spoken of as being "absorbed." When one liquid dissolves another, they are said to be "mixed."

11. Solvent Action of Water.—There is much material existing in the soil that is soluble in water. If you pour water upon a pile of sand, you see the water disappear; it sinks out of sight. Sand, one of the chief ingredients of the soil, gives it porosity, permitting it to drink up large quantities of water. This water, breaking up into small particles, easily passes through the little spaces between the soil grains, dissolving something here, something there, as it goes. No water which comes from the earth, therefore, is perfectly pure, however bright it may appear or pleasant it may be to the taste. The presence of many impurities is frequently easily detected in the waters of many rivers and springs by an unusual color or the possession of some peculiar odor or taste. When mineral constituents are present in rather more than usual amounts, it is said to be a MINERAL WATER, as chalybeate (iron), sulphur, or lithia. If calcium (dissolved lime) is present in notable amounts, the water is said to be HARD.

12. Distillation.—Gases are more soluble in cold water than hot, so we may get rid of them by boiling. The mineral substances that may be dissolved do not volatilize on heating as readily as does water, so we may purify water by changing it into steam and then condensing it (by cooling) to the liquid form again. The first portion of the condensed water is liable to contain the gases, and should be rejected. Life does not continue at the temperature of boiling water; hence by this means all living organisms, called bacteria, the source of many diseases, are destroyed. Water may be best purified, therefore, by distillation.

EXPERIMENTS.

I. (a) Set up apparatus, as shown in Fig. 1, and distil the water collected from the last experiment. (b) Dissolve some salt in water; taste, distil; taste distillate (condensed steam). (c) Place a few drops of red ink in some water and distil. (d) Add a drop of ammonia to some water and distil, collecting in four separate portions; smell each one. Save some of the pure water obtained from a, b and c.

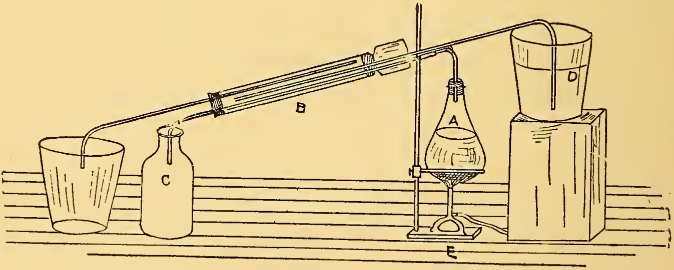

FIG. 1.

A—Distillation flask. B—Condenser. C—Receiver. D—Tank containing water, which keeps the condenser cool.

LESSON III.

ATOMIC THEORY. HYDROGEN.

13. Water a Compound.—Having obtained some pure water, and not being able to decompose it by heat or cold, let us see if we can produce in it any other than a physical change. If we drop a piece of the metal potassium upon some of our purified water, we observe light and heat. A gas, which we shall learn later does not come from the metal, escapes and takes fire. Some great change has evidently taken place in the water. We have decomposed it. Water is a compound, not an element, although it was thought to be so by the Greeks long years ago, and even by people everywhere till near the close of the last century.

14. Atoms.—The minutest amount of water that would still be water would be a molecule. But water may be decomposed, as we have just seen, into something which is no longer water; consequently we know that the molecule can be further subdivided. The very small particles of which the molecule is composed are called ATOMS, and, in the case of water, they no longer resemble that liquid at all. An atom is the smallest possible particle of substance we know of; that is, it matters not how it may be treated, cooled, heated, or what not; it may be caused to combine with some other substance, but no one can now knowingly change or destroy it. It may be done some day.

The old alchemists, living centuries ago, believed that one element could be changed into another, and tried to make gold by heating mercury and sulphur together.

15. Elements and Compounds.—If the atoms composing a molecule are alike, the molecule is ELEMENTARY. If the atoms are unlike, we have a COMPOUND MOLECULE. A water molecule is made up of atoms of two different gaseous elements, as we shall learn, hence it is a compound, and not an element, as was once thought.

16. Atomic Theory.—To review: the smallest indivisible particle is an atom; atoms combine in molecules; if alike, an elementary; if unlike, a compound molecule; a number of molecules constitute MASS—that is, any appreciable amount of matter. As matter has weight, those particles of which it is composed must also have weight; hence the atoms have weight. The atoms of the different elements have their own constant fixed weights, called ATOMIC WEIGHTS, which, as far as we now know, is unalterable. Such in substance is the ATOMIC THEORY as propounded by John Dalton in 1803.

17. Hydrogen.—The gas given off when potassium or sodium acts upon water is the lightest substance known to chemists; the weight of its atom is therefore taken as the standard of atomic weights and placed at unity. This gas is called HYDROGEN. Set it afire and it burns with a pale blue, very hot flame. A mixture of hydrogen with air is very explosive, hence the greatest care should be exercised when experimenting with this gas.

18. Occurrence.—This light substance is found free on the earth only in the gases of some volcanoes and certain natural gas wells. It is said to occur abundantly, however, in the atmosphere of the sun and of many fixed stars. On earth, hydrogen occurs combined in all animal and vegetable matter, and constitutes one-ninth of water (by weight).

19. Preparation.—Hydrogen can best be prepared on the small scale by treating scraps of zinc with dilute hydrochloric or sulphuric acid. As the metallic zinc is dissolved, the gas is evolved. It is collected over water, but, as it is so light, it may also be caught by holding an inverted bottle over the exit tube (upward displacement of the air).

20. Chemical Solution.—In our last lesson we learned that when sugar or salt was dissolved in water a SIMPLE SOLUTION resulted. The sugar or salt may be easily recovered by boiling away the water. No real change has taken place in the sugar or salt. Should we evaporate the water from some of the solution obtained by dissolving the zinc in the acid, we should not recover the zinc as metal, but instead a white crystalline substance possessing very different properties. Quite a complicated process is necessary, in fact, to change this white substance back into the gray metal. This is a solution just as much as the other, only the action is not mainly physical but chemical, brought about by one of those powerful forces of nature called CHEMICAL AFFINITY. Chemistry is that branch of science which treats not only of the different kinds of matter, but as well of the changes brought about by means of the force mentioned.

EXPERIMENTS.

I. Test a little of the pure water obtained in last experiment with red litmus. Drop a piece of the metal potassium on the water. When the violent action has ceased, test again with the litmus paper.

II. Immerse a test tube mouth upwards in a pan of water*; as soon as all the air has escaped, invert the tube, which will now be filled with water. Grasp a piece of sodium half as large as a pea with a pair of dry tweezers, and quickly thrust it into the water and under the mouth of the test tube. When filled with the escaping gas, place the thumb over the mouth of the tube, and while it is still inverted bring close to the flame removing the thumb.

CAUTION.—Never handle sodium with WET hands or tools.

III. Fit up apparatus like that shown in Fig. 2. Add 10 cc.† dilute hydrochloric acid to several small pieces of zinc in the bottle. (a) Collect several tubes full of the gas over the pneumatic trough, throwing away the first three. Why? Collect one tube full by upward displacement. Test all by holding to a flame. (b) Attach the exit tube to a clay pipe by means of a rubber tube, and fill some soap bubbles with the gas. Prepare soap suds by using

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*A vessel filled with water and used in working with gases as above, is called a PNEUMATIC TROUGH, an invention of Priestly.

†Cc.=cubic centimeter; cm.=centimeter.

castile soap and adding a little glycerine. (c) Substitute for the clay pipe a straight hard glass tube about 6 cm. long, drawn out to a small bore. After a minute, set fire to the gas coming out. Hold a DRY tumbler down close over the flame.

IV. Take some of the solution formed by the action of the acid upon the zinc in the above, and boil away most of the liquid. Cool. Do you obtain metallic zinc?

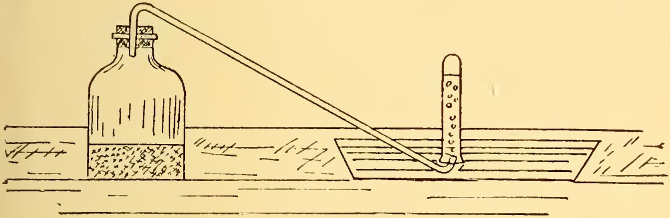

FIG. 2.

LESSON IV.

OXYGEN. OZONE.

21. Oxygen.—Having found that one of the constituents of water is a gas, let us see of what else it is composed. In the latter part of the third experiment in our last lesson we observed that water resulted when hydrogen was burned in the air. As we shall soon learn, the air is mainly a mixture of two gases. In one of these, hydrogen does not burn at all; but in the other it does burn, giving out great heat and producing water as the result of its burning. This latter gas is OXYGEN, which in the free state constitutes one-fifth by volume of our entire atmosphere. It is very abundant, occurring not only free, but combined with almost all known elements. About two-thirds of the earth is composed of this gaseous element.

22. Discovery.—It is generally accepted that Priestley, an English preacher, discovered oxygen in 1774, by heating the red oxide of mercury with a sun glass. Lavoisier, a French *savant*, named it oxygen meaning "acid producer."

23. Preparation.—We may prepare the gas as did Priestley in his original experiment, or best by heating a mixture of potassium chlorate and black oxide of manganese (manganese dioxide). The gas is collected over water.

24. Properties.—Oxygen is a colorless, odorless gas, and what is known as a supporter of combustion. A fire burns in the air on account of the oxygen that is present. If the flame of a lighted match be blown out and only the glowing stick be thrust into a tube of oxygen, it will burst out again with flame. The less oxygen there is present, the less brilliantly does the taper burn. The more oxygen there is present, the more vigorous is the fire; hence oxygen is a supporter of combustion. What is true of the flame is also true of the human body. We breathe solely that we may bring oxygen into our lungs to carry on there and throughout the body a similar process, viz: combustion, but one of very much less violence.

25. Law of Definite Proportions.—By passing an electric current through acidified water we decompose it. At the negative pole of the battery we will observe twice as much gas liberated as at the positive pole. If we test these gases with lighted tapers, we find one to be hydrogen and the other oxygen, liberated in the proportion of two volumes to one, respectively. We have seen that by burning hydrogen in the air we obtain water; we can decompose water into two gases; hence we learn the composition of water by synthetic and analytic means. Further, hydrogen and oxygen are always present in water in the proportion of two volumes to one. If that ratio be altered in the least, that gas which is present in greater amount than necessary for the ratio of two to one (by volume) will be left over when the hydrogen and oxygen are caused to unite to form water. In short, in no other proportion will these two gases unite to produce water. This is true of all chemical compounds, namely, that *the same elements always enter into the same compounds in the same proportion.* If the proportion varies, the compound produced is different.

26. Allotropism.—Ordinarily an elementary molecule is composed of two atoms. Such is the case with hydrogen and oxygen. It is possible to imagine three or even more atoms in the molecule. We have an example of this in oxygen. The ordinary molecule contains only two atoms, while OZONE contains three; that is, a molecule of this latter substance weighs half as much again as one of oxygen. When an element exists in two or more different forms, it is termed allotropism. In one hand you may hold two eggs, comparable to the oxygen molecule; in the other, three eggs, which may be compared to the ozone molecule—an egg representing an oxygen atom.

27. Ozone.—The peculiar odor noticed around some large electric dynamos when they are running is due to ozone. The odor of the oil must not be mistaken for that of ozone, which is produced by the discharge of electricity in oxygen or the air. It is also produced by the slow oxidation of phosphorus in moist air. Aside from the peculiar odor by which it may be detected, we can test for it with paper which has been previously moistened with a solution of potassium iodide and starch. The ozone decomposes the potassium iodide, liberating free iodine, which combines with the starch to produce a blue colored compound.

28. Medicinal Value.—Ozone is very much less stable, but far more active chemically, than oxygen. It is decomposed by decaying organic substances; hence, in and about large cities, where much filth accumulates, there is very little ozone in the air. Lightning produces the ozone in the air. It is said that the waves beating upon the rocks on the seashore produce ozone. This allotropic form of oxygen kills the germs which cause many diseases, as pulmonary troubles. Our doctors send us to the mountains or the seashore, where the ozone renders the air crisp and exhilarating. Air containing too much ozone is very irritating, and in a concentrated form ozone is poisonous.

EXPERIMENTS.

I. Set up apparatus like that shown in Fig. 3. Place half a teaspoonful of red oxide of mercury in the tube A, which should be of hard glass. Heat. Collect the gas evolved over the pneumatic trough. Test about the third tube of gas, open end up, by plunging in a splinter, which, having been lighted and blown out, has a spark on the end.

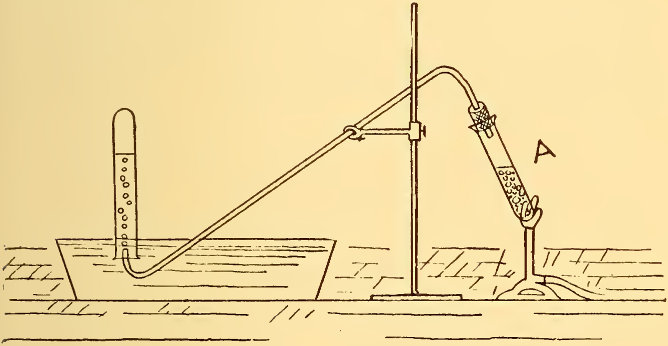

FIG. 3.

II. Place a teaspoonful of a mixture of two parts of manganese dioxide, and three parts of potassium chlorate in apparatus like the above, and heat. After the air has been expelled, collect four tubes and one bottle full of the gas. (a) Test one with a lighted taper. (b) Hollow out the end of a piece of chalk 3 cm. long. Bind this, hollow end up, to an iron wire. Place a piece of phosphorus the size of a pea in the hollow; set fire to it and plunge into the second tube.* (c) Bind a piece of charcoal to a wire with smaller wire; heat one corner of the charcoal until it begins to glow, and then plunge it into the third tube. (d) In the fourth

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*CAUTION.—Never handle phosphorus with the bare hands, except under water. Use tweezers or a pin. Always cut it under water.

tube place a piece of moist potassium iodide starch paper. Prepare the starch paper by boiling several grains of starch in water in a test tube until dissolved; add a crystal of potassium iodide and use the liquid to moisten strips of filter paper. These strips may be dried in the air and kept in a corked bottle. (e) Bind a thin watch spring on an iron wire. Warm the free end of the spring, and touch it to a small piece of sulphur. Set fire to the sulphur, which should be attached to the spring, and plunge into the bottle of oxygen. Repeat experiment if no violent action takes place.

III. Place a piece of phosphorus 2 cm. long in a wide-mouthed bottle; add enough water to almost submerge the phosphorus; cover the bottle with a piece of glass, and allow to stand twenty minutes. Note the odor of the air in the bottle, and drop in a piece of wet potassium iodide starch paper.

LESSON V.

THE AIR.

29. The Atmosphere.—Lying prostrate upon the ground on the edge of some lake or stream, have you not often watched the fish gliding hither and thither in the water? Did you ever realize that we, too, live in a great ocean? We do, but our ocean, as we come in contact with it, is not liquid, but gaseous, and completely envelopes the earth. The whirlwinds and tornadoes are the whirlpools and currents of the ocean in which we live.

30. Thickness and Necessity of Air.—You do not see some gases, but imagine yourself standing afar off somewhere looking at the world with the power of seeing air. The atmosphere would appear to you as liquid water appears to you now. You would see the birds floating around in it, but not reaching the surface, for the highest flying bird, the condor, never soars more than five miles from the earth, and this aerial ocean of ours is at least one hundred miles deep. We people, who live and move around on the earth, would be to you "deep air" creatures, like some of those strange, weird looking animals which live along the sea bottoms. Suppose you fished for us and caught one of us so foolish as to bite your bait and jerked him from his ocean home. You would see him gasp and die just as the many little sunperch you have so frequently drawn with such delight from the pool. The air is not necessarily always a gas, for by complicated apparatus it may be liquefied and then frozen to a solid, as we know is possible for water.

For a long time the ancients regarded air as an element, as they did water. We found water, however, to be a compound and not an element; possibly we may make the same discovery with regard to air.

31. Air a Mixture.—Phosphorus, as we have learned, combines with oxygen, producing flame. If we set fire to a piece of phosphorus enclosed in a definite volume of air confined over water, at first we note a very violent action taking place. Dense white fumes are produced, and after the phosphorus ceases to burn, the white fumes, phosphorus oxides, are gradually absorbed by the water which rises up into the jar. When all the fumes have disappeared, and the jar is quite cold again, we will notice, if we lower the jar into the water until the water stands at the same level on both sides (inside and out), that one-fifth of the air has been burned up by the phosphorus. The portion which has been used up is oxygen, which makes up one-fifth by volume of the atmosphere. Should we test the remaining gas in the bottle with a lighted taper, we would see the fire extinguished, thereby proving that its nature is very different from that of oxygen. It is a very inert gas—called NITROGEN. Air is a mixture of these two gases. It is not an element, nor is it a compound, as no chemical action results when nitrogen and oxygen are mixed in the proportion of four to one; and the mixture has all the properties of air, as may be proven by the experiment just recited.

32. Oxygen Necessary for Life.—The oxygen is that portion of the air we use when we breathe. Should we shut up a mouse in a closed jar, as we did the phosphorus, he would soon use up all the oxygen and die from suffocation. From this all of us can see how foolish it is to live in rooms that are closely shut up or to sleep with the head under the bedclothes.

33. Rôle of Nitrogen in Air.—Doubtless some of you will say: "If oxygen is so necessary for life, couldn't we live much better in pure oxygen?" Let us reason it out. Oxygen combines with phosphorus or carbon producing heat. Our bodies are kept warm by the heat resulting from the union of oxygen with carbon compounds. We breathe out the inhaled oxygen as carbon dioxide. If the air were pure oxygen our bodies would rise to a fever heat and even higher, and all the fires and lights in the world would burn furiously. A fire once started would spread all the world over, and only an Almighty Power could check it. Therefore, this lazy, inert gas, nitrogen, has great value, as it serves to dilute the oxygen, thus retarding its rate of combining with other substances.

34. Other Constituents.—We know that the air contains many impurities dependent entirely upon the locality. Some of these impurities evidence themselves by disagreeable odors noticed about glue, fertilizer, or perfume factories; emanations from new-made ground, stables, and so forth, Suspended dust is almost a constant impurity in the air. Its presence is shown by allowing a beam of sunlight to enter a dark room. But there are two other gases always present in the air, though in small amounts. Inhaled oxygen is converted into carbon dioxide and exhaled as such. In ten thousand parts of air there are about four parts of this carbon dioxide—a small quantity, yet of great value, as we shall see. You scarcely imagine "spirits of hartshorn" existing in the air. It is always present, however, but in exceedingly small amounts, a few parts in the million. These two gases in the air are essential to plant and animal life, but no more so than the vapor of water, which is always in the air.

35. Water Vapour in Air.—Warm or hot air will absorb very much more water vapour than will cold air, hence when warm air, which has been saturated with water, is chilled, we have rain, or snow, or hail. This occurs on a small scale when "dew falls" at night. It does not fall at all. The air during the daytime becomes warm by the action of the sun's rays, and then absorbs more water vapour. The grass and objects of similar nature cool more rapidly at night than does the air, so the moisture of the air is merely condensed on these colder surfaces. If we bring a pitcher of very cool water into a warm room, we see the moisture of the air condense on the sides of the pitcher. Air that is devoid of moisture is very trying, because it dries up the body, absorbing the water from the animal fluids to supply its own thirst. Fortunately, the body can in time adjust itself to new conditions, and the lack of water can be supplied by drink.

EXPERIMENTS.

I. Fill a small, tall bottle with water and place it in the pneumatic trough so that it will project about a centimeter above the surface of the water. On the bottle place a shallow piece of porcelain, holding a piece of phosphorus about the size of a pea. Ignite the phosphorus with a hot wire and quickly cover with a wide mouth bottle; empty quinine bottle will do. Hold the bottle firmly until the phosphorus has ceased burning. Allow the bottle to stand in the water until perfectly cold, and the dense white fumes of phosphoric oxide have been absorbed. Raise or lower the bottle until the water which has risen on the inside is on a level with that on the outside. Why do this? Measure with a rule the height of the bottle above the surface of the water and the entire length of the bottle. State your conclusions.

II. Expose some clear lime water in an open glass vessel to the air for half an hour. What do you see? Lime water, which is prepared by slacking quick lime and filtering, becomes milky, with carbon dioxide, owing to the formation of a white solid.

III. Blow a deep breath into a solution of clear lime water in a test tube by means of a small tube.

IV. Test some of the gas left in the bottle in the first experiment. Does it act like oxygen or hydrogen?

LESSON VI.

OXIDATION.

36. Oxidation. Combustion.—We have seen that phosphorus and charcoal burn more brilliantly in oxygen than in air. We also saw that if no oxygen be present the fire is extinguished, whereas even a steel watch spring will burn in oxygen. The burning of a substance is called COMBUSTION. The substance burned is said to be COMBUSTIBLE, and the substance, usually a gas, which promotes the burning or makes it possible, is called a SUPPORTER of combustion. When the supporter of combustion is oxygen, as in the cases mentioned, the process is termed OXIDATION.

37. Rust. Flame.—We all know that bright metallic surfaces will rust on exposure to moist air. If the air be perfectly dry, no rusting takes place; the presence of water therefore seems necessary. There are other substances which will cause the rusting to be more rapid, but they frequently serve merely to hold a thin film of moisture on the metal. This rust is nothing more than a compound of the metal with oxygen. Its color is red-brown with iron, black with copper, and is the result of oxidation. It is exactly similar to the burning of the steel spring in oxygen, only in the latter case it is more rapid. When oxidation is very rapid, we call it combustion. FLAME is usually one of the results of combustion. The luminosity of the flame is due to the glowing of particles of the burning substance so intensely heated. We may have both light and heat without flame.

38. Incandescence.—Suppose we have three wires, one of lead, one of platinum, and one of magnesium, and heat each of them in a flame. When we set fire to a substance, we say we "ignite it." Sometimes it will melt or fuse, as in the case of the lead wire. Again it will not, but becomes red, then white hot, without melting or oxidizing. The platinum wire thus, we say, becomes INCANDESCENT, it glows. No change, except a physical one, takes place in either of the cases mentioned, except the small amount of oxidation which the lead may undergo, and that is insignificant. Now, when we try the magnesium wire, we see that it burns with a brilliant light. The white powdery substance left is entirely different from the gray metallic wire we used. A CHEMICAL CHANGE has taken place. The white powder is the oxide of magnesium.

39. Oxidation Aided by Heat.—If we expose a piece of magnesium wire to moist air or oxygen for a long time, we should see that the white powder would be produced finally, but no flame would result. To make the oxidation rapid, then, we see that heat must be added to facilitate the union of the substance (metal in this case) with the oxygen. A mixture of one volume of oxygen and two volumes of hydrogen—that is, in the proportion to produce water—is harmless until a lighted match, or other source of heat, is brought in contact with it. The oxidation is then so rapid that a violent explosion results. Such experiments are too dangerous for any one not acquainted with all the conditions to perform, so we shall have to omit an ocular demonstration of this important principle.

40. Heat, a Result of Oxidation.—It is a law of chemical force that, when one element unites with another, heat is produced. When the iron or magnesium wire slowly rusts—becomes oxidized in the air—the same amount of heat is produced as when the rapid combustion takes place, only it is spread out through such a long time that the heat is dissipated or seemingly lost. We have substances, however, which of their own accord become oxidized in the air—slowly at first, then rapidly, and finally burst into flame. Such substances are said to be SPONTANEOUSLY COMBUSTIBLE; that is, by gradual oxidation the heat generated accumulates to such an extent that the temperature of the body is raised to the ignition point, at which oxygen rapidly combines with it. Many fires, especially in mills and factories, result from just such a cause. Old rags, usually greasy, and refuse matter are piled away. Oxygen combines with some of the oily substances, producing heat, which gradually accumulates in sufficient quantity to raise some of the material to the temperature of ignition, and fire is the result.

Therefore, to have fire, we require a combustible substance, a supporter of combustion, and sufficient heat to raise the temperature of the material to the point of ignition.

EXPERIMENTS.

I. Place a small pinch of red lead (oxide of lead), mixed with powdered charcoal, upon a piece of charcoal, and heat it with a blowpipe, directing the flame obliquely upon it (Fig. 4). Oxygen combines very readily with the charcoal, leaving the metal behind. Copper oxide may be substituted for the red lead, but longer heating is necessary.

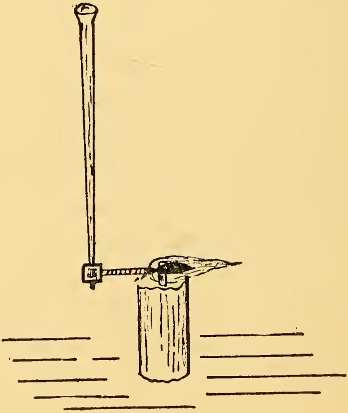


Fig. 4.

II. Carry out the experiment with the three wires, as given in the lesson to day.

III. Expose a piece of dried phosphorus upon a dry plate in a jar of oxygen. Dry the phosphorus by pressing in a dry towel or piece of paper. Do NOT RUB IT!!

LESSON VII.

NITROGEN. AMMONIA.

41. Nitrogen.—After removing the oxygen from the air by burning it out with phosphorus, we found the remaining gas to be a very lazy substance. It is quite different from either hydrogen or oxygen, being neither combustible nor a supporter of combustion. Free nitrogen is, in fact, a very inert, inactive substance, possessing neither taste nor odor. Freshly boiled water can absorb one and a half per cent. of the gas however.

42. Occurrence. — Free nitrogen gas constitutes four-fifths by volume, as we have seen, or three-fourths by weight, of the atmosphere. Combined with hydrogen, it is present in small quantities in the air. It is present in nitrates, commonly as "saltpetre," or "nitre," from which its name is derived. Nitrogen combines directly with very few elements. Indirectly nitrogen may be made to combine with many of the elements, producing bodies varied in character, and frequently, powerfully reactive. Nitrogen is also present in some of the most poisonous vegetable principles.

43. Use.—Nitrogen is not only a diluent for oxygen in the atmosphere, but it plays a most important part in life. Plants can utilize many of the nitrogen compounds directly. There are germs in the soil which have the power of rendering nitrogen almost directly available to the plants. The free gas is not assimilated.

44. Ammonia.—The first compound of nitrogen we shall study is the common substance, AMMONIA, a gas resulting from the union of three volumes of hydrogen, with one of nitrogen. In desert countries, destitute of trees, the natives in need of fires, use whatever fuel comes nearest to hand. The excrement of camels, the beasts of burden of some of these arid wastes, has long proved a profitable source of heat for cooking food. The Arabs, years ago, knew ammonia by the peculiar odor given off when this refuse, parched by the sun, was burned. This method of heating was practiced about the Temple of Jupiter Ammon, whence came the name, ammonia.

45. Preparation.—As nitrogen and hydrogen combine directly only with difficulty, the best method of preparing ammonia is by heating ammonium chloride (sal ammoniac) with quick lime.

46. Properties. — Ammonia is a colorless gas, lighter than the air, and possessing a strong, pungent odor. It is exceedingly soluble in water; one part of cold water dissolves eight hundred times its volume, hence the gas cannot be collected over water, but by upward displacement. The water solution is our ordinary ammonia water or spirits of hartshorn, so named because it can be made by distilling the horns of the hart. The gas and its water solution turn red litmus paper blue, which is a characteristic of many basic substances. All ammonium salts are volatilized on the application of heat.

47. Uses.—Ammonia has many uses in science and the arts. Scarcely a houskeeper can fail to tell one some of the many practical uses to which it is put, as cleaning silver, removing paint, washing windows, and so on.

EXPERIMENTS.

I. Place equal amounts, about 2 g. each, of sal ammoniac and quick lime in a test tube and heat gently. Note the odor of the gas evolved by fanning the fumes under the nose with the hand. Do not breathe too much of it. Moisten a piece of red litmus paper and hold it over the tube without allowing it to touch the glass.

II. Open a bottle of strong ammonia (ammonium hydroxide) near a bottle of strong hydrochloric acid. The dense fumes are ammonium chloride, the material heated with the lime in Experiment I.

III. Insert a cork with an exit tube, 10 c. m. long, in the tube used in Experiment I, and collect a test tube full of the gas by upward displacement. Place the thumb over the open end of the tube used for collecting the gas, and plunge that end of the tube into water. On removing the thumb, what do you see?

IV. Place a small piece of ammonium chloride upon a piece of mica or platinum. Heat it over the flame. Repeat the experiment, placing the sal ammoniac in a test tube and heating it.

LESSON VIII.

CHLORINE.

48. Colored Gas.—So far in our studies we have had only colorless gases to deal with. Many gases are colored, and the one we learn of to-day is greenish-yellow. This gas, called CHLORINE, from a Greek word descriptive of its color, is conveniently prepared by heating strong hydrochloric acid with manganese dioxide. Being soluble in water, and over twice as heavy as the air, it is collected by downward displacement.

49. Properties.—It has a very disagreeable and irritating odor. It is poisonous when inhaled, and one must be careful not to breathe it. When absorbed by water, that liquid acquires a yellow tint. When mixed with hydrogen and exposed to the sunlight, or ignited, a violent explosion results, with the formation of the important compound already mentioned, hydrochloric, sometimes called "muriatic," acid. Since chlorine kills disease germs, it is used as a disinfectant.

50. Chlorides.—Many metals, as copper or antimony, combine directly with chlorine, burning with a bright flame. The compound formed is a CHLORIDE. On account of this affinity for metals, chlorine is never found free in nature, but always in combinations called chlorides; for example, sodium chloride (common salt).

51. Bleaching.—Chlorine has such a strong affinity for hydrogen that it will decompose many bodies containing that element by combining with the hydrogen. Thus, with water, it forms hydrochloric acid and liberates free oxygen. Oxygen liberated in this manner possesses a peculiar power of destroying or changing many coloring matters. Oxygen, as ordinarily prepared, cannot do this. This extraordinary activity of oxygen is accounted for by assuming that it is in the NASCENT STATE; that is, the state of being born. Now, the molecule of oxygen contains two atoms, which saturate one another, as it were; part of the combining energy of the atoms is used up in the union. Therefore, it is obvious that oxygen atoms not so saturated would possess more energy than otherwise, and this state of unsaturation exists before any union has occurred, or at the time of being born, so to speak. Many moist colors, as in leaves, flowers and inks, are bleached by this activity of chlorine in liberating nascent oxygen. If the substance be dry, no water being present to be decomposed, no bleaching results. Printer's ink is not decolorized, because it is mainly composed of minute particles of carbon, which are not affected by nascent oxygen.

The most convenient form in which chlorine can be used for these purposes is in the compound "chloride of lime," or "bleaching powder," which is made by passing chlorine gas over lime.

EXPERIMENTS.

I. In apparatus shown in Fig. 5 place a teaspoonful of manganese dioxide, and cover the black powder with concentrated hydrochloric acid, shaking until thoroughly wet. Add as much more acid and warm gently. Collect the gas by downward displacement, resting the end of the exit tube upon the bottom of the bottle. When the gas rises until it rolls over the edge of the mouth of the receiving bottle, remove, and cover with a piece of paper or glass plate. In this manner collect five dry bottles full of the gas, noting the color and choking odor. If a hood or a draught chamber be at hand,

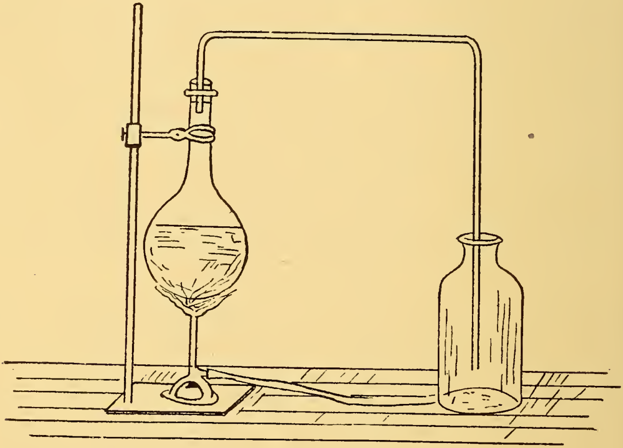


FIG. 5.

perform these experiments under it; if not, work either out of doors, or where there is good ventilation. (a) In one bottle introduce with tweezers a piece of thin copper foil or Dutch leaf. (b) In another, drop a pinch of powdered antimony which has been heated upon the end of a knife blade. (c) In another, place a flower, a moist piece of paper upon which there is writing with ink and a wet piece of calico. (d) In the fourth, place a dry piece of the same calico. What conclusion do you draw from these last two experiments? (e) In the fifth, insert a piece of filter paper which has been moistened with turpentine.

II. Place a teaspoonful of "bleaching powder" in a small beaker. Half fill the beaker with water. In a similar beaker, place an equal amount of dilute hydrochloric acid. Partially dip a strip of chocolate calico alternately in one and then the other several times. Finally wash the entire piece in clear water.

LESSON IX.

HYDROCHLORIC ACID.

52. Gaseous Compounds.—We have gaseous compounds as well as gaseous elements, as we learned in the case of ammonia. Ammonia is a compound of hydrogen and nitrogen, possessing basic properties. To-day we have a compound containing hydrogen and chlorine, which is the very opposite of ammonia in being acid and not basic. It is HYDROCHLORIC ACID.

53. Preparation.—We may prepare this acid by heating ordinary cooking salt with sulphuric acid. This may be done in apparatus similar to that used for the preparation of chlorine (Fig. 5, p. 40). Hydrochloric acid gas is given off and sodium hydrogen sulphate remains behind as a crystalline compound in the flask.

54. Properties.—Hydrochloric acid is a colorless gas, heavier than the air. It has a choking odor, and produces coughing when breathed. Water dissolves about five hundred times its volume at 15°C. The solution of the gas in water is what we usually see in drug stores and use in our experiments, and is sold as "muriatic acid." It is neither combustible nor a supporter of combustion.

55. Solvent Action.—Hydrochloric acid dissolves many metals liberating hydrogen gas. It also dissolves many oxides producing water and a chloride. If added to certain clear solutions, silver nitrate, for example, an insoluble solid is produced. For these reasons hydrochloric acid is not found free in nature. It does occur in small amount in the gastric juice of man and many animals.

56. General Properties of Acids.—As this is the first acid we have studied, we may learn something about the general class of bodies called ACIDS. All acids contain hydrogen combined with some other substance; for instance, hydrochloric acid contains hydrogen and chlorine. Other substances may take the place of chlorine in hydrochloric acid, as, for instance, bromine or sulphur, but a different acid is formed in each case. Although all acids contain hydrogen, all compounds which contain hydrogen are not necessarily acids. Most acids, when diluted, possess a sour taste, and put the teeth on edge. Many of them are corrosive poisons when strong—that is, concentrated. An acid is usually made dilute by adding water to the strong. Acids, as a rule, turn litmus, a vegetable coloring matter, red. Bases, as we shall learn, turn litmus blue (§ 76).

EXPERIMENTS.

I. In apparatus similar to that shown in Fig. 5, place a teaspoonful of ordinary cooking salt and add enough concentrated sulphuric acid to cover the sodium chloride. Heat gently; do not allow the froth to get up into the neck of the flask. (a) Note odor. (b) Expose a piece of moist blue litmus paper to the fumes. (c) Open a bottle of strong ammonium hydroxide and bring near the exit tube. (d) Collect a dry test tube full of the gas by downward displacement. Cover the tube with the thumb, and quickly invert in a pan of water and remove the thumb. The smoking of the acid in the air is due to its great attraction for moisture.

II. In a test tube containing a little dilute hydrochloric acid drop a small piece of zinc; in another, a pinch of magnesium oxide; in a third, a drop of silver nitrate solution. For comparative observations place tubes in rack (Fig. 6). Prepare dilute hydrochloric acid by adding one part of concentrated acid to two of water.

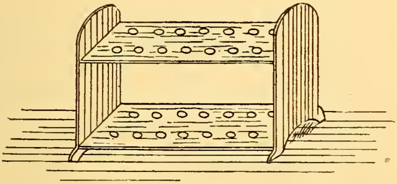


FIG. 6.

III. To ten drops of water add one drop of dilute hydrochloric acid. Taste the liquid. Test it with blue litmus paper.

If hydrochloric acid be accidentally dropped upon the clothing, add ammonium hydroxide to the spot and wash out.

LESSON X.

COMPOUNDS OF NITROGEN AND OXYGEN. NITRIC ACID.

57. Laughing Gas.—Nitrogen forms five compounds with oxygen. One called NITROGEN MONOXIDE, or nitrous oxide, is the "laughing gas" used by dentists to produce insensibility to pain in drawing teeth. Sir Humphrey Davy, an Englishman, was the first to prove the anæsthetic power of the gas. It may be prepared by carefully heating ammonium nitrate, and collecting the gas over warm water. The colorless gas is somewhat soluble in cold water and has a sweet taste. When sufficiently heated, it will support combustion.

58. Other Oxides of Nitrogen.—NITROGEN DIOXIDE, or nitric oxide, is produced when fifty per cent. nitric acid acts upon metallic copper. It can be collected over water and is colorless. On exposure to the air, however, it immediately combines with oxygen to form the TRI- and TETROXIDES of nitrogen, which are deep brown or red in color. The fifth, or PENTOXIDE of nitrogen, is a white solid. It dissolves in water to form NITRIC ACID. All these oxides, except the monoxide, are poisonous.

59. Nitric Acid was called AQUA FORTIS by the old alchemists, who knew the substance only in an impure form. Many still use the name. It combines with metals and metallic oxides to form nitrates. The acid is not formed in nature, but the nitrates are. Potassium nitrate is "saltpetre," used in making gunpowder, and sodium nitrate is "Chili saltpetre," used in fertilizers.

60. Preparation.—Nitric acid is prepared by distilling a mixture of sodium nitrate and concentrated sulphuric acid. It is a colorless liquid when pure, quickly turning yellow under the influence of light, being thus easily decomposed, although it is one of the strongest acids.

61 Properties.—Nitric acid is a very strong oxidizing agent, as we might suppose, when we learn that each molecule contains, besides one atom each of hydrogen and nitrogen, three atoms of oxygen. The strong acid acts upon organic bodies, such as dry sawdust, generating heat sufficient at times to set fire to the substance. It colors the skin yellow. Silk and wool are sometimes dyed yellow by the action of nitric acid.

EXPERIMENTS.

I. A round-bottomed flask of 100 cc. capacity, provided with a rubber stopper and exit tube, is necessary. Place the flask in an iron clamp, and pour in a tablespoonful of ammonium nitrate. Heat cautiously, and collect the gas, which is nitrogen monoxide, in bottles over warm water. After noting odor and taste, insert a lighted match in one bottle of the gas; in another suspend a small piece of burning phosphorus in a chalk cup.

II. Place a teaspoonful of copper clippings in a flask provided with stopper and exit tube (Fig. 3.) Pour in 5 cc. concentrated nitric acid diluted with an equal amount of water, and collect several bottles of the gas over water Use water bottle (Fig. 7) for diluting. It should be colorless when collected. Place a glass plate under the last bottle, and raise it from the water; invert and remove the plate. Explain.

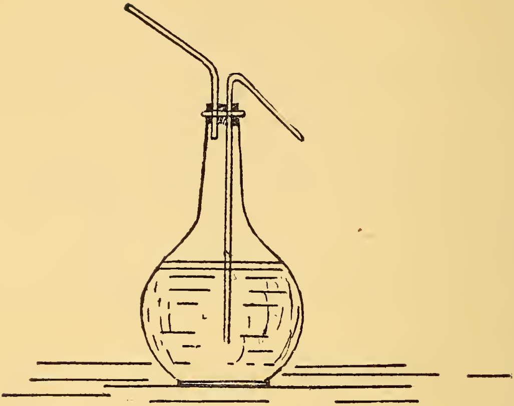


FIG. 7.

III. Put a handful of dry sawdust in a collar box, and pour a tablespoonful of strong nitric acid on it and allow it to stand. Perform this experiment out of doors.

IV. Dip a piece of clean wool for about one minute into concentrated nitric acid. Remove and wash.

LESSON XI.

CARBON.

62. Allotropism.—So far we have had only one case of an element existing, as an element, in more than one form, and that was the gaseous element oxygen. Now we have an element—CARBON—which exists in at least three distinct modifications. The diamond and graphite varieties are crystalline, while the third form, comprising many subvarieties, is AMORPHOUS—that is, without crystalline structure.

63. Artificial Diamonds.—Diamonds are found in South Africa, Brazil and India, usually in a blue cement rock, which fills the craters of extinct volcanoes. They also occur in alluvial deposits from the oldest rock formations. Very small diamonds have been artificially produced recently by an eminent French chemist, Moissan, by dissolving carbon in molten iron, and subjecting this to pressure as it cooled. Diamonds fuse at about 3000° C. We know of no substance which cannot be melted at that extremely high temperature.

64. Jewels.—The diamond, as it is found in nature, resembles very much an ordinary quartz pebble, and is not at all like the beautiful gem we see in the jewelry shops. It has been cut and polished to present a brilliant and attractive appearance. It is one of the hardest known substances.

65. Plumbago.—Graphite, often called plumbago or black lead, is found in Ceylon, Siberia and the United States. It is black and lustrous, and has a greasy feel. It makes a black mark upon paper, and is used as a stove polish, in making crucibles, and as a lubricant for machinery. Bicycle chains are "oiled" with powdered graphite mixed with vaseline. Lead pencils are made by grinding the graphite to a fine powder, washing and pressing it into proper shape to fit the wooden case. Clay in varying proportions is mixed in with the graphite to produce the different degrees of hardness of the pencils.

66. General Properties.—Carbon is heavier than water, almost infusible, and insoluble in all known liquids except molten iron. Melted iron can dissolve about five per cent. of carbon, but, on cooling, some of it separates out in the form of graphite. The forms of carbon vary in their ability to conduct electricity. Graphite is the best conductor, and, in the form of powder, is much used in electro-plating. All forms of carbon, when heated in an atmosphere of oxygen, burn with the formation of carbon dioxide.

67. Amorphous Varieties. Artificial.—CHARCOAL is made by the dry distillation of wood; that is, the wood is heated without the presence of air. Charcoal is heavier than water, although it is the lightest form of carbon. Freshly heated charcoal absorbs large quantities of such gases as ammonia, oxygen and sulphuretted hydrogen; therefore it is frequently used to disinfect places that emit offensive odors. Really, it is merely a deodorizer.

68. Coke is made by heating bituminous coal to red heat without the presence of air. It is used in the manufacture of iron. LAMP BLACK is ordinary soot, which, mixed with boiled linseed oil and a little soap, is used as printer's ink. BONE BLACK is made by heating bones to red heat without the presence of air. It consists of only about ten per cent. of carbon, the rest being bone ash (calcium phosphate).

69. Natural.—COAL contains carbon, hydrogen, oxygen, and small quantities of sulphur and nitrogen, along with silica. It is said to be the result of the partial decay of organic matter in nature, at the same time having been subjected to great pressure, and possibly heat, in certain cases. The purest form of coal is anthracite or hard coal. The next in order of purity is the bituminous or soft coal. Cannel and brown coals contain even more oxygen and hydrogen. Peat is only partially decomposed swamp moss.

EXPERIMENTS.

I. Put some fragments of wood in a test tube, and heat the tube until no more gas is given off. The best grade of charcoal is prepared by heating wood in closed iron retorts, the gases being allowed to escape through a pipe.

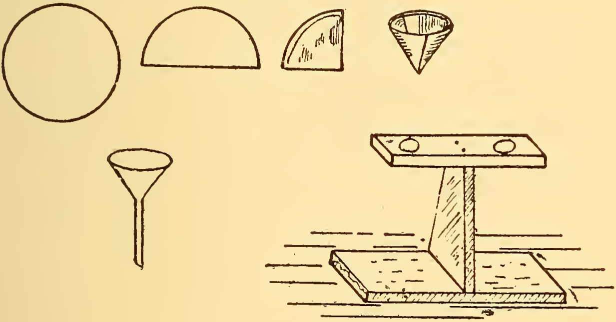


Fig. 8.

II. Compare the combustibility of charcoal and graphite by heating them separately upon a small piece of platinum.

III. Throw a piece of charcoal upon water; it floats. Heat a similar piece quite hot and plunge it suddenly under the water; it sinks. Why do you heat the charcoal?

IV. Fill a test tube one-third full of water; add several drops of ink, and then some coarsely powdered bone-black. Place the thumb over the mouth of the tube, shake vigorously, and filter. Repeat the experiment, substituting a solution of brown sugar. To filter, fold a round piece of porous paper, as shown in Fig. 8, and fit it snugly in a funnel. Wet with water, and pour the solution into the funnel.

LESSON XII.

NOMENCLATURE. CHEMICAL EQUATIONS. VALENCE.

70. Elements.—So far, we have studied five different kinds of elementary matter—hydrogen, oxygen, nitrogen, carbon and chlorine. There are about seventy different kinds of elementary substances known to chemists. These substances have never been decomposed into anything else; they may be some day. Some new elements may be discovered that will have to be added to the list, which is shown on the next page. These elements have different and fixed atomic weights as shown in the table. As we learned before, hydrogen is the lightest substance known to chemists, hence it is taken as the standard of atomic weights; that is, if a liter of hydrogen weighed one gram, a liter of oxygen would weigh sixteen grams, if the gases were under equal pressure, and at the same temperature. In most instances, the names of the elements were given them by their discoverers. The name was sometimes selected to denote a peculiarity the element possessed, or to indicate the source from which it can be derived, and in other cases the name is the legacy from the alchemist, and is apparently without meaning.

71. Symbols.—In arithmetic, we use figures to indicate numbers. In a similar way, in chemistry, we use a symbol to designate the element. It is merely a convenience in denoting the element without writing out the whole name. The initial letter of the Latin name is the usual symbol, though not always, as K. for kalium (Latin for potassium), or H. for hydrogenium. Where there are several having the same initial letter, one of the elements is given the initial letter for its symbol, and each of the others adds another letter from its name to the initial letter for its symbol, as C. for carbon, Cl. for chlorine, Cu. for cuprum or copper. The initial letter is always written with a capital and the other a small letter. In writing a compound we allow the symbols of the elements to follow one another in regular order, as HCl. for hydrochloric acid. Such a formula represents the composition of one molecule of the compound. When an element enters more than once into a compound—*i. e.*, has more than one atom in the molecule of the compound—we designate its repetition by a small numeral below, and to the right of the symbol, as H_2O instead of HHO, or HNO_3 for HNOOO (nitric acid). When the whole compound is taken several times, it is so stated by a numeral just before the symbols, indicating the compound as 4HCl, or $5NH_3$ (ammonia). Do not mistake the memorizing of a lot of symbols for learning chemistry. A thorough knowledge of them will prove of great assistance, but it is not absolutely essential.

72. Valence.—Chlorine, we remember, combines with only one hydrogen atom forming hydrochloric acid (HCl); oxygen with two in water (H_2O); nitrogen with three in ammonia (NH_3); and our last element, carbon, combines with four hydrogen atoms in the compound methane (CH_4), which we have not yet studied. If we arrange them in order, making use of the symbols, we have:

ClH . . . OHH . . . . NHHH . . . . CHHHH.

The number of hydrogen atoms entering into these compound is 1, 2, 3, and 4; so we may say that the VALUE, as we may term it, of chlorine in hydrogen atoms is one, of oxygen two, of nitrogen three, and of carbon four. One atom of chlorine is combined with one atom of hydrogen, therefore we may say that they have the same value. Suppose a hydrogen atom be represented by a nickel, and a chlorine atom by a silver five cent piece. They have the same value. Now if an oxygen atom be represented by a dime, we know that it is worth two nickels or silver five cent pieces. Exactly; oxygen combines with two atoms of hydrogen in water (H_2O) and two of chlorine in chlorine monoxide (Cl_2O). Nitrogen may be regarded as worth fifteen cents in terms of nickels, NH_3 (ammonia), or silver five cent pieces, NCl_3 (nitrogen trichloride), or one dime and a nickel, (NOH) (hyponitrous acid). Further, if carbon be put at twenty cents, it may be worth four nickels, (CH_4) (methane), or four silver five cent pieces (CCl_4) (carbon tetrachlorides), or part of one and part of the other, as CH_3Cl, CH_2Cl_2, $CHCl_3$ (chloroform). To carry this a little further, an oxygen atom, a dime, may be exchanged for two nickels or silver five cent pieces, and we have COH_2 (formaldehyde), or it may be expressed in terms of dimes, for CH_4 we have CO_2 (carbon dioxide).

This property of elements is termed VALENCE or ATOMICITY, hence chlorine is univalent, oxygen bivalent, nitrogen trivalent and carbon quadrivalent, the Latin prefixes being used to designate the figures. Valence is merely the expression of the value of chemical affinity, without regard to its strength.

73. Chemical Equations.—When we wish to write out the effect of one chemical upon another, we use the symbols, placing them exactly as if they were in an algebraic equation. It is a very great saving to use the following :

$$HCl + NaOH = NaCl + H_2O,$$

instead of writing out that "hydrochloric acid acting upon sodium hydroxide produces sodium chloride and water." Now, then, we notice that if we take the atomic weights corresponding to these elements and place them in an equation, the sums of the numbers on each side of the equality marks are equal, as should be. In all chemical equations this is one of the proofs of accuracy.

$$\begin{array}{l}
H \quad Cl \quad + \quad Na \quad O \quad H \quad = \quad Na \quad Cl \quad + \quad H_2O \\
(1+35.5)+(23+16+1)=(23+35.5)+(2+16)=76.5.
\end{array}$$

It is a little trying at first to memorize any of these symbols, but it becomes quite easy with practice.

EXPERIMENTS.

I. Place equal amounts of powdered sulphur and iron filings in a tube, about half a teaspoonful each, and heat to dull red heat for ten minutes. Allow to cool and save.

TABLE OF ELEMENTS, THEIR SYMBOLS AND ATOMIC WEIGHTS.

From the latest report (1898) of Prof. F. W. Clarke, Chairman of the Committee on Atomic Weights, American Chemical Society.

Name of Element.	Symbol.	Atomic Weight.
Aluminum	Al.	27.11
Antimony (Stibium)	Sb.	120.43
Argon	A.	?
Arsenic	As.	75.01
Barium	Ba.	137.43
Bismuth	Bi.	208.11
Boron	B.	11.95
Bromine	Br.	79.95
Cadmium	Cd.	111.95
Calcium	Ca.	40.07
Carbon	C.	12.00
Cerium	Ce.	139.35
Cesium	Cs.	132.89
Chlorine	Cl.	35.45
Chromium	Cr.	52.14
Cobalt	Co.	58.99
Columbium	Cb.	93.73
Copper (Cuprum)	Cu.	63.60
Erbium	Er.	166.32
Fluorine	Fl.	19.06
Gadolinium	Gd.	156.76
Gallium	Ga.	69.91
Germanium	Ge.	72.48
Glucinum	Gl.	9.08
Gold (Aurum)	Au.	197.23
Helium	He.	?
Hydrogen	H.	1.00
Indium	In.	113.85
Iodine	I.	126.85
Iridium	Ir.	193.12
Iron (Ferrum)	Fe.	56.02
Lanthanum	La.	138.64
Lead (Plumbum)	Pb.	206.92
Lithium	Li.	7.03

Magnesium	Mg.	24.28
Manganese	Mn.	54.99
Mercury (Hydrargyrum)	Hg.	200.00
Molybdenum	Mo.	95.99
Neodymium	Nd.	140.80
Nickel	Ni.	58.69
Nitrogen	N.	14.04
Osmium	Os.	190.99
Oxygen	O.	16.00
Palladium	Pd	106.36
Phosphorus	P.	31.02
Platinum	Pt.	194.89
Potassium (Kalium)	K.	39.11
Praseodymium	Prd.	143.60
Rhodium	Rd.	103.01
Rubidium	Rb.	85.43
Ruthenium	Ru.	101.68
Samarium	Sa.	150.26
Scandium	Sc.	44.12
Selenium	Se.	79.02
Silicon	Si.	28.40
Silver (Argentum)	Ag.	107.92
Sodium (Natrium)	Na.	23.05
Strontium	Sr.	87.61
Sulphur	S.	32.07
Tantalum	Ta.	182.84
Tellurium	Te.	127.49
Terbium	Tb.	160.00
Thallium	Tl.	204.15
Thorium	Th.	232.63
Thulium	Tu.	170.70
Tin (Stannum)	Sn.	119.05
Titanium	Ti.	48.15
Tungsten (Wolfram)	W.	184.83
Uranium	U.	239.59
Vanadium	V.	51.38
Ytterbium	Yt.	173.19
Yttrium	Y.	89.02
Zinc	Zn.	65.41
Zirconium	Zr.	90.40

LESSON XIII.

ACIDS. BASES. COMPOUND RADICALS. SALTS.

74. Acid and Basic Oxides.—All the elements mentioned in the table except a few, argon, neon, crypton, metargon (all recently discovered), helium, bromine, and fluorine, combine with oxygen to form compounds termed OXIDES. If the oxide combines with water to form an acid compound, as sulphur trioxide (SO_3), we have an acid forming oxide. If the compound formed is a BASE, as sodium monoxide (Na_2O) and water, we have a basic oxide.

75. Electric Conduct.—There are many substances that are decomposed by the electric current. These are called electrolytes. That portion of the substance which separates at the positive pole (see note, § 25, Lesson IV) is termed electro-negative, and that separating at the negative pole, electro-positive. As we learned that water is one of those substances we shall use it as our type, by which we may better understand the terms ACID and BASE.

76. Base.—The formula for water may be written in this wise: H–O–H. Hydrogen is electro-positive, as we have learned. All the metallic elements are, as a rule, more strongly electro-positive than hydrogen, so we may substitute one of them; for example, sodium, for a part of the hydrogen in water, and obtain Na–O–H, which is a *base*, or hydroxide. The bases are opposed to acids, we may say, as some of them turn litmus paper blue, while the acids, as a rule, turn it red.

77. Compound Radicals.—There are certain compounds which withstood for a long time all efforts made by chemists to decompose them. They acted very much like elements, but were eventually proven to be compounds instead. Substances of compound nature acting much like elements under ordinary circumstances, but capable of decomposition, are called COMPOUND RADICALS. Ammonium (NH_4) is one, and it may be substituted for the hydrogen in water as well as an element, so we have $NH_4{-}O{-}H$. Therefore, a *base* may be regarded as water in which one hydrogen atom has been substituted by an electro-positive element or radical.

78. Acid.—Now suppose we substitute the $-O-H$, called HYDROXYL, by an electro-negative element or radical; for example, chlorine, we have HCl, or by $-NO_3$, when we have HNO_3 (nitric acid). Thus we see that an *acid* may be regarded as water in which the hydroxyl has been substituted by an electro-negative element or radical.

79. Salts.—Bring together an acid and a base, as sodium hydroxide and hydrochloric acid or ammonium hydroxide and nitric acid, and we have a salt produced and water formed. The equations are as follows:

$$\overset{\text{Base.}}{NaOH} + \overset{\text{Acid.}}{HCl} = \overset{\text{Salt.}}{NaCl} + \overset{\text{Water.}}{H_2O}, \text{ and}$$

$$\overset{\text{Base.}}{NH_4OH} + \overset{\text{Acid.}}{HNO_3} = \overset{\text{Salt.}}{NH_4NO_3} + \overset{\text{Water.}}{H_2O}.$$

Such a compound is a SALT. Salts, as a rule, are neutral towards litmus paper, but not always so. Some are soluble in water, and each possesses its own peculiar taste. Some are insoluble in water and also in acids. The result of the union of an acid forming oxide and a base forming oxide is likewise a salt. Example: $SO_3 + Na_2O = \overset{\text{Salt.}}{Na_2SO_4}$. No water is eliminated in this case, however.

80. Chemical Terms.—When we separate a compound into its constituents, as by the action of sodium upon water, we call it ANALYSIS. If we combine elements to make a compound, as in burning hydrogen in oxygen to produce water, we call it SYNTHESIS or building up. If an interchange of the elements of two compounds takes place to produce two new compounds, as shown above in the formation of salts from acids and bases we have METATHESIS, or double decomposition.

EXPERIMENTS.

I. In a porcelain evaporating dish pour 5 cc. of a dilute (about ten per cent.) solution of NaOH. While stirring add to it dilute HCl until a piece of litmus paper is no longer turned blue by a drop of the solution placed upon it with a glass rod. If the solution finally turns the paper red, add a little more of the NaOH solution. Repeat, adding first one and then the other until the litmus paper is no longer affected by the solution. Evaporate most of the liquid by heating over a lamp and allow to cool. State what takes place, writing equation, giving reason for evaporating, and finally taste the substance obtained.

II. Repeat the first experiment, using ammonium hydroxide and nitric acid, instead of sodium hydroxide and hydrochloric acid.

LESSON XIV.

CARBON MONOXIDE AND DIOXIDE.

81. Carbon Monoxide.—Carbon forms two compounds with oxygen, carbon monoxide (CO) and carbon dioxide (CO_2). Carbon monoxide is a colorless, odorless gas, which burns with a pale, blue flame. The lambent blue flame hovering over ignited anthracite coal is caused by the burning of this compound to form the higher oxide, carbon dioxide. The monoxide is very poisonous, and is a prominent constituent of "water gas." Water gas is produced by passing steam over very hot coke. The water is decomposed into free hydrogen, and the oxygen unites with carbon to form the monoxide. This kind of gas burns with a non-luminous, but intensely hot, flame.

82. Law of Multiple Proportions.—You remember that we learned that a chemical compound always has the same elements entering into it in the same proportion. We also learned that nitrogen forms five different compounds with oxygen; they have the same elements entering into them, but in different proportions. We have a similar case to-day with carbon. We have the monoxide and the dioxide. There is a law in chemistry that when one element combines with the same element in different proportions to form different compounds, the amounts of the varying element bear the relation of simple multiples to each other. For instance, CO and CO_2 have oxygen, the varying element, present in the ratio of 1 : 2; again with the oxides of nitrogen, we have N_2O, N_2O_2, N_2O_3, N_2O_4 and N_2O_5, that is, 1 : 2 : 3 : 4 : 5.

83. Carbon Dioxide.—Carbon dioxide (CO_2) is also a colorless, odorless gas. It is produced when carbon is burned in an excess of oxygen or air. The gas given off from the flame of a candle or a lamp is carbon dioxide, as may be shown by passing the gas through lime water. The white solid substance, which is precipitated, is calcium carbonate ($Ca\,CO_3$). When two clear solutions are brought together and a solid is produced, we say it is "precipitated."

84. Result of Oxidation.—Carbon dioxide is given off by the lungs in breathing. Its formation is due to the oxidation of the carbon in organic substances in the body by the oxygen which is breathed into the lungs. Carbon dioxide does not support combustion, nor does it burn when fire is applied to it. It cannot support life. An animal would die of suffocation if allowed to breathe nothing but carbon dioxide; it is not poisonous, however.

85. Properties—Carbon dioxide is slightly soluble in cold water, but more soluble when under pressure, forming what is known as soda water. The solution is called "carbonic acid water," but is really carbonic acid in solution. CO_2 combines with H_2O to produce the hypothetical carbonic acid (H_2CO_3), for it has never been separated. Carbon dioxide may be liquefied by great pressure, and is transported in that form in iron cylinders. When this liquefied gas is allowed to evaporate quickly, intense cold is produced, sufficient not only to freeze water and mercury, but to cause some of the carbon dioxide itself to be solidified into a snow-like mass.

86. Heavier than Air.—It is half again as heavy as air, and should be collected by downward displacement. The gas can be poured from one vessel into another like water, only we do not see the gas as we do the water. Being heavier than the air, this gas, when liberated in large quantities, as it is by some volcanoes, settles down in valleys and remains there, unless disturbing winds displace it. Several places of this kind are known, and are termed "Death Valleys."

87. Wide Distribution.—All combustible material containing carbon, which burns, produces carbon dioxide, thus a great deal of this gas must be given off every hour of the day by the huge furnaces and forest fires. Although the quantity thus produced is enormous, it seldom constitutes in any locality more than about four parts in ten thousand of the atmosphere; yet that amount, the volume of the atmosphere being so great, is more than would be produced if the result of the combustion of all the vegetable and animal matter in the world were uniformly distributed throughout the air.

88. Preparation.—Carbon dioxide is most easily prepared for experimental purposes by treating calcium carbonate (common marble or limestone) with dilute hydrochloric acid.

EXPERIMENTS.

I. Prepare a small bottle of oxygen. Ignite a splinter of charcoal bound to a wire, and insert in the bottle. Remove after the combustion has ceased and introduce a stirring rod moistened with clear lime water. Do not touch the sides of the vessel with the rod.

II. Set up apparatus like that shown in Fig. 9. Gently suck the short tube, holding the funnel over a burning candle or lamp. What is the white precipitate?

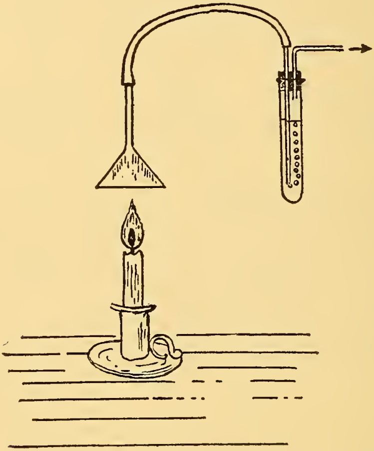


FIG. 9.

III. Put a dozen pieces of marble in apparatus like that shown in Fig. 5. Add HCl. Try to set fire to the gas evolved. Test some with lime water. Collect a large bottle full of the gas and pour it into a wide mouth vessel containing a short, lighted candle.

LESSON XV.

SULPHUR. HYDROGEN SULPHIDE.

89. Brimstone.—So far we have discussed but one solid element, and that one is usually colorless. To-day we have a yellow element. Sulphur occurs free in nature around volcanoes as Mt. Etna, in Sicily, and elsewhere. It has been known for centuries as "brimstone," having been found in large quantities near the Italian and Sicilian volcanoes. It also occurs in large amounts combined with iron, forming a common mineral, iron pyrites, or "fool's gold"; also as galena, a sulphide of lead (which is an important ore), and in similar compounds.

90. Properties.—Sulphur is a yellow brittle solid, which melts a little above the boiling point of water to a clear yellow liquid. On further heating the light liquid changes to a red viscous mass, which, on further heating, liquefies again before it boils. It boils just like water, only at a very much higher temperature, 440°C.

91. Varieties.—Sulphur exists in two main forms, crystalline and amorphous. The ordinary sulphur, moulded in sticks, called "roll sulphur," and the powdered form, "flowers of sulphur," are crystallized. If we add a strong acid to a solution of a soluble sulphide, as of potassium, we get a milky liquid. This milkiness is due to finely divided sulphur, the amorphous variety. Sulphur is insoluble in water, but most of its varieties are soluble in a compound of carbon and sulphur, called carbon di-sulphide.

92. Amorphous.—If we heat sulphur in a test tube until it becomes a dark red liquid, and quickly pour it into cold water, it assumes an India rubber appearance. This is also an amorphous form of sulphur, called "plastic sulphur." It changes to the crystalline variety when kept for awhile however.

93. Compounds.—Sulphur burns with a blue flame, noticed in sulphur matches. The disagreeable smelling, colorless gas given off is an oxide of sulphur, the dioxide (SO_2), commonly, but incorrectly, called sulphurous acid. The compound of sulphur with hydrogen is hydrogen sulphide, commonly called sulphuretted hydrogen.

94. Sulphuretted Hydrogen.—If we heat a mixture of iron filings and sulphur in a tube to a low red heat, we cause the sulphur and iron to combine, forming a sulphide. Allow the tube to cool and add an acid, hydrochloric for example, to this black solid, and a gas will be given off. This gas has a very disagreeable odor; in fact, the bad odor of rotten eggs is due largely to hydrogen sulphide (H_2S).

95. Sulphides.—Sulphuretted hydrogen burns with a pale blue flame. When the gas is passed through acid solutions of many metallic salts, as of copper, arsenic, and lead, the sulphides of these metals are formed.

The sulphides have various colors, hence their production serves as tests for the presence of certain metals.

96. Sulphur Waters.—Sulphuretted hydrogen is quite soluble in water, the solution being prepared by bubbling the gas through pure water about ten minutes. The waters of sulphur springs contain this gas in solution. On shaking with air the H_2S is decomposed with the formation of free sulphur and water. Many tons of sulphur are annually produced artificially by decomposing the sulphuretted hydrogen obtained from calcium sulphide, a by-product in the manufacture of soda by the LeBlanc process, which we shall study later.

97. Similarity to Water.—Under certain circumstances sulphur has the same valency as oxygen. One volume of gaseous sulphur combines with two of hydrogen to produce sulphuretted hydrogen, hence its formula is H_2S. The sulphur seems to take the place of oxygen in water, H_2O, hence it is bivalent. Again, carbon disulphide has the formula CS_2 exactly like carbon dioxide, CO_2. Oxygen and sulphur resemble each other chemically very much, though physically there is no resemblance.

EXPERIMENTS

I. Burn a small piece of sulphur on a piece of porcelain. Note color of the flame and odor of the gas given off.

II. Melt a spoonful of sulphur in a test tube. Note all the changes carefully. When the melted sulphur has become deep red, quickly pour it into a glass of cold water Examine it closely with eye and hand.

III. Dissolve a little sulphur in some CS_2 and pour it out into an open vessel and allow the light liquid to evaporate.

IV. Place some of the ferrous sulphide prepared in Lesson XII. in bottle "a," Fig. 10, and add 30 cc. dilute hydrochloric acid. In bottle "b" have pure water. After the gas has bubbled for five minutes allow the gas from tube "c" to bubble in turn through test tubes containing acidified water solutions of copper sulphate, lead nitrate, antimony potassium tartrate, and sodium arsenite.

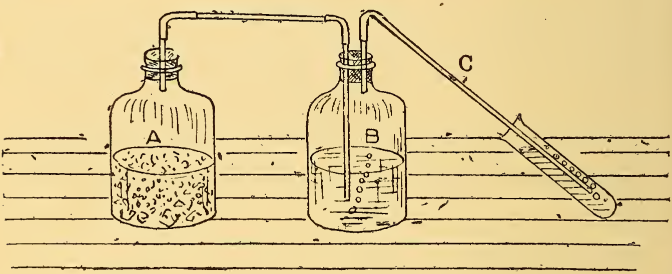


FIG. 10.

NOTE.—*Perform all experiments with this gas under a hood or in the open air.* Dissolved in water the gas has the name of hydrosulphuric or sulphhydric acid. Although the gas is poisonous when breathed, its water solution is an efficacious remedy when taken internally.

V. After smelling the water solution prepared in Experiment IV., test it with blue litmus paper; then shake violently a little of the solution in a stoppered bottle.

LESSON XVI.

SULPHUR DIOXIDE. SULPHURIC ACID.

98. Sulphur Dioxide.—Sulphur burns in the air or oxygen to produce a gas possessing a suffocating and irritating odor. This is SULPHUR DIOXIDE (SO_2), and is soluble in water, the solution being called sulphurous acid (H_2SO_3). The gas is not combustible, nor a supporter of combustion, and is poisonous to both animals and plants. As there is much sulphur in some of the bituminous coals, a great amount of sulphur dioxide is sent out in the air where such coals are burned. It is washed down by the rains, and causes great damage to growing plants. In Cleveland, Ohio, over nine thousand tons of this gas are liberated in the air every year. Sulphur dioxide can be best prepared in the laboratory by heating concentrated sulphuric acid with metallic copper.

99. Bleaching.—Sulphur dioxide destroys certain vegetable colors, and is used for bleaching straw from which hats are made. Its action, which requires the presence of water, is only temporary, however, as straw hats usually turn yellow on exposure. As it kills the disease-producing germs, it is used a great deal in disinfecting houses where infectious diseases have existed.

100. Sulphuric Acid.—Sulphur dioxide can be made to unite with another atom of oxygen, producing sulphur trioxide (SO_3), which is very soluble in water. This solution is SULPHURIC ACID, which, in concentrated form, is commonly called "oil of vitriol" (H_2SO_4), because it was originally made from green vitriol, an iron sulphate (our ordinary copperas).

101. Immense Amounts Used.—The method of manufacturing sulphuric acid is quite complicated. Large amounts of the acid are manufactured, the

United States yielding alone about nine hundred thousand tons annually. The financial and commercial progress of a nation may be judged by the yearly production of sulphuric acid. There is hardly a substance that we make use of in daily life—food, clothing, or what not—that is not directly or indirectly dependent upon the use of sulphuric acid in its production.

102. Properties.—Sulphuric acid is a heavy, oily, colorless liquid when pure, and is soluble in water in all proportions. Ordinary sulphuric acid is colored brown on account of impurities present, some of which are lead, arsenic, organic matter, and the oxides of nitrogen. On mixing the ordinary concentrated acid, ninety-five per cent., with water, an immense amount of heat is developed. Great care should be exercised in mixing. The acid should *always* be slowly poured into and well stirred with the water, and *never* the reverse.

103. Attraction for Water.—The concentrated acid greedily absorbs water from all substances. It has such a strong affinity for water that it will decompose sugar to get the water from it, and leave only black carbon behind. The acid acts upon all organic substances, which it attacks by taking water away from them. It blisters the skin. It is sometimes used by criminals to injure their enemies.

EXPERIMENTS.

I. Burn some sulphur under a large, colorless, bottomless bottle, within which are some moist, bright-colored flowers (Fig. 11).

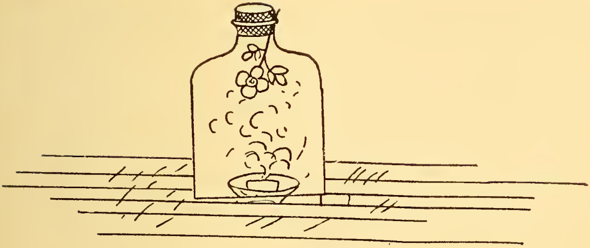


FIG. 11.

When the experiment is completed, note the odor of the gas and test its combustibility. Next, remove the cork and insert a glass rod, to which is bound a tuft of cotton saturated with concentrated nitric acid. State your conclusion.

II. Place 30 c. c. water in a beaker. A thick walled vessel will not do. Add, in a fine stream, 70 c. c. of concentrated sulphuric acid, stirring with a test tube half filled with alcohol, which has been colored with a little ink.

III. Dip a splinter of wood into concentrated sulphuric acid for a few seconds.

IV. Make a thick syrup by dissolving sugar in water in a test tube. Cool, and carefully add concentrated sulphuric acid.

V. Use a glass stirring rod for a pen, and write upon filter paper with concentrated sulphuric acid for ink.

LESSON XVII.

PHOSPHORUS. BASICITY OF ACIDS.

104. To-day we have an element which is so devoted to oxygen that it is never found in nature unless combined in some way with that element. So we do not find PHOSPHORUS free, but as phosphates in bones of animals and in certain phosphate rock. These phosphates always contain calcium, besides phosphorus and oxygen, along with some iron and aluminum. Large deposits of this rock are found, mined, and used as fertilizers. Phosphates in very small amounts, however, occur in all kinds, and is an essential constituent, of fertile soil. Most iron ores contain phosphorus also; some contain it in notable amounts. Like carbon, sulphur and oxygen, phosphorus exists in allotropic forms.

105. White Variety.—Ordinary phosphorus is a very dangerous substance to have anything to do with. It is prepared from bones by a complicated process. Pure phosphorus is waxy and almost colorless, but the kind usually seen, is more or less yellow in color, and frequently contains some arsenic and carbon as impurities. Waxy phosphorus is soluble in carbon disulphide, and oxidizes in the air, often melting and taking fire spontaneously, giving off white fumes possessing the odor of garlic. To prevent this oxidation, phosphorus, usually in stick form, is kept under water. It may be cut with a knife, and it should never be handled with the hands except under water. It is exceedingly poisonous, being used for a rat poison and to make matches. Phosphorus burns are intensely painful and difficult to heal. Bathe the parts as soon as possible after phosphorus burning with a solution of blue stone, copper sulphate This is only effective when some of the phosphorus is left unburned.

106. Insoluble Form.—*Red phosphorus* is another variety, the harmless kind, made by heating the waxy form in an atmosphere of nitrogen free from oxygen. It is red in color, insoluble in carbon disulphide; does not oxidize rapidly at the ordinary temperatures, and may be kept in a bottle without water. It is not poisonous, and does not take fire by friction, as does the waxy phosphorus. However, when heated to a high temperature, it changes back into the white form and takes fire at once. These allotropic forms of phosphorus have the most widely divergent properties.

107. Oxides.—Phosphorus forms several compounds, with hydrogen and oxygen. The former are very poisonous, and take fire spontaneously. Phosphorus trioxide (P_2O_3), is formed by the slow oxidation of phosphorus, as when ozone is prepared. Phosphorus pentoxide (P_2O_5) is produced when phosphorus burns rapidly, or with an excess of air. It is a white solid, possessing the most intense attraction for water, phosphoric acid (H_3PO_4) resulting on solution.

108. Phosphates.—The salts of phosphoric acid, called phosphates, are formed by the replacement of the hydrogen of the acid by metals. The important ones are calcium phosphate (bone phosphate), superphosphate, and the sodium and ammonium phosphates. Phosphates are the principal constituents of fertilizers.

109. Basicity of Acids.—There are acids, hydrochloric for instance, which contain no oxygen, but many acids have that element in them. At one time, it was thought that all acids contained oxygen, therefore, that element received its name from the Greek, meaning "acid producer." On adding water to the acid forming oxides, as we have learned, we obtain the acids.

$N_2O_5 + H_2O = 2HNO_3$—Nitric acid.
$2SO_3 + 2H_2O = 2H_2SO_4$—Sulphuric acid.
$P_2O_5 + 3H_2O = 2H_3PO_4$—Phosphoric acid.

Comparing these three acids HNO_3, H_2SO_4 and H_3PO_4, we note that we have 1, 2 and 3 hydrogen atoms present respectively. One or all of these hydrogen atoms may be replaced by a metal or electro-positive radical to form salts. These acids are therefore termed respectively mono-, di- and tri-basic.

110. Salts.—When all the hydrogen atoms are replaced, as $NaNO_3$, K_2SO_4 and $(NH_4)_3 PO_4$, we have a NEUTRAL SALT, chemically speaking. The fact of the matter is, that the last named salt is actually alkaline to litmus paper, but it is neutral chemically, because all the hydrogen has been replaced; that is, it is no longer acid. If only a portion of the hydrogen be substituted, we have an *acid salt*; that is, a salt with acid properties, for example, $NaHSO_4$ or $CaHPO_4$. Calcium is bivalent, hence replaces two hydrogen atoms. It is obvious from the above, that hydrochloric or nitric acid can form only neutral salts, as from one of our first definitions an atom can not be divided. All the hydrogen need not necessarily be replaced by the same metal, but the metals or radicals may be different, giving DOUBLE SALTS, as $(NH_4)MgPO_4$, ammonium magnesium phosphate, magnesium being bivalent.

EXPERIMENTS.

I. Dissolve a little ordinary phosphorus in carbon disulphide in a test tube and pour the liquid upon a filter paper. Lay it on an iron stand to dry.

II. Touch a small piece of phosphorus with a test tube filled with hot water.

III. Place a small piece of phosphorus in a dry test tube, drop in a tiny crystal of iodine and warm. The re-action begins at once. A small amount of phosphorus iodide is volatilized and the rest is changed to the red variety. Cautiously add water, pour off, and test the solubility of the red phosphorus remaining in carbon disulphide. Does it take fire as easily as the other kind?

IV. Set fire to a small piece of phosphorus in a bottle with a hot wire. Close the bottle loosely, and after the reaction is over, insert a piece of moist, blue litmus paper into the bottle.

LESSON XVIII.

ORGANIC CHEMISTRY.

111. Old Divisions.—Carbon combines with hydrogen in different proportions to form thousands of compounds. The number is so very large that a certain division of chemistry treats alone of those compounds. It is called ORGANIC CHEMISTRY. Years ago chemistry was divided into three great divisions—chemistry of the mineral, animal, and vegetable worlds. It was held that distinct and separate laws governed each kingdom. Later, the barrier between the animal and vegetable chemistry was torn down, but it was still maintained that the laws of inorganic or dead nature could not hold for substances which were the products of vital force. In fact the religious fanatics, who ruled all intellectual life of the day, regarded it as sacrilegious to attempt the artificial production of substances ordinarily found only as the result of the action of life.

112. Important Date.—In 1828 a great chemist, named Wöhler, proved that an organic body (urea) could be produced in the laboratory from inanimate material. From that time on hundreds of chemists have worked in that field. For convenience the results of their great labors are incorporated under one head, but the laws for Inorganic and Organic Chemistry are the same.

113. Vast Number of Organic Compounds.—Organic chemistry treats not only of the compounds of carbon and hydrogen, but of carbon and hydrogen with oxygen, sulphur, nitrogen, and phosphorus. All substances used for illuminating our homes, except electric light of course, such as petroleum, gas, and oils, kerosene, coal gas, and tallow candles, come under that head. It also treats of the composition of plant fiber, clothing materials, foods, animal juices, muscles, etc. Many medicines, as antipyrine, quinine, morphine; poisons, as strychnine and nicotine, are organic compounds. Stimulating substances, as alcohol; sedatives, as chloral, ether, and chloroform; fats, vaseline and wax, are included in that division. Many of these substances are very simple in their composition, and some are very complicated in structure. Their technical names usually accord with their chemical composition, as "phenyl-dimethyl-pyrazolon"; but frequently there is a simpler name, as "antipyrine," which is used in commerce, but gives no indication of its structure or composition.

114. Some Properties of Organic Compounds.—All organic compounds may be burned with oxygen. In the decomposition the carbon burns to carbon dioxide, hydrogen to water, and nitrogen to the oxides, or escaping free. Many of the compounds burn very easily; some form violently explosive mixtures with oxygen. The flame resulting from the explosion is due to burning of a gas. Organic compounds, containing nitrogen when heated without oxygen, usually give off ammonia.

115. Safety Lamp.—Iron is one of the best conductors of heat, so a fine wire gauze conducts away the heat so rapidly that it does not permit inflammable gas on the other side to reach the kindling point. This principle, that flame will not pass through fine wire gauze, was made use of by Sir Humphrey Davy in his safety lamp. These lamps are now much used in coal mines, where explosions would occur as the result of the ignition of the combustible gases by the flame of an ordinary miner's lamp.

116. Luminosity.—Flame itself is colorless or pale blue, such as we see when an alcohol lamp is burning. The LUMINOSITY of the flame, however, is due to the intense heating of very small particles of substance or incandescent gases. The ordinary white or yellow light of a gas burner, lamp, or candle is due to finely divided carbon heated to a white heat. A platinum wire held in the colorless flame glows and becomes luminous. Colored flames result when other substances are substituted for carbon. The green fireworks are due to barium; yellow to sodium; and crimson to strontium.

117. Oxidizing and Reducing Flame.—The colorless flame is hotter than the luminous flame. In the yellow flame the carbon is not burned all at once, but is, finally, in the brightest flame. If it were not a smoky flame would result. If we can so arrange, by the introduction of more oxygen, that the combustion of the carbon is at once complete, the flame becomes colorless and very much hotter. This may be accomplished by a blowpipe. The blue flame, which shows an excess of oxygen, is called the OXIDIZING, while the yellow one, having more carbon than it can burn, is the REDUCING flame. The blowpipe promotes the burning and concentrates the heat at one point.

EXPERIMENTS.

I. Pour five drops of ether into a tin can. Cover it with a piece of paper. After ten minutes slip the paper just enough to leave a small opening. Standing at a distance apply a lighted taper to the opening.

II. Hold a piece of wire gauze over a gas burner. Turn the gas on and light it *above* the gauze. Now light the gas issuing from the burner, remove the gauze, and bring it down to the middle of the flame.

III. Sprinkle some powdered charcoal in a colorless flame, or file a nail or other metallic substance above it. Hold a cold piece of clean porcelain in the yellow flame for a few seconds.

IV. Prepare a glass tube by fusing a platinum wire into one end, as shown in Fig. 12. Drop a pinch of sodium chloride into a watch glass containing dilute hydrochloric acid. Moisten the wire with this solution, and hold in the colorless flame, which may be had by use of the plowpipe (Fig. 4, p. 34). Pour fresh acid into the watch crystal, and cleanse the wire by wetting with the acid and burning. Repeat until the flame is no longer colored. When the wire is clean repeat the experiment, using strontium and barium chlorides.

FIG. 12.

LESSON XIX.

MARSH GAS. CHLOROFORM. ALCOHOL.

118. Marsh Gas.—Who has followed a squirrel through the woods into a swamp and not noticed the bubbles of colorless gas rise as you trod upon the decaying leaves lying under water? That very gas is a compound of carbon and hydrogen, called METHANE, or "marsh gas." It is produced in these swamps by the decay of leaves and other organic matter. It also occurs in large amounts in some coal mines, being then termed "fire damp," because when mixed with the air and exposed to the flame it explodes with great violence. These explosions are very destructive to life and property.

119. Substitution Compounds.—Methane (CH_4) has one carbon atom and four of hydrogen to its molecule. Applying the principle of valency, we may substitute one chlorine atom for one hydrogen atom, as we remember both are univalent. In succession, chlorine may be caused to take the place of one, two, three, or even all four of the hydrogen atoms, giving compounds with the formulas, CH_4, CH_3Cl, CH_2Cl_2, $CHCl_3$, and CCl_4. The only one we need speak of is $CHCl_3$, called CHLOROFORM.

120. Preparation of Chloroform.—As we regard chloroform as methane with three hydrogen atoms substituted by an equal number of chlorine atoms, it may result from the direct action of the sunlight upon a mixture of these two gases. Chloroform, a colorless volatile liquid, may be more easily prepared by distilling dilute alcohol with chloride of lime. Chloroform is used to dissolve many substances, but it is called for mainly as an anæsthetic in surgery.

121. Organic Hydroxides.—Going back to methane, suppose one of the four hydrogen atoms to be substituted by the hydroxyl group, -O-H. The formula would be CH_3OH, and the compound, an ALCOHOL. This one is the so-called "wood alcohol." There is another hydrocarbon, a gas called ethane, very similar to methane, which has the formula, C_2H_6. We may regard it as two molecules of methane, less two hydrogen atoms, or methane in which one hydrogen atom has been substituted by CH_3—*i. e.*, $CH_3 . CH_3$.

122. Ethyl Alcohol.—Suppose -O-H be substituted for one hydrogen atom in this ethane. We have C_2H_5OH, or CH_3CH_2OH, the formula for the ordinary alcohol. Another way of looking at the formula is to regard alcohol as water, in which one hydrogen atom has been replaced by the compound radical, C_2H_5, called ethyl, as :

$$H–O–H \;.\;.\;.\;.\;.\;.\;.\; CH_3CH_2–O–H.$$

123. Fermentation. — Alcohol is ordinarily prepared by the FERMENTATION of sugar or starch. Fermentation is generally regarded as the result of the action of living organisms, which possess the power of breaking up such substances as grape sugar, giving off carbon dioxide. The equation :

$$\underset{\text{Grape sugar.}}{C_6H_{12}O_6} = \underset{\text{Alcohol.}}{2C_2H_5OH} + \underset{\text{Carbon dioxide.}}{2CO_2},$$

expresses this decomposition. The ferment is added to a weak solution of molasses or sugar in water and the surface of the liquid exposed to the air, as the germs require oxygen for existence and growth. The organisms are most active at the temperature of a warm summer day. The alcohol produced is dissolved in the water present. As alcohol boils at 78° C, and water at 100° C, they may be partially separated by distillation, the alcohol vapor passing off first. ABSOLUTE alcohol contains no water. It is obtained from the ordinary alcohol, which is ninety-five per cent. alcohol, by treatment with quick lime, which is slaked by the water present, and redistilling the alcohol from the slaked lime formed.

124. Properties. — Pure alcohol has a peculiar, pleasant odor, and burns with a non-luminous hot flame. It has a great attraction for water. Alcohol is the principal solvent for substances of organic origin. It is extensively used in the arts, as in the manufacture of varnishes, perfumes, and the tincture of drugs.

125. Not a Food.—In the dilute form, for it can be mixed with water in all proportions, it acts upon the human system as an intoxicant; when concentrated, or taken in large quantities, it is a poison. When alcohol is taken internally a sense of warmth is experienced, but in fact the temperature of the body is actually lowered from one-half to two degrees Fahrenheit. Many ignorantly suppose they can endure greater cold, or cold for a greater while, if they fortify themselves with alcohol. Such is a mistake, as proven by the experience of the members of various Arctic expeditions. Those who eschewed the use of brandy or whiskey withstood the intense cold better than those who had recourse to it. Alcohol is not a food in the sense we ordinarily think of such, and should not be so considered.

126. Beverages.—The many beverages in use depend for their efficacy upon the amount of alcohol that is present. The milder forms of beer, some European beers, for example, contain from two to three per cent.

of alcohol. American beer has from six to eight. Beers, besides alcohol, contain as a rule some nourishment. Light wines, as claret, have eight per cent., and whiskey and brandy sometimes contain as much as sixty, or even seventy per cent. of alcohol. These drinks are nothing but solutions of alcohol of various strengths with small amounts of different kinds of essences obtained from the fruit or grain from which the liquor is made. The evil effect of these strong drinks is so well taught in physiology that it is almost needless to warn one against their use. Suffice it to say, that knowing that the snake will bite, one is a fool to put his hand upon it. Most all the ill health, and nine-tenths of the unhappiness of the world, is the direct result of the abuse of the use of alcoholic stimulants. Like many poisonous substances, alcohol, when properly used under the direction of a physician, is a good medicine.

EXPERIMENTS.

I. Burn a little alcohol in a porcelain dish. Note flame. Add to the alcohol in turn a solution of sodium, strontium and barium chlorides. Burn in dish, washing the dish well after each experiment and using fresh alcohol each time.

II. Pour a little alcohol in the palm of the hand. How do you account for the sensation felt?

III. Half fill a test tube with water. Cautiously fill the tube to overflowing with alcohol. Avoid mixing the two liquids during the pouring, then place the thumb over the mouth of the tube. Mix well by shaking. Remove the thumb without spilling any of the liquid. What do you notice? Explain.

IV. Place a little of the white of a fresh egg in test tubes. In one add 1 cc. alcohol; in the other, 5 cc. whiskey.

NOTE.—It has been learned recently that fermentation may occur without the presence of germs.

LESSON XX.

ETHERS. ORGANIC ACIDS. ESTERS.

127. Ethers. Class.—Suppose we subtract a molecule of water from two molecules of alcohol; we have

$$\begin{matrix}CH_3CH_2OH\\ CH_3CH_2OH\end{matrix} - H_2O = \begin{matrix}CH_3CH_2\\ CH_3CH_2\end{matrix}{>}O$$

a body called an ETHER. Alcohol is regarded as an hydroxide of ethyl; ether is an oxide of ethyl. Let us go back to water to make it clear.

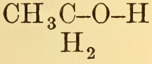
Water.	Hydroxide.	Oxide.
H–O–H	Na–O–H	Na–O–Na
	Alcohol.	Ether.
	$CH_3CH_2–O–H$	$CH_3CH_2–O–CH_3CH_2$.

Sulphuric acid has the power of tearing water away from certain substances, hence we may use that reagent for removing water from alcohol to produce ether.

128. Properties.—Ether is a colorless, very light liquid, which boils ever so easily (at 34.9° C.) In fact, on a hot summer day, it exists as a gas; hence it evaporates with great rapidity, producing cold. It is very inflammable, and possesses a peculiar odor and taste. It is an excellent solvent for resins, fats and many organic substances. It has great value and use as an anæsthetic.

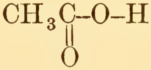
129. Organic Acids.—In the production of alcohol by the fermentation of sugar, it was necessary to cut short the operation before the ferment had done all the damage possible. If the reaction, resulting from the ferment, be allowed to continue, the alcohol will be further oxidized and changed to ACETIC ACID. Suppose we imagine alcohol with two hydrogen atoms substituted by one oxygen atom, we have,

Alcohol.



$$CH_3\underset{H_2}{C}-O-H$$

Acetic Acid.



$$CH_3\underset{\overset{\|}{O}}{C}-O-H$$

130. Acetic Acid.—Acetic acid is an organic acid, soluble in water in all proportions. All organic acids have the group –CO.OH, called CARBOXYL, present in them. An impure form of acetic acid, containing about four per cent. of the acid, is vinegar. Acetic acid is prepared on the large scale in two ways—by the oxidation of alcohol and by the distillation of wood. Pure concentrated alcohol, when exposed to the air, does not undergo change. Poor wines containing small per cents. of alcohol, however, on exposure to the air, gradually become sour. This is called "wine vinegar," and the oxidation of the alcohol is due to the action of the ferment called "mother-of-vinegar." As great quantities of vinegar and acetic acid are used annually, a "quick vinegar process" has been invented. It consists essentially of barrels filled with beech wood shavings, which are covered with the "mother-of-vinegar." Dilute alcohol is allowed to trickle down through the shavings, the air passing in the opposite direction. The liquid which went in at the top as alcohol emerges at the bottom as vinegar.

131. Organic Salts.—Alcohols act like hydroxides, as they really are, and the organic acids behave as ordinary acids do; therefore they may combine to form ORGANIC SALTS. Some of these organic salts, called ESTERS, often wrongly called "ethers," are among the most important substances of which we make use in daily life.

132. Saponification.—FATS, as lard, beef fat, and butter, for the most part consist of these esters. They are composed of glycerol, commonly termed glycerine, which is an alcohol, combined with the complicated acids, palmitic, stearic, and oleic. Fats may be decomposed into glycerol and the acids, if boiled with dilute acids or alkalies; the process is called SAPONIFICATION, the act of making soap, for such it is.

133. Soaps.—The hydroxides of some metals, as potassium or sodium, combine with the organic acids just as they combine with what are known as the mineral acids—hydrochloric, for example. By boiling fat with potash we obtain glycerol and a mixture of potassium palmitate, stearate, and oleate, or SOFT SOAP. If sodium be used instead of potassium, we have HARD SOAP, which is less soluble in water than the former. Now, calcium palmitate and stearate are quite insoluble in water. Hard water, as we learned, contains calcium salts in solution, which immediately form a calcium palmitate or stearate as soon as the soap is placed in the water. Consequently, hard water does not at first form a lather with soap. Soap is easily and cheaply made at home by saving all the wood ashes in a barrel and then treating them with water. This leaching dissolves out the potassium carbonate, which, when boiled with refuse fat, otherwise to be thrown away, gives soft or kitchen soap. Expensive soaps are usually made of purer material, colored, perfumed, and moulded into cakes of convenient size.

134. Butter.—BUTTER is a mixture of fats as produced by the various milk-giving animals. In addition to the fats mentioned above, it usually contains butyrates also. If the butyrates are acted upon by certain ferments causing partial decomposition, rancid butter is the result. Other ferments give the sweet odor to butter. OLEOMARGARINE is an artificial mixture of fats (palmitates, stearates, and oleates), colored yellow as a rule, and is frequently substituted for butter. If these fats were pure originally, there could be no objection to the use of oleomargarine as a substitute for butter, only it should be sold as such, and not as butter.

EXPERIMENTS.

I. Pour a little ether in the palm of the hand, noting volatility, cold, and odor.

II. In a test tube place a little butter. Pour in some ether. Shake. After solution place the tube in hot water. Have no fire near in working with ether. Why?

III. Test a little vinegar with blue litmus paper.

IV. Place in a test tube equal portions of alcohol and acetic acid. Add a little concentrated sulphuric acid. Boil and note odor. Add precipitated chalk until no further effervescence is noticed. This is done to neutralize the acid. Filter and taste. The esters constitute many of the soda-water flavors.

LESSON XXI.

CARBOHYDRATES.

135. Carbohydrates.—In the formula given for sugar, $C_6H_{12}O_6$, we notice that the hydrogen and oxygen stand in the same ratio to each other as they do in water—that is, H_2O, $6(H_2O)$. A large number of compounds, which, besides carbon, contain hydrogen and oxygen in that relationship, are classified as CARBOHYDRATES (hydrate meaning water). Some of these substances are sugar, starch, and cellulose.

136. Great Variety of Sugars.—There are a great many kinds of sugars. We do not mean by that brown, granulated, cut-loaf, or maple sugar, but that some sugars are solid and crystallizable; some liquid and non-crystalline, as honey; some are sweet, and many have no taste at all. It is hard for you to understand that there is a sugar which does not possess a sweet taste, but such exist. Further, anything that is sweet is not necessarily sugar. We cannot go into a discussion of the complicated structure of these sugars. Some have the formula given above, some have double that formula less one molecule of water, as ($C_{12}H_{22}O_{11}$), $C_5H_{10}O_5$, and so on.

137. Common Sugars.—We get a sugar from grape juice called GRAPE SUGAR. If sugar cane is compressed and the liquid obtained evaporated in large pans, CANE

SUGAR ($C_{12}H_{22}O_{11}$) crystallizes out. The thick molasses left is an impure solution of sugar in water. The treacle obtained from certain kinds of cane has a slightly bitter-sweet taste, due to small amounts of mineral salts present. This is commonly called SORGHUM. We may also obtain a sugar from sweet milk, entirely different from the others. It is called MILK SUGAR or lactose.

138. Acids from Sugars.—All of these sugars, under the proper conditions, may be made to undergo fermentation with the formation of corresponding acids. For instance, milk sugar can be converted into lactic acid; in fact, this is the change which occurs in the souring of milk. The acid produced coagulates the casein, a constituent of milk, causing it to thicken, forming "bonny clabber."

139. Starch.—STARCH contains, apparently, one less molecule of water than sugar, $C_6H_{10}O_5$. Starch is found everywhere in the vegetable kingdom in large quantity—particularly in all kinds of grain, as wheat, barley, etc.; in tubers, as the potato; in fruits, as chestnuts and acorns. In the United States starch is mainly manufactured from maize, and in Europe from potatoes.

140. Properties.—Starch has a granular structure, not crystalline. In the usual condition it is not soluble in water, but on being heated to 60°C with water, the starch cells burst and the contents go into partial solution. On cooling it forms a translucent mass, called STARCH PASTE. To prevent starch paste from souring or undergoing fermentation, a little salycilic acid should be added. When dried, as by a hot iron, it presents a certain inflexibility, and is used to stiffen collars, cuffs, etc., in the laundry. On being moistened it loses its rigidity.

141. Cellulose.—CELLULOSE forms, as it were, the groundwork of all vegetable tissues. The coarse fiber of the sturdy oak contains it no less than the tender shoots of the delicate heliotrope. It is, of course, more or less altered in form in these different substances. Cotton, hemp, and flax are almost pure cellulose.

142. Chemist's Power.—Cellulose is amorphous; insoluble in all ordinary solvents. It dissolves, however, in concentrated sulphuric acid, which, when diluted and boiled, produces dextrin and dextrose. The latter is one of the sugars. From an old shirt the chemist can make a sugar, then alcohol and ether or vinegar.

143. Explosives.—Concentrated nitric acid in the presence of strong sulphuric acid acts upon cellulose to form a nitrate. This so-called nitro-compound is the very explosive GUN-COTTON. It is soluble in a mixture of alcohol and ether, producing COLLODION-SOLUTION, which is used in medicine and photography. A mixture of gun-cotton and camphor is used under the name of CELLULOID. Celluloid is plastic at slightly elevated temperatures, and can then be moulded into any desirable shape, hardening as it cools. It takes fire very easily.

144. Paper, in its many forms, is for the most part cellulose, which may come from any source. Linen or cotton rags, wood or straw, are cut or torn into fragments. Caustic alkali is frequently used for cleansing and to aid in the disintegration. The pulp suspended in water is spread out in thin layers and passed between heated rollers to give it compact form and to dry it. The large sheets are cut into any desired shape. Paper is made white by bleaching, and the various tints are due to the different coloring matters used.

With the carbohydrates are included a number of substances which occur in nature called gums; for example, gum arabic or tragacanth.

EXPERIMENTS.

I. In one test tube make a solution of glucose in water. In another add a pinch of saccharine to water. Taste both. The former is a sugar, the latter is not, yet it is 284 times as sweet as ordinary sugar.

II. Boil some starch in water in a test tube. Note what takes place; add a small piece of iodine. Heat some ground wheat, flour, with water in a test tube, and test also with iodine.

III. Dissolve some cotton in cold concentrated sulphuric acid. Dilute quickly by pouring into a large amount of water.

IV. Pour a little collodion solution (obtained from drugstore) upon a glass plate. After the liquid has evaporated, set fire to the remaining film.

V. Apply a lighted match to a small piece of celluloid.

LESSON XXII.

ORGANIC BASES. OILS. GUMS.

145. Urea.—The inert black carbon can be made to unite with lazy nitrogen to form an active and very poisonous compound, gaseous cyanogen, C_2N_2. Hydrogen can also combine with these two to form another gas equally as poisonous. It is said that one-fifth of a breath of either of these gases is sufficient to produce death. The latter compound is hydrocyanic or prussic acid, (HCN). It forms salts called cyanides. HCN can further unite with oxygen forming cyanic acid, $HCNO$. This acid forms salts called cyanates—for example, ammonium cyanate, $(NH_4)CNO$. This body is one of the most important organic substances known, for it was this substance, which when heated allowed a re-arrangement of its atoms so that a body named UREA was produced. This body had never before been made except through the intervention of life. Urea, the result of the decomposition of nitrogenous substances in the body, served to break away the barriers previously existing between Organic and Inorganic Chemistry.

146. Organic Bases.—All the compounds containing carbon, hydrogen, oxygen and nitrogen are not acids; some are basic. Some of them are called ALKALOIDS, because they resemble the alkalies in their chemical action. They are usually found in plant and animal juices combined with acids to form salts. In opium, for example, there are seventeen of these alkaloids upon which the medicinal properties the drug possesses depend. In small doses either as the plain tincture (laudanum), or the flavored camphorated tincture (paregoric), opium is a sedative; in larger ones it is a narcotic poison. Morphine (from Morpheus, the god of sleep) is the chief narcotic principle of opium, and is also used to produce sleep and to alleviate pain.

147. Some Alkaloids.—Quinine obtained from Peruvian bark is used in fever and as a tonic. Nicotine forms from two to eight per cent. of the volatile principle in tobacco. One drop of it will kill a dog. Strychnine from the nux vomica is a deadly constituent of the poison used on arrows by the South American Indians. In minute amounts it is an excellent tonic; but in doses of any size tetanic spasms ending in death result. It is one of the most deadly of the alkaloids. It is said that five one-hundredths of a gram was known to kill a man. Strychnine possesses in the highest degree an almost unfailing property of all alkaloids, namely, bitterness of taste. One gram will impart a bitter taste to four thousand liters of water. In medicine these alkaloids are always used as salts, as sulphates or hydrochlorides.

148. Tea and Coffee.—There are two alkaloids which many of us take into our systems every day. Coffee and tea contain an alkaloid called CAFFEINE or THEINE, according to the source. Besides one per cent.

of caffeine, coffee contains about fourteen per cent. of oil and fat with an essential oil, which on roasting is volatilized, giving coffee the pleasant aroma. Tea contains twenty-odd per cent. of extractive matter with essential oils, and about fifteen per cent. of tannin.

149. Tanning.—Besides the stimulating effect of a decoction of tea, it possesses a very astringent taste due to the tannin. This tannin is an organic acid found also in the leaf and bark of the oak and hemlock, and in nut galls. It is soluble in water and coagulates albuminous substances, as gelatin. On this last property depends its value to us in our every day life. Skins of animals are treated with milk of lime and scraped to remove the hair and fatty substances and then exposed for variable lengths of time, the longer the better, in vats with water and ground hemlock and oak bark. The water dissolves out the tannic acid (as tannin), which passing into the pores of the skin combines with the albuminous substances to form an insoluble compound, the basis of our leather. Leather is blackened by washing it on one side with a solution of copperas, ferrous sulphate. Tannic acid combines with iron to form a dark iron tannate—in fact, an ink.

150. Essential Oils.—The essential oils we mentioned above differ from the ordinary oils in not being saponifiable with an alkali; they are not esters, but a large number of them are alike in composition possessing the same formula, $C_{10}H_{16}$. They are all of vegetable origin, being present in the petals of flowers, as in the violet; in seed, as caraway; in leaves, as mint; or root, as sassafras. Oil of turpentine, "spirits of turpentine," is an important member of this class of bodies. It is much used in making varnishes and in medicine. Two of its hydrogen atoms seem to unite with one of oxygen when exposed to the air, and a resin is formed. Thus we note that the turpentine around the mouth of a bottle in which it is kept becomes sticky, then resinous. This is in fact oxidation. Turpentine is an excellent solvent for many substances. In using it to remove grease spots fresh turpentine should be used, else a worse spot, more difficult to remove, will result from the resin present. The value of turpentine in paints is largely dependent upon this oxidation and formation of resin, hence painters call it a "dryer."

151. Resins.—If this resinous substance, called rosin, be dissolved in an essential oil, we have a balsam. Pitch, which exudes from incisions in certain trees, as the pine, is a balsam. By distillation the turpentine is separated from the rosin, which constitutes about three-fourths of the pitch. Amber is fossil resin. These resins do not decay, but are excellent preservatives. The ancient Egyptians embalmed bodies with these materials, and the mummies found in the Pyramids to-day are thus preserved after a lapse of over two thousand years.

152. Rubber.—India rubber or caoutchouc is a gum, which exudes from certain trees in South America.

Pure rubber is white, but the kind we usually see is darkened by smoke from the fires used to dry it, or by an admixture of some coloring matter. Vulcanized rubber is made by heating the rubber with a little sulphur. When strongly heated with a large percentage of sulphur, it becomes hard, brittle and capable of receiving a high polish, and is used for buttons, combs, knife handles, etc.

EXPERIMENTS.

I. To a crystal of strychnine the size of a pin head, in a watch glass, add a drop of dilute HCl, and dissolve in a liter of water. Taste.

II. To a little dilute solution of gelatine in water, add some water solution of tannic acid.

III. Mix together a solution of tannic acid or nut galls and copperas in water. Add gum to thicken, and make the fluid run smooth from the pen.

IV. Expose a tablespoonful of turpentine to the air for day.

V. Treat a cloth spotted with grease with fresh turpentine, wash this out with alcohol, and that in turn with water.

VI. Heat a small piece of hard rubber upon a silver coin. Sulphur turns silver black.

LESSON XXIII.

LIGHT METALS.

153. Historical.—To-day we must go back and pick up some seemingly lost threads. The ancients knew seven heavy lustrous bodies, which they called METALS. They were gold, silver, copper, iron, tin, lead, and mercury. Since that time a number of substances more or less like those mentioned have been discovered and classed with the metals

154. Characteristics.—By the term metal is ordinarily meant a substance that is heavy, opaque, but possessing lustre and capable of being hammered into thin sheets and drawn into fine wire. Now, there are several metals that actually float upon water, and some are too brittle to be beaten into plates, so we see that the classification possesses more convenience than accuracy. But metals are generally solid at the ordinary temperatures, and are good conductors of heat and electricity. Mercury is a liquid, however. Many of them liberate the hydrogen in acids when brought into contact with them, salts resulting. Combining with oxygen, as a rule, the metals form basic oxides, but in rare cases acid oxides are formed—arsenic oxide, for example.

155. Electrolysis.—We have learned that acidified water is decomposed when an electric current is passed through it. Some compounds of the metals are also decomposed by the electric current. As the metals always separate out at the negative side of the current, we say that they are electro-positive. This power of decomposition by electricity, called ELECTROLYSIS, was made use of in the early part of this century by Davy, who discovered the two light metals, sodium and potassium, by this means.

156. Light Metals.—Some metals, as sodium and potassium, have the power of decomposing water, liberating hydrogen, as we have seen. When thrown upon water the reaction produces a great deal of heat, so much that if the metals, which melt and roll about upon the water are confined, the gas takes fire, the metal being changed to a hydroxide. It is not necessary to confine the potassium, which takes fire anyhow. The bright metals, on exposure to the air, rapidly tarnish, the oxides being formed. This tarnishing is the same as the rusting of other metals. To avoid the "rusting," the metals must be kept covered with petroleum or kerosene oil. Care must be taken never to allow large pieces of sodium or potassium to come into contact with water. Neither of these metals should ever be handled with damp hands.

157. Caustics.—The sodium hydroxide made by the action of the metal upon water is commonly called "caustic soda." It is more usually made by heating slaked lime with sodium carbonate and plenty of water. Caustic soda is a white solid of a very disagreeable alkaline taste. It corrodes the skin, having a greasy feel when dissolved in water, and is classed among the corrosive poisons. It comes from the manufacturer in white lumps or sticks. It is largely used in making "hard" soaps, as is the "caustic potash," potassium hydroxide, in making "soft" soap.

158. Flame Reactions.—We have learned that sodium burns with a bright yellow flame. Potassium, which is very similar to sodium, gives a bluish violet-colored flame. If a mixture of their salts—the chlorides serve best—be burned, the color of the potassium is hidden by that of the sodium. When testing under such conditions, hold a dark blue glass between your eye and the flame. The blue glass cuts off the sodium light and you only see the potassium as a gray streak.

159. Occurrence.—On account of the strong attraction these metals possess for other substances, we do not find them occurring free in nature. As compounds, however, they are widespread. The presence of potassium in the soil is necessary for crops.

EXPERIMENTS.

I. Float a filter paper upon water, and drop a small piece of metallic sodium upon it.

II. Make a concentrated solution of sodium hydroxide in water. Dilute a portion largely; try the "feel" and taste. Place a piece of cloth in the strong solution and heat.

III. Expose to the air, on a dry surface, a piece of freshly cut sodium.

IV. In a colorless flame hold a platinum wire moistened with a solution of sodium chloride, then potassium chloride, then a mixture of the two. (See Fig. 12, p. 77.) Look at the flames through a blue glass. Slap a book in front of the flame.

LESSON XXIV.

MAGNESIUM. ALUMINUM. ZINC.

160. Flash Lights.—There are two other so-called light metals that deserve our attention—magnesium and aluminum. MAGNESIUM is a silver white metal, which retains its brightness in dry air, but tarnishes if moisture be present. We usually see magnesium in the form of ribbon, wire, or powder. When heated, it takes fire, burning with a brilliant light of high actinic power. The "flash light" used in taking photographs at nights is produced by the burning of magnesium powder. The magnesium light has been seen twenty-eight miles at sea. The white powder produced when magnesium burns is magnesia or magnesium oxide.

161. Aluminum.—How many of you have seen pictures of flying machines? Very few flying machines have been successful inventions because of the great weight of the material used in construction. The bluish-white, very light metal—aluminum—promises to help the flying enthusiasts not a little.

162. Source and Manufacture.—Aluminum in the form of compounds occurs everywhere. It is the main constituent of clay. The difficulty of obtaining it in metallic form has made it too expensive until rather recently. At one time it cost as much as forty dollars per kilogram; now it may be had for one dollar or less. The materials used as a source for aluminum are bauxite, corundum and emery. They are either hydroxides or oxides of aluminum. The purity is frequently dependent upon the locality from which they come. A very intense heat is necessary to cause carbon to rob these compounds of their oxygen. The difficulty of obtaining this high heat economically was the obstacle in the way of reducing the cost of this metal. Large electric furnaces are now used in the electrolysis of a solution of alumina in fused cryolite, a mineral containing sodium and flourine, besides aluminum. The carbon combines with the oxygen, forming carbon dioxide, the molten aluminum being drawn off from the bottom of the large carbon-lined crucible.

163. Properties.—Aluminum takes a high polish, and is very sonorous. It is half as heavy as iron, but not so strong. It is, however, tenacious, very malleable and ductile. It is little altered on exposure to the air. It is soluble in hydrochloric acid and the alkalies —sodium and potassium hydroxides—hydrogen being evolved. Because it is so easily attacked, it does not serve well for cooking utensils. However, it is not poisonous, even when taken into the system. Large amounts would derange the digestion.

164. Uses.—Aluminum has a wide range of usefulness. It is used in making optical instruments and delicate balances, where lightness of material having only moderate strength is necessary. Aluminum bronze, which contains ninety per cent. of copper and ten of aluminum, is very hard and malleable, having the tenacity of steel. It is gold colored, and can be highly polished. Accoutrements for soldiers, who must march great distances, are being made out of aluminum. It can best be worked when heated a little above the boiling point of water.

165. Zinc is a peculiar metal belonging to the same class as magnesium. It is grayish-white and crystalline in structure. At ordinary temperatures and at 200° C., zinc is very brittle, but at 140° C. it is quite malleable, and may be rolled into sheets.

166. Galvanized Iron.—Zinc decomposes most acids with an evolution of hydrogen. Commercial zinc is scarcely ever pure, but contains small amounts of lead, iron, carbon, and traces of arsenic and antimony. Zinc is not much affected by air, either dry or moist; therefore it is used to coat iron piping to prevent rust. Iron so coated is called "galvanized iron," and is made by placing the cleansed iron for a moment in a bath of molten zinc. Zinc salts are poisonous. Almost all natural waters contain some acid, hence galvanized iron should not be used to conduct such waters, if they are to be used for drinking purposes. Zinc readily precipitates most metals from solutions of their salts. In the case of lead, it produces the "lead tree."

EXPERIMENTS.

I. Ignite one end of a piece of magnesium wire held by tweezers.

II. Place a mixture of three parts magnesium powder and one of potassium chlorate in a piece of paper and set fire to the paper. Do this in a dark room and avert the eyes.

III. In three test tubes containing dilute hydrochloric acid, add separately small piece of magnesium, aluminum and zinc. Test the gases given off with a lighted taper.

IV. In a test tube containing sodium hydroxide solution, add several pieces of aluminum, and heat. Test the gas given off with a lighted taper.

V. Suspend a strip of zinc in a strong water solution of lead acetate (sugar of lead).

LESSON XXV.

MERCURY. TIN. LEAD.

167. Quicksilver.—Closely akin, chemically, to zinc is the only liquid metal we know, MERCURY, often called quicksilver. Mercury sometimes occurs alone in nature. It is a silvery lustrous, very heavy liquid. It also occurs combined with sulphur in the red mineral cinnabar. Many of the compounds of the metals with the non-metals or electro-negative elements have entirely different colors from any of the uniting substances. Tin is white; one of the sulphides is yellow, another brown. Red copper combines with chlorine to form a green compound, and the gray lead unites with oxygen to make yellow and red compounds.

168. Properties.—Mercury remains a liquid through a wide range of temperatures, boiling at 360°C, freezing at —40°C. It expands almost uniformly for equal additions of heat, therefore it is used extensively in making thermometers.

169. Amalgams.—Mercury dissolves many of the metals producing AMALGAMS. Mirrors are made by taking advantage of this solubility of metals in mercury. A thin sheet of tin is spread upon a perfectly level surface and mercury poured over it. A clean glass plate is then pressed down upon the mercury, the greatest care being exercised to prevent the presence of any air bubbles. Weights are added to press out the excess of mercury. The tin dissolved in mercury adheres closely to the glass, giving the bright reflecting background to our mirrors.

170. Crystallized Metal.—Speaking of TIN calls to mind the value of that beautiful white lustrous metal. Tin obtained in sticks creaks or "cries" when bent. This is due to the grinding of the crystals upon one another, for tin is highly crystalline. Sheet tin, used so much for roofing purposes, is, in fact, sheet-iron which has been covered over with a thin film of tin in a somewhat similar manner to that by which galvanized iron is made.

171. Sheet Tin.—Tin withstands the action of the weather even better than zinc, and as it does not corrode, serves excellently for roofing purposes. Sheets of tin alone would be too expensive. Iron is cheap, but would soon rust away and be of no value. In ordinary sheet tin, the strong cheap iron is made use of and the non-corroding covering of tin obtained as well. We must bear in mind that there must be no small cracks in the tin coating permitting the exposure of the iron to the weather, for if there were, the iron would corrode even faster, as galvanic action would set in. To avoid this corrosion, as invariably the tin is cracked some in the handling, tin roofing is painted with a mineral paint, which is not attacked. Thus the cracks are filled up and no corrosion can result.

172. A Soft Metal.—LEAD is a near relative of tin. Lead is soft, however, and corrodes in the air. Lead, which may be easily cut, at first exhibits a bright surface; it rapidly becomes dull, however. Lead melts very easily and can be cast into moulds of any shape. When warm, lead may be squeezed through round holes something like sausage coming from a sausage mill, except that an iron plug may be so arranged to make the tube hollow. This lead piping is made use of in plumbing on account of the ease with which it may be bent or unions made.

173. Lead, a Poison.—Lead salts are poisonous and cumulative in effect. Water containing dissolved oxygen, when passed through lead pipes, causes them to corrode, forming the hydroxide, which is soluble in water. This kind of water, which has passed through lead pipes, when drunk continually, will prove poisonous. If the waters be carbonated, the insoluble carbonate of lead is produced, and this acts as a protective coating, preventing further corrosion. As a rule, the amount of lead piping used is so small that there is practically no danger attending the use of the water for drinking purposes.

EXPERIMENTS.

I. Place several drops of a mercuric chloride solution (corrosive sublimate) upon a silver coin and rub it. What do you notice? Conclusion? Repeat, substituting zinc for the silver. Zinc poles in electric batteries are thus amalgamated.

II. Cut a piece of lead with a knife and notice the exposed surface at once and again the next day.

III. Dissolve a No. 8 bird shot in dilute nitric acid. Pour the solution into a liter of water and pass hydrogen sulphide into the solution. This is the method of detecting lead in drinking water, only the water is usually concentrated to one-tenth of the original volume by boiling.

LESSON XXVI.

COPPER. SILVER. GOLD. PLATINUM.

174. Colored Metals.—All the metals we have studied so far are white. There are a few that are colored; copper is red. "Silver threads among the gold" tells beautifully of the gray hairs of on-coming age.

175. Red Metal.—That metal, COPPER, is one of great use in the arts and manufactures. Hammered into sheets it is used to sheathe wooden ships; drawn into wire it serves to conduct electricity. Although the metal is red in color, many of its compounds are green or blue. The red color of copper serves as a danger signal, for its compounds are poisonous. The alloys of copper are very important, as we shall see.

176. Occurrence.—Copper occurs free in a conglomerate rock along the shores of Lake Superior. It also occurs elsewhere as an oxide and sulphide. In the latter state it is usually accompanied by some silver sulphide. The process for obtaining the pure metallic copper from its ores is quite complicated.

177. Ores.—Metals in mineral form may be very widespread in their occurrence. When these minerals are found in sufficient quantity to be profitably worked, they are termed "ores."

178. White Metal.—SILVER is a beautiful brilliant white metal, softer than copper, but harder than gold. It is exceedingly malleable and ductile, and the best known conductor of heat and electricity, except gold. Copper stands next to silver in this conducting power. Silver is unaltered on exposure to the air, and resists the action of hydrochloric and sulphuric acids. Like copper, it dissolves readily in nitric acid.

179. Uses.—Silver is much used for jewelry and plate because it retains it high polish. If pure silver were used for coins it would soon wear away on account of its softness. To give hardness, silver in coins is usually alloyed with ten per cent. of copper. Silver is not acted upon by the fused hydroxides of the alkali metals. On exposure to sulphuretted hydrogen it turns black. This black sulphide of silver goes under the incorrect name of "oxidized silver."

180. Noble Metals.—Now there are some metals that are not acted upon by the acids when used singly. The so-called "noble metals," GOLD and PLATINUM being taken as examples, are soluble, however, in a mixture of concentrated hydrochloric and nitric acids. This mixture is called "aqua regia," because it dissolves these royal metals. These metals are seldom found in nature combined with other elements. Although they come into contact with oxygen and sulphur through long ages, as all lie together buried in the earth, they seem to disdain any intimate relationship. This is more especially so of platinum, as gold occurs in places as sulphide and telluride.

181. Malleability and Ductility.—GOLD is a beautiful yellow metal. When in the finely divided state it appears ruby red. Much of the red colored ruby glass is due to gold in this state of subdivision. Gold is the most malleable and ductile of all the metals. It may be hammered into such thin sheets, that it requires one hundred thousand of them to make one centimeter. This gold leaf transmits green light. One gram of gold may be drawn into four kilometers of wire.

182. Measure of Fineness.—Gold is very soft, therefore it is alloyed with copper to harden it to prevent its wear in use. It is extensively used for coins, jewelry, gilding and other purposes. The purity of gold in jewelry is estimated by "carats." Pure gold is "twenty-four carats fine." Ordinary watches and rings are of fourteen or eighteen carat fine gold.

183. Fusibility.—PLATINUM is a heavy, soft metal of tin white color. Gold melts at about 1100°C, and platinum at almost 2000°C. Platinum is malleable and ductile and is not acted upon by most chemicals, hence it is hammered into different shapes for use in chemical laboratories. Such metals as lead when heated in platinum ware form a fusible alloy, producing a hole in the vessel. Platinum may be welded like iron at red heat. It possesses the property of absorbing large amounts of gases after being heated. This absorption of gas by a metal is called "occlusion." Under the influence of heat and cold platinum expands and contracts almost equally with glass, therefore it is used to conduct electricity through glass in incandescent electric lights.

EXPERIMENTS.

I. Dissolve (out of doors) a few copper scraps in concentrated nitric acid. Add to a portion of the solution ammonium hydroxide in excess. Note the color of both solutions.

II. (a) Stick a bright nail in a solution of copper sulphate. (b) Add a few drops of sulphuric acid to the solution, and insert narrow strips of platinum and zinc, allowing the outer ends to touch. Electroplating.

III. Expose a dime to the fumes of hydrogen sulphide.

IV. Try to dissolve a small piece of platinum wire in HCl, HNO_3 and H_2SO_4 separately. Finally try a mixture of two parts of HCl and one of HNO_3, both concentrated. Save the solution.

V. Heat a small piece of lead upon a platinum wire.

LESSON XXVII.

ALLOYS.

184. Alloys.—In some of our previous lessons, we used the term ALLOY, without defining it. An alloy is a mixture or compound of several metals brought about by melting them together. Alloys assume new and different properties from those of the constituents, depending upon the proportions of the metals used. The alloys are frequently more useful than the pure metals themselves. Gold and silver are so soft that articles made of them would soon wear away in the necessary handling in commercial transactions with a consequent depreciation in value. These metals alloyed with ten per cent. of copper have been found to be sufficiently hard to withstand much loss in handling.

185. Brass.—There are other alloys of copper containing that metal in large amount. Doubtless, so far as usefulness is concerned, copper is more valuable than gold or silver, for it is the main constituent of brass, bronze and bell metal. Brass is composed of copper, two or three parts, and zinc one part. This alloy is extensively used for ornamental purposes, etc. The greater the proportion of copper, the more malleable and ductile is the alloy. One per cent. of lead is ordinarily added during the melting to make the alloy soft enough to be cut on a turning lathe and to prevent the metal from sticking in the teeth of files used in shaping the material. When a larger amount of lead is used the alloy assumes the same color as gold, and is used as spurious gold. It soon tarnishes on exposure, however.

186. Bell Metal.—Bronze is an alloy of copper and tin. In one proportion for awhile it served to make cannon, and in another it still serves to make our church bells. An alloy to be used in making ordnance, must be tough to prevent explosion of the gun, elastic to yield slightly to the large volume of gases generated, and hard to avoid being worn away by the projectiles used. Usually, a little phosphorus is added before casting to make these properties more pronounced. Bell metal must, of course, be sonorous, and it is always hard and brittle. Bronze is easily cast—that is, when melted, it will assume the shape of any mould into which it may be poured. For this reason many works of art are made of bronze. They do not suffer on exposure to the weather, but take on a beautiful rich rust, called *patina*. The peculiar lustre of the metal may be seen through this green coat.

187. Casting.—Alloys containing zinc, on cooling from the molten state tend to shrink, therefore brass is not a good substance for casting shapes with fine lines and corners. If, however, it be desired to cast around something—for instance, an iron ball—brass is an excellent material for that purpose. There is another metal, however, used in alloys which does not shrink, but even causes expansion upon cooling. It is antimony.

188. Type-Metal.—In ancient times, knowledge was preserved by verbal tradition from one generation to another. This method was in turn replaced, after the art of writing was introduced, by handwritings upon parchment. About the fourteenth century, wooden blocks with letters cut on them, were used for printing purposes. These, however, were too perishable. And in the century following, metal was for the first time used to make types. Type-metal should be strong, able to withstand pressure, easily melted, so that it can be recast, as constant use wears it away. On comparing a capital "I" and a small "l," or an "e" and "c," we see how necessary it is to have a metal which, in casting, will flow out, or on cooling, expand into the smallest depression in the mould that the true character of the type may be represented. Four parts of lead mixed with one part of antimony give us this type-metal, which is usually cast under pressure.

189. Solder.—Alloys frequently melt at lower temperatures than the metals of which they are composed. Tin melts at about 230° C., and lead at 330° C., yet a mixture of equal parts of these metals melt at 186° C. This is the "soft" solder used by tinners. "Hard" solder contains copper, zinc and tin. There are many kinds of solder composed of different metals according to the kind of union to be made. The surfaces of the metals to be joined together by soldering must be bright and free from oxide, else the union will not be close. The air must be excluded during the operation also, to prevent the formation of oxide under the influence of heat. This is usually done by rubbing powdered borax or rosin over the clean surface before applying the heated soldering iron These substances also dissolve the metallic oxides formed during the heating, and thus help to keep the surfaces to be soldered untarnished. A good "non-rusting" soldering fluid is made by dissolving zinc in hydrochloric acid and adding ammonium hydroxide until the excess of acid is neutralized. (How may we recognize this neutral point?)

Instead of having experiments with this lesson, a printing office, preferably where electroplating is done, and a tin shop, could be visited with profit.

LESSON XXVIII.

IRON.

190. Base Metal.—Of what value would a king be, if he had no subjects? It may be asked of what value would the noble metals be, if there were no base ones? A nation is not rated according to the king, or president, but according to the people and their industrial value. We may consider just one more metal, the value of which is almost incalculable; without which there would be no life. This common metal, IRON, seems to have been placed on earth not only as a means for man's progress, but as a necessity for his very existence. Very likely the doctor has prescribed an "iron tonic" for most of us.

191. Occurrence. Impurities.—The occurrence of iron combined with other elements is universal, not only in the mineral world, but as well in all the realms of life, animal and vegetable. The iron ores, magnetite (sometimes called lodestone), brown and red hematites and so forth, occur in vast quantities all the world over. The great and serious difficulty is to obtain the metal comparatively free from impurities. Ores of all kinds always occur surrounded by and mixed with other rock. The main constituent of rocks is silica, an oxide of silicon. Sulphur and phosphorus sometimes occur in iron ores to the extent of three per cent. Phosphorus has the peculiar effect upon iron of making it brittle when cold, "cold short." Sulphur causes the iron to crumble when hammered or rolled hot, "hot short." It is easily seen that the presence of these substances is very deleterious to the usefulness of the metal, yet perfectly pure iron is a substance rarely seen.

192. Manufacture.—To get iron in a metallic state the ore in lumps from the size of one's fist to the head is mixed with coke and limestone. The carbon of the coke combines with the oxygen of the ore forming carbon dioxide and leaving molten metallic iron. If limestone, calcium carbonate, be heated intensely with silica, they will combine, forming a glass, which is liquid at the high temperatures necessary to deoxide the iron ore. This glass is called "slag," and dissolves out a great deal of the phosphorus and sulphur, and prevents the reduced iron from being reoxidized by the air pumped into the furnace near the bottom. This process of removing impurities is called "fluxing."

193. Iron Furnace.—Very large quantities of iron are worked with; hence huge furnaces, twenty-five meters high and six meters in diameter, are used. The ore, coke, and limestone are dumped into the top of the furnace in cart loads at a time. A high temperature is necessary for the chemical changes to take place, and, as the quantities used are enormous, to get that temperature immense blowing engines are kept pumping air into the sides near the bottom of the furnace, like the bellows at a blacksmith's forge. The temperature reaches as high as 1500°C, or higher, in the hottest part of the furnace. The molten iron, being heavier than the slag, sinks to the bottom of the furnace and is drawn off and run out into trenches in sand. From time to time the slag, which occupies considerable space, is drawn off from holes higher up in the furnace. It is cooled by water being thrown upon it, which makes it porous. This is sometimes used for railroad ballast. If the amount of phosphorus is considerable, as it is in basic steel slag, it is used as a fertilizer.

194. Cast Iron.—The iron, drawn from the bottom of the furnace, after cooling, may be either white or gray, and when broken presents a crystalline appearance. This is ordinary "cast iron," and contains from 2 to 6 per cent. of carbon. The carbon has come from the coke used to reduce the iron oxides to the metallic state. In the WHITE CAST IRON, the carbon is combined with the metal, causing it to fuse more easily than the gray. In dissolving white cast iron in an acid some of the hydrogen liberated combines with the carbon, forming disagreeably smelling hydrocarbons, easily detected in the gases given off. The GRAY CAST IRON has some of its carbon combined, but most of it is in the form of minute crystals of graphite. The more combined carbon there is present, the easier is the iron fused and worked.

195. Steel and Wrought Iron.—This crude, impure form of iron requires further special treatment, according to the use to which it is to be put; for instance, by reducing the amount of carbon to from 0.2 to 1.5 per cent. we obtain STEEL. When the percentage of carbon becomes even lower than 0.2 of one per cent. we have WROUGHT iron. Wrought iron is the least fusible of all the varieties. The blacksmith buries his bar of iron to be forged in a bed of hot coke, not only to heat it so that he may hammer it into shape, but also that it may absorb carbon to replace that burned out by previous work. The absorption of this carbon makes the iron melt more easily.

196. Tempering.—All iron in bars is crystalline. Steel is only one of the varieties of iron possessing peculiar properties of its own. For instance, if we allow it to cool slowly after heating, it becomes quite tough, but comparatively soft. It may easily be dented with a hammer. If it be cooled very quickly, as by plunging into water when white or red hot, it becomes exceedingly hard, but brittle. This is called "TEMPERING."

197. Importance.—When we stop to think that it takes several thousand dollars to put an iron furnace into operation, or "blowing in," as it is called, in addition to the cost of construction, and then recall the immense number of such furnaces in the world, and the innumerable uses made of iron, from a small watch spring to huge engines, we stand in awe at the stupendous amount of capital necessary to manipulate this simple base metal.

EXPERIMENTS.

I. Dissolve a little iron in a mixture of HCl and HNO_3. Evaporate to dryness in a porcelain dish; add a little more HNO_3, and dry again, finally ignite. One of the oxides of iron is produced.

II. Place some blood in a porcelain dish, add several drops of HNO_3 and evaporate to dryness, and heat until no more fumes come off. Old blood stains upon cloth are usually due to the iron present.

LESSON XXIX.

METALLIC OXIDES.

198. Oxides of Metals.—All metallic elements form compounds with oxygen, hence there is a large number of oxides. Many of them exist free in nature, but many require the most careful laboratory manipulation for their preparation.

199. Colors of Oxides.—Most of the iron ores are oxides. They differ in color. In fact, the metallic oxides are of almost all colors, from white to black. Mercuric oxide, from which Priestley obtained oxygen, is red. One of the oxides of lead is yellow, another red, and still another is brown. Manganese, too, has several oxides, the most prominent one being the black or dioxide, frequently called binoxide.

200. Lime.—If marble—calcium carbonate, also called limestone—be heated intensely, it gives off carbon dioxide. The oxide of calcium, or our ordinary "quick lime," remains. When quick lime is placed in water it "slacks." Sufficient heat is sometimes produced to cause the water to boil. This heat is the result of the chemical action brought about when the water unites with the calcium oxide to produce the hydroxide or "slaked lime,"

$$CaO + H_2O = Ca(OH)_2.$$

If quick lime be exposed to the air for a long time, it absorbs moisture and carbon dioxide, forming the hydrate and carbonate:

$$CaO + CO_2 = CaCO_3.$$

In such a state, it no longer has value as a constituent of mortar.

201. Paints.—The red oxide of iron, amorphous ferric oxide, Fe_2O_3, mixed with oil, is much used as a paint for metallic surfaces. This "mineral paint" is not affected by moisture or oxygen, hence forms an excellent protective coating against rust. By paints, we ordinarily mean substances that are, as a rule, insoluble in water, and are mixed with either a weak glue solution (being then termed water-colors), or with linseed oil (called oil-paints). These insoluble substances may be colored or white with the addition of sufficient coloring matter to give the desired tint. Many of the coloring matters are soluble in water.

202. White Paints.—The basis of most white paints is a compound of lead, called white lead. Sulphuretted hydrogen forms with lead a black compound, lead sulphide. Therefore, if this gas be in the atmosphere, and it is desired that the paint retain its original brightness, something else than a lead compound must be used as the body of the paint. Zinc oxide, called "zinc white," is much used as a substitute, the sulphide of zinc being white. It is much more expensive than white lead, and possesses less "covering" power—that is, more of the substance is required to cover the surface—but it gives a peculiar beautiful white gloss. White lead is sometimes adulterated (weighted) with the heavy barium sulphate.

203. Gems.—The light metal, aluminum, is the most abundant element upon the earth's surface, except oxygen and silicon, and it has but one oxide. This oxide, alumina, Al_2O_3, is white when pure, but it occurs in nature as the gray corundum, and crystallized as the ruby and sapphire, which have their characteristic brilliant colors. Emery is an impure corundum. All these minerals are very hard, ranking next to the diamond, which is the hardest of all minerals. On account of their color and scarcity, rubies and sapphires are very valuable, being classed as "gems." Emery and corundum, on account of their hardness, are used for grinding and polishing hard surfaces.

204. Oxides for Illumination.—Some of the oxides not only rank amongst the hardest known materials, but are the least affected by heat—only the most intense heat fusing them. To utilize this infusibility, the Auer-Welsbach gas-burner was invented. A bag of cotton gauze is dipped into a solution of thorium and zirconium nitrates and allowed to dry. When strongly heated, the cotton cloth is burned away and the nitrates decomposed, and we have a white gauze made of the oxides of the metals which had been in solution. This infusible gauze, when ignited sufficiently, gives a pleasant, soft, white light.

205. Calcium Light.—What is known as the "calcium light" is produced by directing the intensely hot oxyhydrogen blowpipe flame upon quick lime. The light is blinding in its brilliancy, and the lime is heated to over 2000° C. without melting.

EXPERIMENTS.

I. Drop a piece of quick lime the size of a walnut into half a liter of water. After all action has ceased, filter the solution into a porcelain dish and pass in carbon dioxide until the precipitate produced is partly redissolved. Boil off the excess of carbon dioxide and filter again; dry the solid, which is "precipitated chalk." This is not our ordinary crayon chalk, which is calcium sulphate; nor tailor's or French chalk, which is talc.

II. Pass H_2S over some white lead paint; zinc white paint.

III. Place several lumps of marble in a hot bed of coals or in an open fire, and allow them to stay there for two hours. Remove with tongs, allow to cool, and place in a wide-mouth bottle with a cork. What is thus prepared? Test its action on moistened blue litmus paper.

IV. Direct the flame of a blowpipe cautiously, but continuously, upon a small pencil cut out of quick lime.

LESSON XXX.

CHLORIDES.

206. Salts.—One of the simplest salts is common "salt." This condiment, actual food, is sodium chloride. It must not be understood from this that all salts possess a salty taste.

207. Salt Beds.—Sodium chloride occurs in small amounts all the world over. In some places, however, as a mineral, it occurs in large beds, usually in sedimentary geologic formations. Such beds have been discovered in Germany, England, at Petit Anse, Louisiana, Saltville, Va., Syracuse, N. Y., etc. It also occurs in salt springs. These saline springs are the result of the solution of this crystalline mineral in underground water. In some places the water from such springs has accumulated, and as the liquid has been removed by evaporation, the solid has remained behind. Thus were our salt lakes in Utah produced. The salt lakes are so strong in sodium chloride that it crystallizes upon the bodies of persons bathing in them. No fish exist in these lakes. Again, these saline springs have poured their waters into numerous flowing streams, and the salt has collected in the great oceanic reservoirs. The dry air, as a sponge, has soaked up the water vapor from the oceans. The rivers have continued to carry salt to the sea, in infinitely small amounts at a time to be sure, so that the amount of salt has gradually increased until it now reaches from two to four and one-half per cent. The solid constituents of the sea water are practically the same in all parts of the world, sodium chloride predominating.

208. Salt Gardens.—Salt for a long time was obtained from the sea by an arrangement of "salines" or salt gardens, as it were. On a level shore in a hot climate a large reservoir was constructed. The water from the sea was allowed to run in through connecting canals. The heat of the sun evaporated most of the water, leaving a strong brine into which more sea water was run and allowed to evaporate. In this way a concentrated solution of the salts was secured. The sodium chloride being the least soluble of the salts present, separated out in crystals and was raked up on the ridges between the reservoirs. These crystals were not pure, but contained admixtures of magnesium and potassium chlorides. These latter absorbed water on exposure to moist air, and becoming liquid, sank into the ground underneath. The piles of salt were covered with straw to prevent washing away by rain.

209. Working Salt Springs.—Salt is frequently got by evaporation of the water from salt springs. The more extended the surface exposure the more rapid is the evaporation of the water which holds the salt in solution. To obtain this large surface, racks about ten meters high are built, and the space intervening between the top and the ground is filled in with branches of trees. The spring water trickles down from the top, and the salt appears in sparkling crystals on each twig and is easily collected in baskets. Where salt occurs in large beds, it is mined as if it were an ore, or by dissolving the salt in water, which leaves the insoluble impurities, and evaporating the brine thus obtained.

210. Impurities.—Sodium chloride crystallizes in cubes, sometimes separating in a hat-like form. The varieties seen in trade depend upon the size of these crystals, differing according to the rapidity of the crystallization. Fine table salt is crushed to the powdery state in which we usually see it. Pure salt is not deliquescent, but most salt contains some little potassium chloride, which does absorb water, causing the salt to cake, as is frequently noted in salt cellars used upon the table. Usually salt contains from three to five per cent. of water, not as a constituent, but so bound up in the crystals as not to give evidence of its presence unless the salt be heated. On the application of heat the crystals burst with a crackling sound, and are thrown about with some violence. This is called DECREPITATION. Decrepitation is due solely to the escape of water changed into vapor or steam by the heat.

211. Uses.—Sodium chloride has a number of uses, being a necessary condiment in food. It is essential for the well being not only of man, but beast. A man weighing seventy-five kilograms contains in his body one-half kilo of sodium chloride. In order to maintain this amount, as so much is got rid of daily, he must eat as much as seven and a half kilos of salt every year. All of us know the necessity for "salting" cattle. A great deal of salt is used in making soda, chlorine, in tanning and in glazing certain rough brown kinds of earthenware. It is also much used as a preservative of wood, as well as meat, butter, etc. In fact, England produces over two million and the United States about one million tons per annum.

212. Horn Silver.—Silver chloride, a white curdy substance, is of great interest, as it is so much used in photography. It is very sensitive to light, changing under its influence into a body, which is not soluble in the various mixtures called fixers, which do dissolve the unchanged chloride.

EXPERIMENTS.

I. Dissolve as much sodium chloride as possible in 200 cc. of boiling water; evaporate off three-fourths of the water by heating. Cover, cool slowly, and note the form of the crystals.

II. Direct the flame of a blowpipe upon some dry salt on a piece of charcoal.

III. To a solution of silver nitrate in a test tube add several drops of HCl. Allow to stand in sunlight, and note the change on such exposure.

IV. PHOTOGRAPHY.—Float a piece of sized paper upon a solution of sodium chloride. Dry and then float it for three minutes on a silver nitrate solution in a dark room. Press a fern leaf upon a glass and place the dried paper upon the leaf, and cover that with a board the size of the glass. Clamp all together with clothes pins, and expose the glass side to the direct sunlight. When sufficiently dark, remove the clamps and place the sheet of paper in a strong solution of sodium hyposulphite (ten per cent.), and wash thoroughly with water. The "hypo" solution dissolves the unchanged chloride, which has been protected from the light by the leaf.

V. DELIQUESCENCE.—Expose some dry sodium chloride, potassium chloride and calcium chloride for a day or two. Which are deliquescent substances?

LESSON XXXI.

NITRATES. EXPLOSIVES.

213. Saltpetre.—The salts of nitric acid are called nitrates. Potassium nitrate does not occur in large quantities in any one place except in India, and then in only fair amounts. In small amounts, it is present in almost every soil. As enormous quantities of this so-called "nitre" or "saltpetre" are used every year, some artificial mode of production has had to be resorted to. Nature has been imitated as closely as possible. It has been known for a long time that where organic matter containing nitrogen was decomposing in the presence of potassium carbonate, the nitrate always resulted.

214. Nitre Plantations. Purification by Crystallization.—Wood ashes contain a high percentage of potassium carbonate. "Nitre plantations" are made by mixing these ashes and lime with manure and all kinds of refuse animal matter piled in layers, made porous by straw. These heaps are kept moist and exposed to the action of the oxygen of the air for two or three years. The nitrogenous matter is decomposed to nitric acid, which, having combined with the lime, is in turn decomposed by the potassium carbonate. The potassium nitrate thus obtained is secured by leaching the beds with water. As all nitrates are soluble in water, they are thus collected in solution, which, on evaporation to a small bulk and cooling, allows the nitre to separate out in large crystals. One hundred parts of boiling water will dissolve two hundred and forty-four of nitre, whereas only thirteen parts are dissolved if the water be at zero. Taking advantage of this, we can purify the nitre by repeated crystallizations, a small portion of the salt being lost at each crystallization; the impurities remaining in solution are washed away.

215. Chili Saltpetre.—Sodium nitrate is called "Chili saltpetre," because enormous beds of it are found in that country. Sodium nitrate is much used as a fertilizer, as we shall see; it is also used for the manufacture of potassium nitrate, which has more commercial value.

216. Gunpowder.—Nitric acid and the nitrates have a large amount of oxygen present in them, therefore one would imagine that substances mixed with them should burn more vigorously than otherwise. This fact is made use of in the manufacture of gunpowder. If we heat a piece of potassium nitrate upon charcoal, we note a very rapid burning of the carbon. The principal upon which gunpowder is used, is based upon the fact that a large volume of gas confined to a very small space exerts intense pressure or force. The more gas generated for the amount of material, the greater the explosive force.

217. Products of Explosion.—The average gunpowder consists of a mixture of seventy-five per cent. potassium nitrate, twelve per cent. sulphur, and thirteen per cent. carbon. This seems to be the best proportion of the three materials to produce the most rapid burning in order to convert the whole substance as quickly as possible into gaseous products. In this case, carbon dioxide and nitrogen are the principal gaseous products, and a small quantity of potassium sulphide is the main solid product formed. As we learn from our physics that the higher the temperature the greater is the volume of the gas, therefore it is desirable to have the gas heated as much as practicable. We must "keep the powder dry," else it will not burn, or at least burn slowly, thereby decreasing its effectiveness. Sodium nitrate attracts moisture, hence its use in gunpowder is undesirable; it does very well in blasting powder, where cheapness is a great factor. Potassium nitrate is universally used for gunpowder. The noise of the explosion is caused by the confinement and sudden release of the gases, driving the air away, which, on its return, produces the report.

218. History.—The amount of powder used every year is enormous. One of the largest cannon in the forts or on board of one of the tremendous war ships, requires nearly five hundred kilograms for each discharge. The early history of gunpowder is very obscure. Its first use in war was in the fourteenth century.

219. Nitroglycerine.—There are a great many explosives, all dependent upon the principles stated above. They are also dependent upon the use of nitric acid in their manufacture. If we treat glycerol (glycerine) with a cold mixture of concentrated sulphuric and nitric acids, we get an oily substance incorrectly called nitro-glycerine—better, glyceryl nitrate. It is extremely explosive, and when absorbed in saw dust or kieselguhr, chalky earth, it is sold as dynamite.

220. Lunar Caustic.—The white crystalline silver nitrate, called "lunar caustic," has use in medicine as a cautery to destroy diseased flesh.

EXPERIMENTS.

I. Direct a blowpipe flame upon a piece of potassium nitrate on a piece of charcoal. Deflagration.

II. Place a small amount of black powder on a piece of tin and set fire to it by means of a match attached to a long stick. Repeat, using moistened powder.

III. Make a mixture of one part shellac, four parts strontium nitrate, and one-fifth part potassium chlorate. Make another, substituting barium nitrate for strontium. Set fire to small amounts of each. The colors can best be seen at night.

LESSON XXXII.

CARBONATES.

221. Acid Salts.—We learned that when hydrochloric acid was passed into a solution of sodium hydroxide the one hydrogen atom in the acid was replaced by sodium. If carbon dioxide be passed into water the hypothetical carbonic acid, H_2CO_3, is produced. We do not really know the acid, but we do know, that if we pass carbon dioxide into a solution of sodium hydroxide we can obtain two salts. In one of these the sodium takes the place of only one hydrogen atom and in the other both are replaced. This carbonic acid is, therefore, like sulphuric acid, in that it can form an acid and a neutral salt.

222. Cooking Soda.—If carbon dioxide be passed through a strong solution, ten per cent. of sodium hydroxide, we obtain at first the neutral salt, Na_2CO_3. If we continue to pass the carbon dioxide through the solution, very soon we notice white crystals separating out. This is the acid carbonate, according to the equation,

$$Na_2CO_3 + H_2CO_3 = 2NaHCO_3.$$

This is our common cooking soda, and is used in making Seidlitz powders. Cooking soda, an ingredient of the many artificial yeasts or baking powders, is used with sour (acid) milk to make bread " light."

223. Chemical Action Promoted by Solution.—The acid carbonate contains a large excess of carbon dioxide, which is liberated when an acid is brought into contact with it in the presence of water. The powder in the white paper of Seidlitz powders is acid potassium tartrate, and the blue paper has acid sodium carbonate mixed with Rochelle's salts, according to the dose. If these two powders be mixed in the dry state, we have no evidence of chemical action at all; but if we add a little water effervescence is noticed at once. As chemical action is between molecules and atoms in the molecules, to promote chemical action they must be brought as close together as possible. When a substance is brought into solution it is resolved at least into the molecular state, hence the molecules come closer together and chemical action takes place more readily. In the case mentioned the gas, carbon dioxide, is produced with the solid, sodium potassium tartrate, a double salt, which remains in solution. The gas has an exhilarating effect upon the system, and the tartrate possesses its own peculiar medicinal value.

224. Soda.—The manufacture of the neutral or saturated sodium carbonate constitutes one of the most important branches of chemical industry. Immense quantities, thousands of tons, are used every year in the manufacture of glass and soap, as well as in making the numerous sodium compounds. It is commonly called "washing soda". As sodium chloride is abundant and cheap, it serves as the starting point in the production of sodium carbonate. There are two great rival processes which have contended for the supremacy in the production of this material.

225. Leblanc Process.—The Leblanc process depends upon the formation of sodium sulphate by the action of sulphuric acid on sodium chloride. The sulphate is mixed with two parts of carbon and one of limestone, and the mixture heated in a special furnace. The sodium carbonate produced is leached out with water from the impure material, and the solution concentrated by boiling. The carbonate combined with ten molecules of water separates out on cooling in large crystals. This exhibits further the great value the process of purification by crystallization possesses in aggregating like molecules and leaving the impurities in solution.

226. Solvay Process.—The other method, known as the Solvay or ammonia-soda process, also makes use of a strong brine. The salt solution after being saturated with ammonia gas is placed in iron cylinders in which carbon dioxide is forced with a pump. The cylinders are kept cool by a water jacket on the outside. Sodium hydrogen carbonate and ammonium chloride are the result of the chemical action which takes place. The former of these two compounds is rather insoluble, while the latter is very soluble and may be easily washed away. On heating the acid carbonate changes to the neutral carbonate, CO_2 and H_2O being driven off

$$2NaHCO_3 = Na_2CO_3 + H_2O + CO_2.$$

A new electrolytic process, the Castner-Kellner, promises to displace both of these.

227. Other Carbonates.—Potassium carbonate is made by evaporating the lye obtained from wood ashes. As the evaporating was formerly done in pots, it assumed the name of "potash." Ammonium carbonate is the ordinary "smelling salts."

228. Stalactites.—Calcium carbonate is insoluble in water, but easily soluble in acids, even carbon dioxide and water (*i. e.*, carbonic acid). Calcium carbonate is held in solution by waters carrying an extra amount of carbon dioxide. When that gas is evolved, the carbonate separates out usually in a crystalline state. This is seen in many caves where the stalactite and stalagmite formations present such beautiful outlines. Temporary hardness of water is due to the presence of calcium carbonate held in solution by the carbon dioxide present. As soon as the carbon dioxide is got rid of the carbonate separates out. This is accomplished by boiling, the heat liberating the gas. The shells of snails, mollusks, etc., are composed mainly of calcium carbonate.

229. Basic Salt.—As an example of a basic carbonate, we may mention white lead. Lead hydroxide is somewhat soluble in water, as noted, but it has a good covering power. Crystalline lead carbonate is transparent, but is insoluble in water. White lead is a compound of these two, in which the good qualities of each are utilized—that is, white lead is a basic carbonate, used as the ground work for most paints.

EXPERIMENTS.

I. Pass a steady stream of CO_2 into a ten per cent. solution of sodium hydroxide until a white precipitate begins to form.

II. Mix intimately equal parts of powdered dry cooking soda and acid potassium tartrate (*i. e.*, cream of tartar). Add water to the mixture.

III. Leach some wood ashes with water and evaporate the solution to dryness. To a portion add HCl. Test this solution with a platinum wire held in a colorless flame. Allow the other portion to stand until the next recitation. "Pearlash," *i. e.*, potassium carbonate, absorbs water.

IV. To some powdered snail or oyster shells add HCl. Test the gas given off.

LESSON XXXIII.

SULPHATES. PHOSPHATES.

230. Acid and Neutral Salts.—With sulphuric acid, we may have the acid and neutral salts. Sodium replaces one or both of the hydrogen atoms, giving, in the latter case, "Glauber's salts" (Na_2SO_4), named after the chemist who discovered the substance. Much of this salt is produced in the Leblanc soda process. It possesses valuable use as a medicine.

231. Epsom Salts.—Now, just as we have a di-basic acid, so may we have a di-acid base or metal which replaces two hydrogen atoms instead of one. Magnesium is such a metal, as is calcium also. Magnesium sulphate ($MgSO_4$), "Epsom salts," so named on account of its abundant occurrence in a spring by that name, is much used in medicine.

232. Gypsum.—Calcium sulphate occurs in large quantities as the crystalline mineral gypsum. A mountain range in Switzerland is made up almost entirely of it. The other sulphates we have just mentioned are very soluble in water, whereas calcium sulphate requires five hundred parts of water for solution. As a rule, substances soluble in water are more soluble when the water is hot than when it is cold. In the case of calcium salts, the reverse of this is true. Calcium and magnesium sulphates, or chlorides, when dissolved in water, produce permanent hardness, as that property of such water is not lost by heating.

233. Plaster of Paris.—Calcium sulphate, when mixed with water, forms a paste, which, after a portion of the water has evaporated, "sets" or hardens. It is called "plaster of Paris," and is calcium sulphate combined with two molecules of water. By gentle heat this water may be driven off and the sulphate will harden again when mixed with water; for, in fact, this so-called plaster of Paris is the anhydrous sulphate obtained by heating gypsum ($CaSO_4 + 2H_2O$), which drives out the water of crystallization. If it be too strongly heated, however, it loses the power of forming a compound with water. Plaster of Paris is largely used in making casts and in surgery. One anhydrous form of calcium sulphate, called alabaster, is almost transparent, and is often used for carving beautiful statuary.

234. Colored Sulphates.—Copper sulphate is "blue stone," so much used in telegraph batteries. Ferrous, or iron sulphate, is a green colored substance, commonly called "copperas," although there is not a trace of copper in it. It is used as a deodorant.

235. Tri-Acid Bases.—Aluminum is a metal that can replace three hydrogen atoms; and, as there are only two in sulphuric acid, we see that one atom will require one and a half molecules of acid. Since we cannot have fractions of a molecule existing alone, we must have two atoms of aluminum and three molecules of sulphuric acid to produce the salt aluminum sulphate, $Al_2(SO_4)_3$.

236. Alums.—Aluminum sulphate can combine with the sulphates of the alkalies—potassium, sodium, ammonium, etc.—to form double sulphates, an important class of bodies called ALUMS. The aluminum can be substituted by iron or chromium in such compounds, the iron or chrome alums resulting. The common alum (aluminum potassium sulphate), with about an equal amount of water, is much used in medicine and adulterating baking powders. These beautifully crystalline compounds are much used in dyeing—not so much as dyes themselves, but as fixers of the colors. Many colors do not affect the cloth to be dyed, and require to be fixed or bound to it. Such dyes are "adjective colors." The alums possess the property of fixing these colors, and are termed *mordants* or biters, as they are said to bite the color into the fabric.

237. Phosphates. Universal Distribution.—Nature provides for all those dependent upon her. Being essential for all forms of life, the phosphates are distributed in small but generous amounts in the different kinds of soil. Plants require them for growth, and thirty per cent. of all bones is calcium phosphate. When unhindered, all thing are ordered for the perfect growth of each member of Nature's kingdom. Thus we find the grandest trees and the graceful deer in the uninhabited forests. But where man steps in and disarranges Her plans, he must then, by his own efforts, make good any drain he has caused upon Her goodly store. Nature had fully provided for those trees and deer. Let man clear away the forest and plant the field. By the removal of the wood, he has carried away much that, in decaying, would have been valuable for other trees in their turn. Now he plants the field, harvests the crop, bearing away year after year great quantities of plant food. Soon "poor land" results, as Nature cannot forever, unaided, supply the demand.

238. Phosphate Mines.—Phosphoric acid is one of those valuable foods removed every year. Man has learned that he must replace every kilogram of phosphoric acid he removes if he wishes to continue to derive profit from tilling the soil. Many years ago, long before the time our histories tell us of, the bones and teeth of many animals were thrown together and buried by the action of the water and other geologic agencies. At this time these deposits are very valuable to us, as those bones and teeth are mainly composed of neutral calcium phosphate. This form of phosphate is insoluble in water, and less directly available for plant food than it would be if it were soluble. The phosphate rock is therefore mined, crushed and treated with concentrated sulphuric acid. This treatment causes the formation of the hydrogen calcium phosphates, one of which is soluble in water. This is called "acid-phosphate," and is an important constituent in all fertilizers.

EXPERIMENTS.

I. Paste a narrow strip of paper, 5 mm. wide, around the edge of a coin. Spread with a camel's hair brush a milky solution of calcium sulphate over the surrounded surface of the coin, carefully removing all air bubbles. Now fill up the cavity with a thick paste of plaster of Paris in water. Allow to dry, and carefully remove the paper strip and coin.

II. Taste a little alum. Heat some on charcoal. Observe all the changes closely. Intumescence.

III. Dip one end of a strip of cotton cloth into a strong solution of potash alum, then into a decoction of madder (a dye-stuff). Remove and wash. Repeat the experiment, substituting iron alum. Decoctions are made by boiling the substance in water.

LESSON XXXIV.

SILICATES:—NATURAL AND ARTIFICIAL.

239. Natural Silicates. — There is an element called silicon, very similar to carbon. The element does not occur free, however, but combined with oxygen it is very abundant, composing about one-half of the earth's crust. Quartz and the whitest varieties of sand are almost pure silica, silicon dioxide. The free acids, and there are several, can be prepared in the laboratory, but they never occur in nature. Their salts, brought about by the union of silica with the different metallic oxides, constitute our rocks and soil. Of these natural silicates we might note feldspar, a double silicate of aluminum and potassium, and mica, sometimes improperly called isinglass, a complex silicate of the same metals and sometimes with iron.

240. Artificial Silicates.—Besides the great number of natural silicates there are some artificial ones of the greatest importance and use in the arts and our every-day life. The silicates are as a rule insoluble in water, but there are two, sodium and potassium silicates, which, because they are soluble, go by the name of "water glass." This is much used by soap makers and calico printers. It is also used to make artificial stone, for when mixed with calcium hydroxide or carbonate it forms calcium silicate, which is not only insoluble in water, but acts as a cement for sand and rocks.

241. Mortar. Cement. Ordinary MORTAR is a mixture of slaked lime with sand. When most of the water of the mixture dries out the crystalline calcium hydroxide is formed. This acts as a binding material to hold the sand in place. Carbon dioxide is slowly absorbed from the air, and the hydroxide gradually changes after long exposure to the firmer calcium carbonate. Ordinary mortar can be washed away or disintegrated by water. We frequently need a similar binding material for bridge piers, wharves, etc., that are exposed to the continuous action of water. HYDRAULIC CEMENT is prepared by igniting a mixture of limestone, clay, and powdered quartz. When stirred up with water this mixture hardens, which depends upon the formation of the insoluble calcium silicates and aluminates.

242. Glass.—The most important use of sodium and potassium silicates is in the making of glass. The various kinds of glass of commerce are mixtures of highly siliceous sodium and potassium silicates with the silicates of calcium or lead. Most of the natural silicates are crystalline, but the transparent value of glass depends upon its amorphous condition. The silicates of the alkali metals, when not too siliceous, are amorphous, but soluble in water. The silicates of the other metals are insoluble, but tend to crystallize when cooled from the molten state. Window glass must not only be transparent, but withstand the solvent action of the rains; therefore, the silicates of sodium and calcium are mixed in such judicious proportions as to retain both the non-crystalline and insoluble properties. Some cheap forms of window glass scale, and become dull as the result of the action of air and water.

243. Varieties of Glass.—Bohemian, or HARD glass, so much used by chemists, consists of the silicates of calcium and potassium. It is only slightly acted upon by acids and alkalies, and melts only at a high temperature. FLINT GLASS contains lead, causing it to melt more easily. On account of its valuable optical properties it is used for making lenses for telescopes and other optical instruments. As these lenses must be cast, cut, and polished, a soft glass must be used. Ordinary bottle glass is composed of the silicates of calcium, iron, aluminum and potassium. As many metallic silicates are colored, the oxides of these metals are mixed with the silicates to produce the various colored glasses. For instance, green bottle glass is due to the presence of ferrous silicate; cobalt gives a blue, manganese a violet, uranium a green-yellow, copper-green, or red and gold ruby red.

244. Borates.—Borax (sodium tetraborate) acts very like the silicates. The glasses formed are much more easily melted.

EXPERIMENTS.

I. Take some plastering or mortar from an old building, and add HCl to it. Pass the gas through lime water.

II. Pour 30 cc. water glass into a vessel; pour on it half as much dilute sulphuric acid, having care that the two do not mix. Pour immediately into another vessel; silica should separate in long tubes resembling stalactites.

III. Powder finely some window glass in a mortar; place some of the powder upon moist red litmus paper, and allow it to remain there for several minutes.

IV. Try to dissolve some sand in water. Mix some of it in a powdered state with four times as much sodium carbonate and potassium nitrate, and heat strongly with the blowpipe for ten minutes on a piece of platinum or charcoal. When cold, test solubility in water.

LESSON XXXV.

THE SOIL.

245. The Soil.—What we are so accustomed to see is not likely to attract our close attention. All admire a beautiful green field of grain undulating with the breezes. A hillside covered with golden stacks of ripe oats is a beautiful sight. But how little thought do we give the ground, that repulsive dirt which fostered and nourished Nature's lovely growth! It is well to shun filth, but the soil should claim not only our attention and respect, but affection, in fact. All animal and plant life are directly or indirectly dependent upon the soil's generous allowances.

246. Source of Soil.—In our other nature studies we learn more about the soil. For our purpose, however, we may say that the soil is decomposed rock, the disintegration being brought about by nature's agents—the weather, oxygen, erosion, plants and animals. All of us have seen stones worn away and crumbled by long exposure to the action of the atmosphere, rain and changes in temperature. In many cases, this powdered rock is borne away by the action of water, either as rain, streams or springs, and deposited anew in some other place. Such soil as is found in the Mississippi bottom lands and the Yadkin Valley, is called alluvial. It has been estimated that the Mississippi river carries, dissolved, one hundred and fifty million tons of matter, and three hundred and fifty million tons of suspended matter annually, into the Gulf.

247. General Similarity in Composition of Soils.—Since the soil is decomposed rock, it contains the constituents of the original rocks. As the kinds of rocks, dependent upon their mineral constituents, vary in different places, the soils should be unlike. Soils do vary in character, of course, but most of the rock constituents are a little soluble in water. The continuous action of the water would thus carry portions from one locality through and across the surface of the soil and over great distances to another resting place. The winds and surface water, with burrowing worms, and plants, too, tend to alter, change, transport and mix the soil, so wherever it is formed there is a general similarity in composition.

248. Elements Present in Soil.—Of the many chemical elements found in the soil, the following are the most abundant and are of the greatest importance from an agricultural standpoint: Non-metals—oxygen, silicon, carbon, sulphur, chlorine, phosphorus, fluorine, nitrogen and boron; metals—hydrogen, aluminum, calcium, magnesium, potassium, sodium, iron and manganese. Silicon and oxygen are the most abundant elements in the soil, the former always being combined with the latter. This oxide, silica, which we know as quartz, is a very hard insoluble substance, hence is always most abundant in soil, as it wears so slightly and is not easily washed away. This sand is present even in the very clayey soils.

249. Essential Elements.—The clay of soils is dependent upon the third most abundant element, aluminum. This metal is a prominent constituent of feldspars and the different micas, rocks which, on decomposing, produce clay. The other elements mentioned are all quite necessary when the soil is called upon to fulfil its greatest duty, but they are present in much smaller amounts than the other three. In relation to the soil, carbon is present in humus (organic matter), carbonates and gaseous carbon dioxide. Sulphur is present in iron sulphide (fool's gold) and sulphates. Hydrogen, not free, but combined with oxygen in water, is most important. Phosphorus as phosphates, is a constituent of the oldest rocks, hence, through the process of decay and rock building, it has found its way into the soils of all geologic ages. The phosphates are absolutely essential, and are always present, as are also calcium, magnesium, iron and manganese. Potassium is one of the indispensably necessary plant foods, and it is almost universally distributed in the soils, but its salts are very soluble in water. Sodium is very similar, chemically, to potassium, and is always present, but somehow it cannot in any way take the place of potassium in the life of land plants.

250. Nitrification.—Nitrogen, though so abundant in the atmosphere, is least abundant as a rock ingredient in the earth's crust. It is present in the fragments of decaying tissues of plants and animals. Plants secure much of their necessary nitrogen by the action of the rains which scrub the air, removing the ammonium compounds with nitric acid.

251. Bacterial Inhabitants.—Besides these chemical elements existing in all kinds of combinations, there are infinite numbers of lower microscopic forms of life which accomplish wonders, especially in the process of nitrification—that is, the rendering of nitrogen available to the plant. It is through this agency that the organic matter is caused to decay, thereby changing the nitrogen into nitric acid, or its salts, as potassium nitrate, which becomes at once useful as plant food. Each farmer has an unpaid corps of faithful laborers.

252. Source of Heat.—The soil permits the sunshine to accomplish the great work of nature. The manner in which soil is tilled, permitting loose disintegration for draining away the excess of water and maintaining the proper amount during drought, and the way in which it is otherwise strengthened, have a great deal to do with the amount of work the sun's heat is able to do.

LESSON XXXVI.

SYNTHETIC LIFE.

All of us know that

> "Large streams from little fountains flow,
> Tall oaks from little acorns grow."

But how does it happen?

253. Plant Life.—The acorn, the fruit of the oak tree, contains the seed, which under proper conditions should give rise to another oak. From your botany lessons you likely learned the structure of the acorn. Buried under the sod, through the influence of moisture and the sun's heat, the tiny plantlet sends up its plumule and downward its root. Nourished by the rich food encased in the shell, this product of vital force, which man can neither give nor create, heaves aside the crust of the earth, and its two tender leaves greet the sunshine. Immediately this small but perfect object exercises acute judgment. Its little root noses around for some soluble food, as an infant takes its milk. The leaves gather strength and life from the sun and air, and the full grown tree in turn is evolved from one tiny cell. All life is such ; all life begins in just this way. No one can explain it. It is a point at which man with all his learning must stop and acknowledge the existence of some Higher Being, who can place life within a grain of wheat from which may rise an hundred more.

254. Osmosis.—Of the seventy known elements about sixteen are invariably found in plants of different kinds. These elements are not found free in any instance, but in almost innumerable combinations. The plant existing half in the air and half in the soil, has two channels through which it may obtain its food. The roots and roothairs are covered with a porous membrane, through which all material entering the plant from the soil must pass. If we separate two liquids by a porous membrane, each tends to pass through the membrane into the other. The lighter one moves much more rapidly. This diffusion of liquids through membranes is called OSMOSIS, the principle upon which the acquisition of all liquid plant food from the soil depends. The pressure upward in plants resulting from this osmotic action is very strong at times.

255. Liquid Vehicle.—Water, as we have seen, holds many of the salts of the soil in solution. The water carrying these salts of potassium, calcium, magnesium, sodium, iron, etc., usually as phosphates, nitrates, chlorides or sulphates, is sucked through the root membrane. Once in this wonderful little cellular workshop, the plant forms and fashions each to be most suitable for its use.

256. Tobacco.—Some plants have peculiar appetites and make a special demand upon the soil for their gratification. Tobacco is especially greedy about potassium. About eight per cent. of the ashes on the end of a cigar is made up of that element. In fact, its use is not only a wasteful, expensive yet gratifying luxury, but a plant which produces very exhausting effects upon land. For example, fourteen tons of wheat, fifteen tons of corn, or twelve tons of oats remove no more of the principles of fertility than does one ton of tobacco.

257. Abundance of Plant Food in Soil.—Each plant requires more of a certain kind of food than others. Each fertile soil by calculation contains, when new ground—that is, virgin soil—enough plant food to supply alternating crops of corn, oats, and clover for several hundred years, but with successively decreasing yields. If any one crop be grown continuously upon the same spot of ground, necessarily that form of inorganic food most required will gradually become exhausted in part, if not altogether. This exhaustion must be prevented by artificially replacing that which has been removed. Now phosphoric acid must always be present, also calcium and magnesium are necessary, as they are found in the seeds of the plants grown. As strange as it may seem, the refractory metal iron is necessary for the plant's growth. Hydrogen enters the plant combined with oxygen as water. Not a blade of grass nor a stalk of corn can grow without the gaseous element nitrogen. Although the plant is surrounded by millions of tons of it, it tastes it not as such.

258. Necessity for Fertilizers.—The plant thus extracts rich food from the earth. Dying in place and decaying where it knew its birth, it returns to the soil all that was loaned for its temporary use. Thus primeval forests always present a most luxuriant growth. If these plants be removed with their borrowed riches, a generous soil lends it aid as long as its impoverished resources permit, but in time they become exhausted and "poor land" results. Man, for it is the result of his work, must make good the loss, if gain is still to be had. This is the cause of our having to fertilize the fields. Those substances which can least afford being removed, but are as a rule carried away, are phosphoric acid, potassium, and nitrogen compounds. The amounts removed depend upon the crop. For instance, an acre of wheat yielding twenty-five bushels requires in straw and grain twenty-two kilograms of ammonia. Results of careful experiments show that under the most favorable circumstances only five kilograms of ammonia would be carried to the soil by rain. If all that were assimilated, seventeen kilos would still have to be added to meet the wants of this one wheat field. In actual practice this is not quite so, for as we learned there is another source from which much nitrogen is derived—namely, bacterial agents.

259. Kinds of Fertilizers.—Fertilizers are made up in accord with the results of these experiments, the amounts of these three constituents varying with the kind of crop to be fertilized. The phosphoric acid is obtained as bones or phosphate rock, treated with sulphuric acid, and utilized as "acid phosphate." Potassium is used in the form of any of its salts, or in cotton-seed meal. Nitrogen is added as ammonium compounds or nitrates or nitrogenous organic material, as fish scale or scrap. Farmers sometimes think that the more the fertilizer resembles the black rich dirt of the woods the more valuable it is. Although that has nothing whatever to do with the value of the fertilizer, manufacturers frequently darken their fertilizers with soot or powdered charcoal.

260. Fixation of Carbon.—Now, there is one element which constitutes by far the greater portion of the plant, if we except the constituents of water, and which is absolutely necessary for its growth, and yet the farmer does not have to pay for it. That brings us to the other road by which plants acquire food—the leaves. You who have studied botany know that each leaf has many mouths, called stomata, through which we might say they breathe. Just underneath the thin outer skin of the leaf in the tiny cells there exists a complicated chemical compound, which contains carbon, hydrogen, oxygen, and iron, called chlorophyll. This is the green coloring matter of the leaves. Through the tiny leaf mouths the sluggish gas, carbon dioxide, passes; combining with water, which has come up from the roots, this gas passes into the little cells, where, under the influence of the sun's rays and the guidance of the chlorophyll, new fiber is generated and oxygen gas is liberated. The oxygen is breathed out from the pores, and diffuses into the air. The fresh air, hovering over a dense forest, thus contains a trifle more oxygen than does the ordinary air. In fact, this is the source from which we obtain all the oxygen we breathe.

EXPERIMENTS.

I. Place in a dish spring water, to which has been added one-fourth as much carbonated water, obtained from any soda fountain. Place several sprigs of mint or other succulent plant in an empty quinine bottle, and fill it with the water. Invert the bottle in the dish, being careful that no air gets in. (Fig. 13.) Expose to direct sunlight for two hours. Carefully remove the plants, detaching the bubbles from each leaf, without admitting any air. Place a card under the mouth of the bottle and invert. Test the gas with a lighted taper.

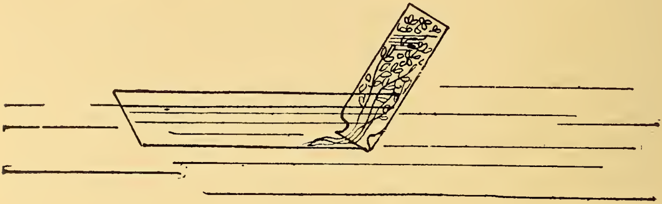


FIG. 13.

LESSON XXXVII.

THE CYCLE OF LIFE.

From our last lesson we learned that plants are always gathering together simple materials, and constructing complicated highly organized products. Now what part does the animal play in Nature's grand system of economy?

261. Animal Growth.—By the goodness of the Divine Being we are given life. Neither chemistry nor any branch of science can or dares to explain that. Our existence is the result solely of God's wonderful beneficence. Science seeks to explain as far as possible what takes place under the influence of that God-given power, life. Suppose we take any one of us young people as an example. We have life; our tiny bodies were given us to begin with, and we by our efforts have grown. To cause the body to increase in size we must furnish the material for its upbuilding in different forms of food. We have learned something of the chemical elements which enter into plants. Directly or indirectly we live upon plants, therefore the substances which build up our structure serve us in the same way. The architect, life, builds not only according to the material furnished, but fashions each as it seems best.

262. Cooking.—By means of the digestive organs and various liquid systems learned of from your physiology, the food material is made available, and properly distributed throughout the body. It may be well to remark here that while some kinds of food stuff would require less work on the part of the animal body to digest if it were not cooked, that the cooking of food makes it more palatable, and promotes a greater desire for its consumption, just as proper seasoning of food causes the more ready flow of the digestive fluids, hence greater ease in digestion. For example, the albumen of an egg, the white, is thoroughly digestible raw, but if it be cooked, hard boiled, it becomes less digestible and requires great expenditure of energy to get it once more into an available state.

263. Similarity in Composition.—Arising from the glowing embers of a fire is the colorless gas, carbon dioxide, the result of the oxidation of the carbon in the coal or wood. Carried out over the earth in the air this gas is sucked in by the bright-green leaves, and serves as the foundation principle of the prettiest rose or sweetest sugar or rankest smelling onion. This plant material is then utilized by man as food at once, or he eats the ox that has devoured it. Thus we see that carbon may in time give beauty to the cheek of our prettiest girl friend, the result of chemical action and life force. We are all composed alike. Carbon has been selected merely as an illustration of a principle which is true of all the elements entering into the composition of the animal body.

264. Air a Necessity.—Now men have been known to live quite awhile without food. Dr. Tanner fasted forty days. Shipwrecked sailors on the ocean have been known to remain without food and water for six days, but no man has ever been heard of who lived for more than a few minutes without breathing. Although food is essential for our upbuilding, growth, even our mere continued existence, and while water is also necessary as a vehicle of food, and for keeping up the liquid state of the many animal fluids, yet the air is the most essential. Why?

265. Animal Heat.—The air, active oxygen diluted with inert nitrogen, is carried into the lungs where it comes into close contact with the blood by diffusion through a thin membranous wall. A constituent of the blood, hæmoglobin, a chemical compound, has the power of combining with oxygen and bearing it away from the lungs throughout the body. When this oxygen is thus carried into the capillary blood vessels, it becomes exceedingly active and burns up certain portions of the animal body, mainly carbon compounds. We have seen that when chemical action takes place heat is produced. Not only is our temperature kept up to the normal by this, but carbon dioxide is liberated. It is dissolved in the blood and borne back to the lungs, and there turned loose in the air, where it is rapidly diffused. Pure air breathed into the lungs returns ladened with impurities, carbon dioxide being the greatest.

266. Animal Life Analytic.—Thus we see that plants are forever withdrawing carbon dioxide from the air and building up food material; animals are using that food and throwing the carbon dioxide back upon the air. Plant life is therefore on the whole synthetic, upbuilding; animal life is analytic or breaking down. Yet with all these great changes the amount of carbon dioxide in the air remains almost constant, about four parts in ten thousand. This has been so for years and is likely to continue.

267. Debt to God.—When we stop to think, although seemingly a small matter, this is a most delicate balancing of our very existence. No such nice arrangement could be maintained except by an all powerful, and, as it is exactly suited for our best existence in this world, an Allwise Providence. And chemistry, to which some have attributed an opposition to religion, can do nothing but beg all to join in praise to God " from Whom all blessings flow."

APPENDIX.

Apparatus and Chemicals Required.

10 common Blowpipes.
Sheet platinum, 2 x 10 cm.
10 Funnels, top 5 cm. diam.
1 Thermometer, 0—360° C.
3 lbs. soft glass Tubing, 4 mm. diam.
Corks, 2 doz. each 1 and 2 cm. diam.
Fourth quire Litmus paper, half red.
100 Filter papers, 595 S. & S. 11 cm. diam.
5 doz. Test Tubes, 2 cm. x 15 cm.
10 porcelain Dishes, 5 cm. diam.
2 small porcelain Mortars.
12 round bottom Flasks, 100 cc. cap.
1 book Dutch Leaf.
Platinum Wire, 1 m. (small).
Iron wire Gauze, 12 cm. x 1 m.
12 Erlenmeyer Flasks, 250 cc.
3 metres soft rubber Tubing, 4 mm. diam.
Common Bottles, 12, each 200, 250 and 500 cc. cap.
1 quire Absorbent Paper.
10 Iron Tweezers (small).
5 Clay Pipes.
3 Graduates, 50 cc.
10 pieces Mica, 2 x 6 cm.

Acetic acid............	½ kilo.
Hydrochloric acid...	3 kilos.
Nitric acid............	2 kilos.
Sulphuric acid........	2 kilos.
Tannic acid...........	150 grams.
Ammonia............	1½ kilos.
Ammonium chloride	120 grams.
Ammonium nitrate.	120 grams.
Sodium (metallic)...	60 grams.
Sodium silicate (water glass)....	½ kilo.
Sodium hyposulphite..............	¼ kilo.
Sodium hydroxide, crude..............	½ kilo.
Iron sulphide.........	½ kilo.
Iron alum..............	30 grams.
Iodine....................	10 grams.
Potassium iodide....	30 grams.
Potassium (metallic)	15 grams.
Potassium chlorate,	¼ kilo.
Potassium nitrate, crude cryst......	120 grams.
Lead oxide (red lead)	120 grams.
Lead nitrate..........	120 grams.
Lead acetate (com.),	½ kilo.
Sulphur (brimstone)	120 gr.
Antimony powder...	15 gr.
Magnesium wire......	30 gr.
Magnesium powder,	60 gr.
Silver nitrate.........	30 gr.
Bone black............	250 grams.
Carbon disulphide..	500 grams.
Copper sulphate, crude..............	250 grams.
Bleaching powder...	500 grams.
Aluminum wire.......	30 grams.
Saccharine.............	1 gram.
Mercuric chloride...	15 grams.
Strontium nitrate....	250 grams.
Barium nitrate.......	250 grams.
Madder..................	30 grams.
Phosphorus............	60 grams.
Manganese dioxide powder...........	500 grams.

The above, enough for ten pupils, may be procured for twenty-five dollars from Eimer & Amend, New York. One dollar per pupil should cover cost of material and breakage.

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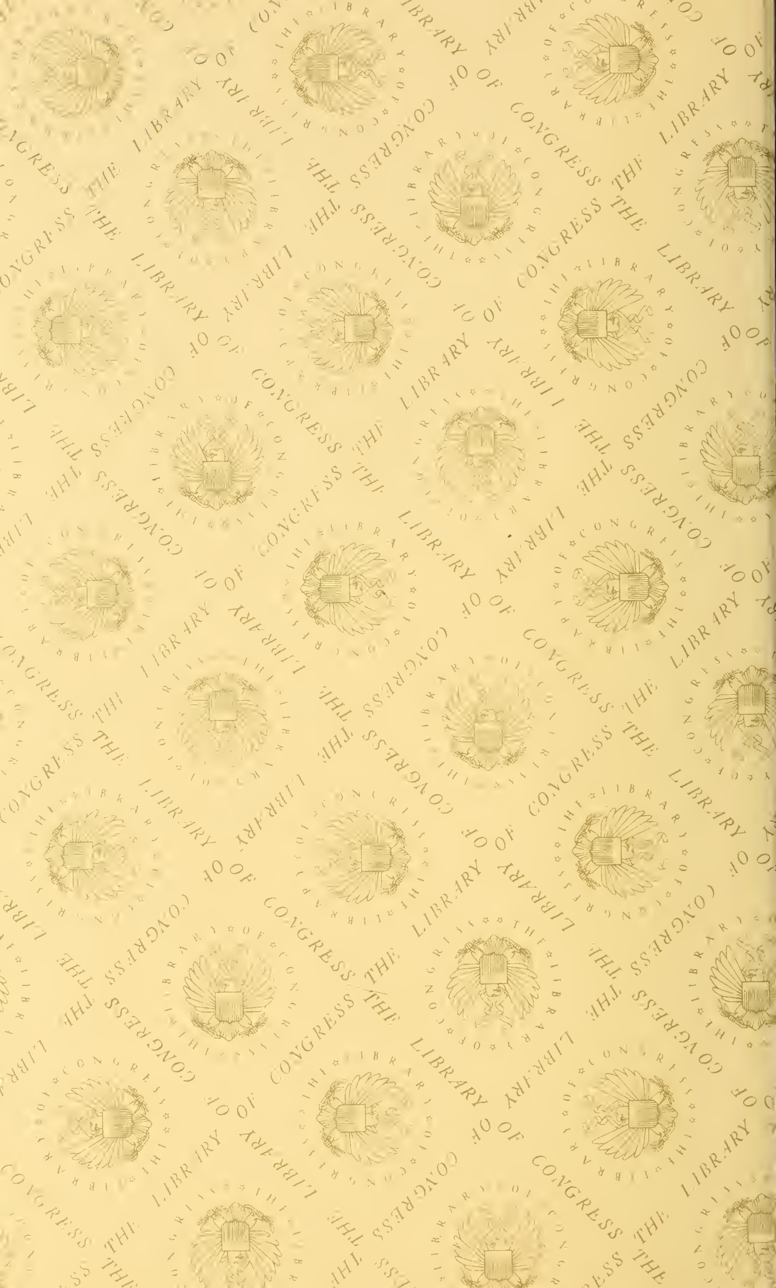
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