THE NEW TEXT-BOOK OF CHEMISTRY

FOR USE IN

HIGH SCHOOLS AND ACADEMIES

BY

LE ROY C. COOLEY, PH.D.

PROFESSOR OF PHYSICS AND CHEMISTRY IN VASSAR COLLEGE

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PREFACE.

THE sole purpose of this volume is to provide an elementary course of chemistry suited to the wants of students and teachers in high schools and academies.

With this purpose in view I have endeavored, —

I. To seize upon the fundamental facts and principles of the science, and submit them to a simple and concise, but at the same time a clear and accurate, treatment.

II. To present the subjects which will be of greatest value to a student whose course of study must end when he leaves the preparatory school, and which are, at the same time, most desirable, as a preparation for higher courses, to those who may in future pursue them.

III. To thoroughly systematize the whole, securing a logical order of subjects and an arrangement of topics adapted to encourage the best methods of study and to facilitate the best methods of teaching.

IV. To encourage a complete mastery of what is undertaken, by concentrating study upon a brief course, and providing for thorough reviews at short intervals by means of summaries of principles and exercises with problems at the ends of sections or chapters.

V. To present the science in the light of modern theories, sparing no pains to make the work represent the *present state* of chemistry fairly and fully, so far as its elementary character will permit.

VI. To make the experimental method of reaching facts prominent and practicable, not by requiring the student to read the descriptions of large numbers of experiments which he can not make and may never see, but rather by inviting his attention to a few which are within reach, or which may be clearly understood by means of cuts, and then providing an additional number which are easily and cheaply made, and *which are to be studied only when they are seen.*

Chemistry is eminently an experimental science. Nothing whatever can be learned in it by speculation, and very little by direct observation of nature. All its laws and principles have been established by experiment. The experimental method is therefore the method of study which is true to the character of this subject, and it should be prominent in an elementary course.

Nevertheless it is not thought that a student should be compelled to learn every thing in chemistry by the study of experiments. He will acquire a larger knowledge and a good discipline if he take many established facts and laws without witnessing the experimental evidence on which they rest. And, indeed, this he must do, or accomplish little, so long as chemistry is compressed into the small space which it usually occupies in our schools. But while so much of the world's progress in art, in industry, and also in those beliefs which sway the minds and characters of men, is based upon the experimental method of reaching the facts and laws of nature, chemistry, above all other sciences, should take to itself the privilege and the duty of teaching the young to know what an experiment really is, how to translate it, how to weigh its testimony, and how, by means of this method, facts may be revealed and laws established.

These considerations have prompted me to furnish two sets of experiments. I have in the first place illustrated the text by experiments, the study of which constitutes a part of the student's daily preparation. These experiments should be witnessed if possible, but will, in any case, impress the fact that the principles of chemistry rest on experiment for their demonstration. In the second place, I have provided, *for the teacher's use*, a selection of *additional* experiments, easy, cheap, and appropriate for illustration, which may be used at discretion, and which, having the charm of novelty to the student, will more effectually fix his attention, and arouse his thoughts.

LeR. C. C.

June, 1881.

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CHAPTER I.

GENERAL PRINCIPLES OF THE SCIENCE.

SECTION I.

PHYSICAL AND CHEMICAL CHANGES.

1. Physical Changes. — Bodies of matter are constantly changing. They move: how varied are their motions! They change their shapes, as when rocks are rounded by the flow of water over them, or shattered by a blast of gunpowder. They change from solid to liquid forms, as when the snows of winter melt; and from liquids to gases, as when water passes into steam. In such changes as these, however, the nature of the substance is not affected. After a body has moved, it is the same body as before. Water, in the forms of ice and dew and vapor, is water; and, if it change from one to another, it is water still. Such changes are called Physical Changes. They are changes during which the nature of a substance remains unaltered.

2. Illustrated by Experiment. — A simple experiment with an apparatus shown in Fig. 1 will show that air will yield to the slightest pressure, and suffer a change in its volume and in its density.

A glass tube with a bulb at the upper end is joined to another by a piece of rubber-tubing. A colored liquid fills the rubber, and stands at equal heights in the glass tubes. The bulb and stem above the liquid are filled with air. Let the lips be applied now to the top of the open tube, and the breath be gently forced into it: the fluid quickly responds to the pressure, and rises towards the bulb, pushing the air before it into a smaller space. The volume of the air is diminished, and its density is increased.


Fig. 1.

But the nature of the substance is not altered. The same air remains, after the pressure is exerted, as before. These changes in volume and density are examples of physical changes.

3. Chemical Changes. — But all changes are not like these. Wood burns: in doing so it ceases to be wood; it is changed to smoke and ash. Gunpowder explodes: it is no longer gunpowder. Fluids from the soil, and gases from the air, are taken into the roots and leaves of plants, and are there changed into substances which form the plant itself. Such as these are changes in the nature of substances. They are called CHEMICAL CHANGES. They are changes in which we discover the production of new substances, substances totally unlike those which enter into the action.

4. Illustrated by Experiment. — A piece of cardboard is put over the top of an ale-glass (Fig. 2). A teaspoonful of sugar and another of potassium chlorate are powdered and mixed, and then laid upon the cardboard. Three or four drops of strong sulphuric acid are allowed to fall from the end of a glass tube upon this mixture. Almost upon the instant when the acid touches the powder, violet-colored tongues of flame leap up from it with a hissing sound, accompanied with large volumes of white vapor passing off into the air. When the sound has ceased, and the colored flames have died out, only a coal-black mass is left upon the card-board.

The white mixture of sugar and potassium chlorate, with the drops of oily-looking acid, have been changed into vapors, which have gone into the air, and a black coaly mass, which is left behind. Not a particle of either of the substances used can be found remaining.

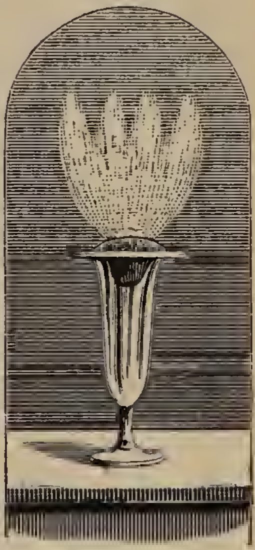


Fig. 2.

Burning of Sulphur in Oxygen. — Another illustration is seen in the action of oxygen gas and ignited sulphur. A large bottle is filled with oxygen. A small pendent spoon is fixed to the cork, and filled with powdered sulphur. The sulphur, first set on fire, is thrust into the bottle (Fig. 3), when the combustion goes on with a fine blue flame, and the bottle is soon filled with whitish vapors.

The sulphur has now disappeared. Much of the oxygen has been also used up. The new gas formed contains them both, but its properties are not like those of either. It is very suffocating when breathed: oxygen is not. Pour a little solution of blue litmus into the bottle: it is first reddened and then bleached. Neither oxygen nor sulphur can bleach blue litmus. The mutual action of the sulphur and oxygen is called a CHEMICAL ACTION, because it has brought about a change in the nature of these substances, producing a new kind of matter having properties different from those of either sulphur or oxygen.

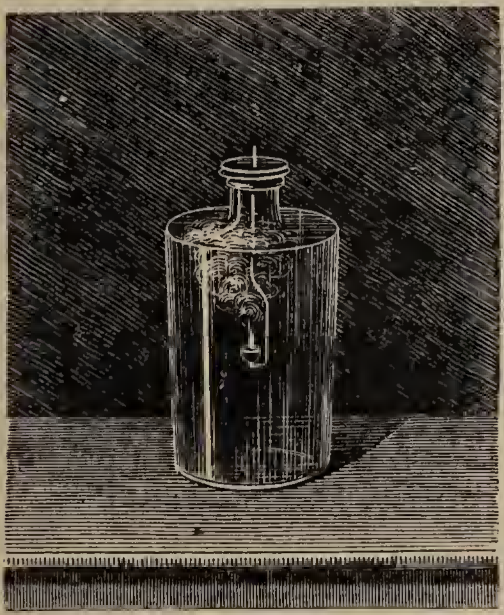


Fig. 3.

5. Chemical Combination. — In this chemical action of sulphur and oxygen, we find two substances uniting to form a third. This is a case of chemical combination.

In the following experiment (Fig. 4), we may see another example of the same kind of action. A small quantity of pure mercury is put into a glass flask over a source of heat. A bent glass tube reaches from the flask over into a vessel of mercury. The flask and tube are filled with oxygen. The end of the tube is bent upward; and over it is placed a glass jar, mouth downward, also filled with oxygen. Mercury and oxygen are the only two substances in the flask. If the mercury be heated for several hours, without boiling, two things will be discovered: yellowish or red scales make their appearance upon the surface of the mercury in the flask, increasing in quantity until, if the heat be continued, the metal is completely covered. At the same time the mercury rises in the jar which contains the oxygen, showing that a part of that gas has been removed.

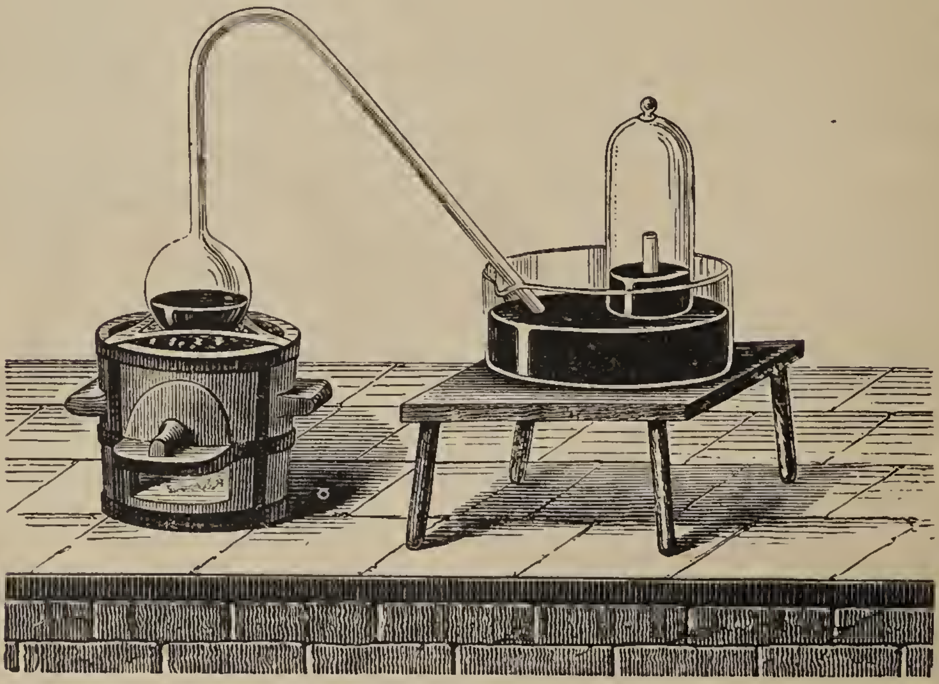


Fig. 4.

Some of the mercury and some of the oxygen have united to produce this new red substance. Their action is an example of chemical combination, since the product contains them both, and is a substance quite unlike either.

6. Chemical Decomposition.—If the red powder, which contains mercury and oxygen, and which the chemist calls mercuric oxide, is heated, both the mercury and the oxygen can be regained.

For this purpose a tube (Fig. 5) containing a small quantity of the oxide is tightly closed with a perforated cork, from which a bent tube passes just through the cork of a glass flask full of water. Another bent tube, through this same cork, reaches to the bottom of the water, while its other end passes into the neck of a second flask. By heating the tube, first the red oxide becomes black, and soon after a colorless gas will push the water over into the second flask.

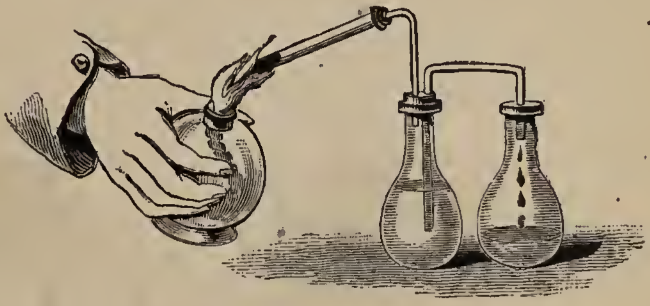


Fig. 5.

Little globules of shining mercury will be found clinging to the sides of the tube; while, if the cork and tubes be removed from the first flask and a lighted taper be plunged into it, the flame will burst into brilliant combustion, showing the gas to be oxygen.

Here we find a single substance broken into two quite different ones. This is an example of chemical decomposition.

Of Water.—The decomposition of water is another example of this kind of chemical action. It is done by means of electricity in an apparatus shown in Fig. 6.

Two tall glass tubes are first filled with acidulated water, and then inverted over two strips of platinum at the bottom of a vessel containing the same liquid. These pieces of platinum may be joined by wires to the poles of a galvanic battery. The moment such a connection is made, bubbles of gas begin to rise from each platinum, and are collected in the tubes above. After a little time the tubes will be full of the colorless gases into which the water has been converted.

If then one tube be lifted from the water, a lighted taper may be thrust into it, when the flame will instantly burst into tenfold brilliancy, showing that the gas is oxygen. But if the other tube be lifted, and the flame applied, the gas instantly takes fire, and burns with a slight explosion. This is characteristic of hydrogen.

The electricity has produced a chemical decomposition of the water: it has changed it into oxygen and hydrogen, two substances altogether different from that from which they were derived.

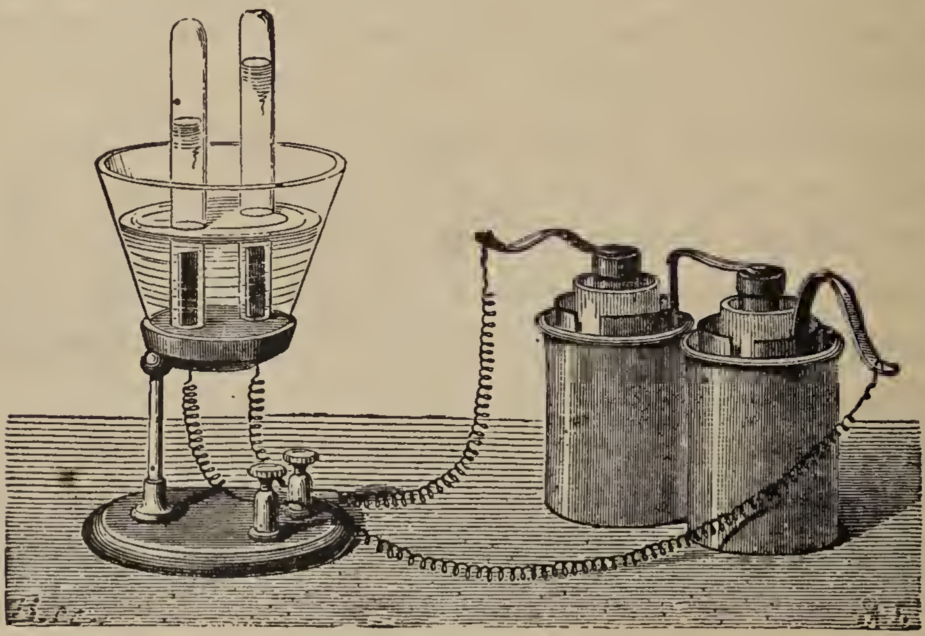


Fig. 6.

7. Definition of Chemistry. — Chemistry is the science which describes the chemical changes which occur in bodies of matter, discovers the circumstances under which they take place, and the laws by which they are governed.

8. Observation and Experiment. — Many things can be learned by simply watching what occurs in the ordinary operations of nature. The rusting of iron, for example, is a very common occurrence. All have seen the bright metal, on exposure to moist air, become covered with a dull red coating; and, since iron-rust is a substance totally different from either iron or moist air, we may say that these two substances, on contact, enter into a chemical action. This method of learning the truths of nature is called OBSERVATION.

But it is likely that we should never have learned by observation that water is composed of two invisible gases, because its decomposition does not happen in nature under circumstances in which we are able to observe it. It was discovered by means of an experiment. *An experiment is an operation performed under conditions which we ourselves arrange for the purpose.* Nicholson and Carlisle, in the year 1800, placed the poles of their galvanic battery in a vessel of water, and in this way brought electricity to act upon that liquid in a manner not to be seen in nature; and the result was the decomposition of the water into oxygen and hydrogen. This method of learning the truths of nature is called the EXPERIMENTAL METHOD.

Chemistry is an experimental science. Very little of what it teaches has been discovered by observation. Nearly every thing we know of chemistry has been established by experiment; and the student should pursue the study of it by the experimental method.

An experiment in chemistry is of no value except as it teaches some chemical truth. However attractive it may be, or however startling, if it have no other merit than beauty or novelty, it is worthless to the student of chemistry.

In order to learn by an experiment, one must see clearly the conditions under which the substances are brought together, observe accurately the changes which occur, and the products of them; and, finally, he must endeavor to see that the fact or principle, which the experiment claims to prove, is a correct interpretation of the results.

REVIEW.

I. — SUMMARY OF PRINCIPLES.

9. Physical changes are such as do not affect the nature of a substance. They are the result of *mechanical action.*

Chemical changes are such as cause an alteration in the nature of a substance, converting it into a different kind of matter. They are the result of *chemical action.*

Chemistry is the science which treats of the chemical changes in all kinds of matter, discovers the conditions under which they occur, the laws which govern them, and describes the qualities of substances which take part in them, or are produced by them.

Chemical combination is the uniting of two or more substances to form a single one with entirely different properties.

When sulphur burns in oxygen, these two substances unite to form sulphurous oxide, a suffocating gas quite unlike either sulphur or oxygen. When mercury is heated in oxygen, these two unite to form mercuric oxide, which resembles neither of them. These are examples of chemical combination.

Chemical decomposition is the breaking-up of a single substance into two or more simpler ones, with entirely different properties.

Mercuric oxide, when strongly heated, is changed into mercury and oxygen. Water is broken, by the electric current, into oxygen and hydrogen gases. These are examples of chemical decomposition.

Neither oxygen, nor hydrogen, nor sulphur, nor mercury, can be decomposed by any known process. On this account they are called ELEMENTS.

An element is a substance which has never yet been decomposed.

When two or more substances enter into chemical combination, the new substance formed is called a COMPOUND.

A compound is a substance made up of two or more simpler ones, and whose properties are unlike those of either of them.

Besides elements and compounds, there are substances called MIXTURES.

A mixture is a substance made up of two or more which have not combined. Its properties are the same as the properties of its constituents.

II. — EXERCISES.

How many kinds of changes occur in bodies?

Give examples of physical changes. Why are these called physical changes?

Give example of chemical changes. Why are these called chemical changes? *Salt dissolves in water: is the change physical, or chemical? Iron rusts: is this a physical, or a chemical change? The loftiest tree may be shattered into fragments by a lightning-stroke: which kind of change is produced?*

Define chemistry.

What phenomena are to be studied in chemistry? In natural philosophy? *Which science explains the fall of a stone? the decay of wood? the manufacture of glass out of potash and sand?*

What is meant by chemical combination?

What is meant by chemical decomposition?

What is the experimental method?

What is the method of observation?

What is an element? A compound? A mixture?

What is produced by burning sulphur in oxygen? By heating mercury a long time in air? Why are these new substances called compounds?

What are the elements in water?

How may water be decomposed?

SECTION II.

CHEMICAL ATTRACTION.

10. Attraction. — If we suspend a long and light wooden bar by means of a thread attached to its middle point, or balance it on a pivot, and if then we briskly rub a glass tube with a piece of silk or flannel, we shall find that the end of the bar will swing toward the glass, even when the glass is several inches away from it. (Fig. 7.) In this experiment we see the action of a force which tends to bring these two bodies together. Any force by which bodies tend to approach is called Attraction.

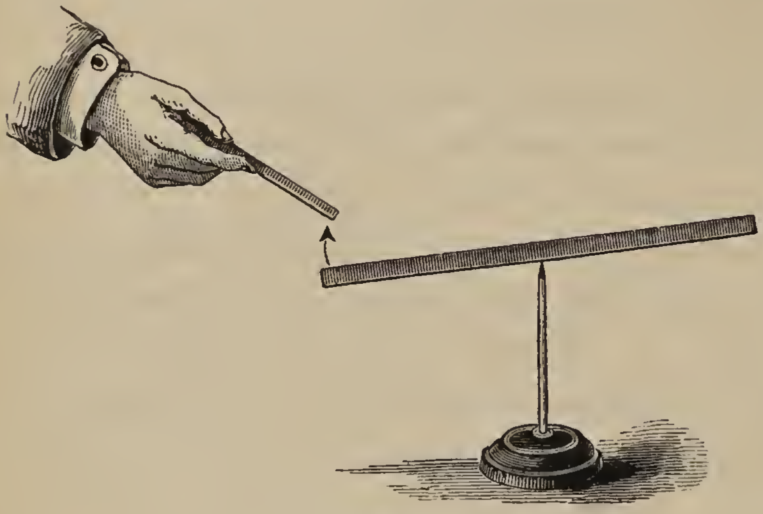


Fig. 7.

Varieties of Attraction. — We know that all bodies, if not supported, will fall to the ground: this we learn by observation. Experiments have been made which show that there is also an attraction between bodies of every size and kind. This attraction which acts between all separate bodies, and throughout all distances however great, is called Gravitation.

Again, we know that the parts of a solid body are not to be separated with perfect ease; and the attraction which holds the particles of a body together is called Cohesion.

Neither gravitation nor cohesion produces any chemical changes whatever. But we have seen that when sulphur and oxygen are heated, they fall together into such intimate contact, that neither sulphur nor oxygen can be distinguished, but sulphurous oxide is found instead. The attraction which brings these two elements into chemical combination is called CHEMICAL ATTRACTION.

The production of new substances, by uniting those on which it acts, is the characteristic effect of chemical attraction. In a word, all chemical changes are effected by the force of chemical attraction.

11. Influence of Cohesion on Chemical Action. — If we place a crystal of potassium chlorate, about one-half as large as a grain of wheat, in contact with a piece of brimstone of about the same size, no chemical change occurs. The particles of each are held together by cohesion, and kept from uniting with those of the other. But if we grind them together in a mortar, a sharp explosion follows. The chemical action converts both into gases, which disappear in the air. This experiment illustrates the fact, that, in the solid forms of matter, cohesion is generally too strong to be overcome by chemical attraction.

To weaken cohesive attraction is to facilitate chemical action. In the experiment, cohesion was overcome by *pulverizing* the solids. It may be overcome or weakened in other ways, among which *solution* and *fusion* are most important.

Effect of Dissolving the Substances. — If sodium carbonate and tartaric acid, both in the finest powder, be mixed most thoroughly, or rubbed together violently, no chemical action will take place. But, when a little water is added to this mixture, a violent chemical action will quickly follow. Let the mixture be made at the bottom of a tall jar (Fig. 8), and the vigor of the action will, very likely, carry the foam to the top and over upon the plate below.

Fig. 8.

Now, in this case, the water by dissolving the solids has overcome cohesion to such an extent, that chemical attraction can bring about a chemical change. In this way solution very generally facilitates chemical action.

Effect of Fusing the Substances. — The melting of solids together often has the same effect. If we take a little sulphur, twice as much potassium carbonate (well dried), and three times as much niter, powder them well separately, then mix them thoroughly on paper, and introduce a *little* of the mixture into an iron spoon, we may melt them together over the gas-lamp. The cohesion is overcome by the heat; and the mixture explodes violently, announcing the chemical action produced. In this way fusion very generally facilitates chemical action.

Exceptions. — Yet we must not suppose that no two solids can combine without their cohesion being first overcome by the methods just described. If, for example, a crystal of iodine is laid upon a thin slice of phosphorus, the two will shortly burst into a rapid and curious combustion. But such examples are quite rare: the liquid form is most favorable to chemical action.

12. Influence of Heat. — Although the liquid state is most favorable to chemical action, yet in many instances no chemical change takes place until the mixture is heated. A high temperature is favorable to chemical action.

We have already seen the power of heat to produce chemical changes. Sulphur was *ignited*, before it would combine with oxygen to form sulphurous oxide. Mercury was *heated*, before it would combine with oxygen to form mercuric oxide.

Attraction of Phosphorus and Oxygen. — Phosphorus and oxygen are two elements between which there is a powerful chemical attraction. Nevertheless at ordinary temperature their combination is very slow. Phosphorus is kept under water to prevent the action of oxygen in the air, which will consume it slowly on exposure, and very rapidly if its temperature be raised. The heat of the fingers may start the combustion, and hence phosphorus must be handled with great caution.

Rapid Combination by Heat. — But let us cut a thin slice from the end of a stick of phosphorus *under water*, and carefully dry it by very gentle pressure between two folds of blotting-paper. We may then place it upon a piece of card-board supported by a test-glass (Fig. 9). Delicate clouds of white vapor float away: these are the new substance formed by the slow combination of phosphorus and oxygen at ordinary temperature. Now let a red-hot iron rod be brought near. The phosphorus will burst into rapid combustion, evolving large volumes of dense white fumes, even while the hot rod is several inches away.


Fig. 9.

All these are examples of chemical combination; but heat very often produces chemical decomposition also. We have already learned that the red mercuric oxide is resolved into mercury and oxygen when strongly heated. Chemical decomposition by heat alone is called Dissociation.

Different Effects of High and Low Temperatures. — The experiments with mercury and oxygen illustrate another curious and important fact, — that a compound which is formed only at a high temperature, is again decomposed if heated to a temperature still higher. In order to compel mercury and oxygen to combine, we must heat them almost to the boiling point of mercury; and then by heating the red oxide thus formed to a still higher temperature, we reproduce the elements.

Most compounds are more readily made under the influence of high temperature; and yet at some higher temperature, it is thought, every compound may be resolved into its elements again.

13. The Influence of Electricity. — Electricity is also a powerful agent in chemical action. The decomposition of water is an example. Many other liquids may be decomposed in a similar way. This decomposition by electricity is called ELECTROLYSIS.

In the decomposition of water the oxygen was liberated at the *positive* pole of the battery, and the hydrogen at the *negative* pole. The oxygen is described as the *electro-negative* constituent, and the hydrogen as the *electro-positive* constituent, of water. Whenever a compound is decomposed by electricity, that constituent which is found at the negative pole of the battery is called the ELECTRO-POSITIVE constituent, the other is called the ELECTRO-NEGATIVE constituent.

Chemical Combination by Electricity. — But electricity may also produce chemical combination.

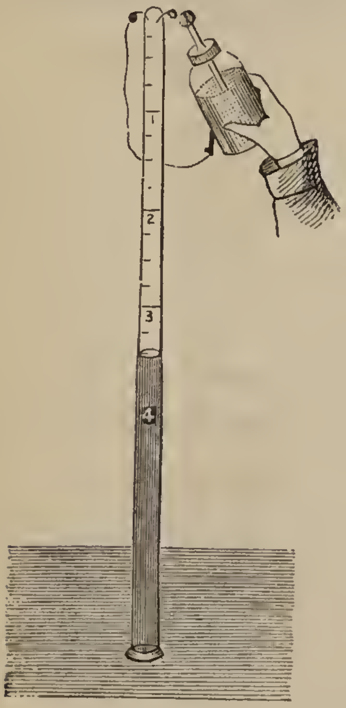


Fig. 10.

The tall glass tube represented in Fig. 10 is closed at the top, and open at the bottom. Two platinum wires are fused into the glass near the upper end: they almost touch one another on the inside. The tube was first filled with water. Afterwards a few bubbles of oxygen were introduced, and a little more than twice as much hydrogen was added. Now let a spark of electricity pass through the gases by means of the platinum wires, and a violent explosion instantly follows. The gases combine under the influence of the electricity, and disappear, only the small excess of hydrogen remaining.

14. The Influence of Light. — Both decomposition and combination are also brought about by light.

We may witness an example of decomposition if we will dampen the surface of a piece of paper with a solution of silver nitrate, and expose it to the action of the sunlight. Darker and darker the white surface becomes, until finally it is almost black. This darkening of the surface is due to the decomposition of the nitrate by light.

A large number of compounds of silver are darkened by light. The art of photography is founded upon this decomposing action of light.

Combination by Light. — Combination is also brought about by light, as when hydrogen and chlorine gases are mixed in a glass vessel. The direct rays of the sun will drive them into combination so rapidly as to produce explosion. A new gas — hydrochloric acid gas — is formed by their union.

Let a strong tube or bottle be filled with a mixture of the two gases, taking care to have a little more of one — say of hydrogen — than of the other. Let this mixture be prepared in a dark room, the tube containing it being inverted in a vessel of water, firmly fixed in place (Fig. 11), and covered with a black cloth. Thus prepared, take the apparatus at once to a place where the *direct rays* of the sun may fall upon it, and by means of a long handle remove the cloth. A violent explosion will quickly follow: the water, speedily dissolving the acid gas, will rise in the tube, and would strike the top of it with violence, did not the excess of hydrogen act as a cushion to prevent it.

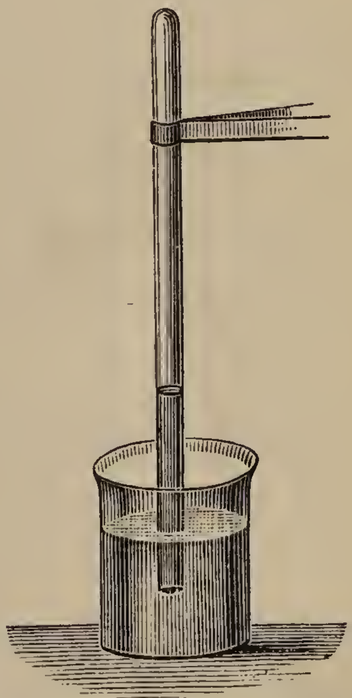


Fig. 11.

Exposed to *diffuse* light, the combination of the mixed gases is gradual; but, if prepared and kept in the dark, no combination occurs.

These experiments clearly illustrate the fact, that sunlight has power to produce chemical action. Other intense lights, the electric light for example, may be used with similar results.

REVIEW.

I. — SUMMARY OF PRINCIPLES.

15. Attraction is a force under whose influence bodies of matter tend to approach one another.

Attraction is called *gravitation* when it acts between separate masses of matter. It is called *cohesion* when it acts between the particles of a body without causing any chemical change. It is called *chemical attraction* when it causes two or more substances to produce a single substance quite unlike themselves.

Chemical attraction is opposed by cohesion. We facilitate chemical action by reducing the strength of cohesion.

Cohesion may be reduced by pulverizing the solid, by dissolving it, and by melting it. All these processes are favorable to chemical action.

Chemical action is most likely to occur when the substances are brought together in the liquid form.

The application of *heat* is favorable to chemical change. *Electricity* and *light* also, in many cases, produce chemical action.

Electrolysis is the decomposition of a compound by electricity. The constituents are distinguished as electro-positive and electro-negative.

II. — EXERCISES.

Define attraction. Name the three varieties of attraction. Define each. To which of these are chemical changes due?

What is the influence of cohesion on chemical action? Why should pulverizing solids facilitate their chemical action? By what other means may we weaken cohesion? Name two solids whose cohesion does not prevent their combination.

What is the general influence of heat on chemical action?

Give an example of chemical combination, brought about by heat. Give an example of chemical decomposition, brought about by heat. Define dissociation.

Give an example of chemical decomposition by electricity. Define electrolysis. What is meant by electro-positive and electro-negative constituents?

Give an example of chemical combination by electricity.

What is the effect of light upon silver compounds?

What is the effect of light on mixed chlorine and hydrogen gases?

SECTION III.

THE INDESTRUCTIBILITY OF MATTER.

16. No Loss occurs in Chemical Combination. — In many cases of chemical action, substances disappear and seem to be lost. This happens when a candle burns: its material slowly wastes away. Nevertheless, the quantity of matter in all such cases remains unchanged. Much or even all of it may become invisible, but by appropriate tests its existence can be demonstrated.

Test for the Presence of Matter. — Weight is the appropriate measure of the quantity of matter, and experiment will always decide whether the weight of the products of chemical action is exactly equal to that of the substances which take part in it.

Application of this Test. — Take a piece of phosphorus as large as a pea, and support it on top of a small wire, whose lower end is bent into a flat spiral so that it will stand upright. Select a thin, light beaker, and a large, thin, and light flask, of about 1,500 cubic centimeters capacity. Put about 400 cubic centimeters of water, colored blue with litmus, into the beaker, stand the phosphorus wire in it, and then invert the flask over the phosphorus, with its neck resting on and *completely closing* the top of the beaker, and its mouth reaching almost to the bottom of the colored water. The outside must be dry.

In this condition, let the whole apparatus be placed upon one pan of a pair of scales, and exactly balanced by weights in the other pan. (Fig. 12.)

If we leave the apparatus for several hours, — perhaps until to-morrow, — we shall find that water has risen into the flask, as shown in the figure, as if a part of the air had been annihilated. Moreover, we think we can see that the phosphorus has wasted away, as if a part of its substance also had been destroyed. The balance, however, *remains undisturbed :* the apparatus weighs exactly as much as at first, showing that no loss can be detected. And yet, if what substance has disappeared had been destroyed, the apparatus would weigh at least .43 of a gram less. A good common balance would easily detect much less loss than that.

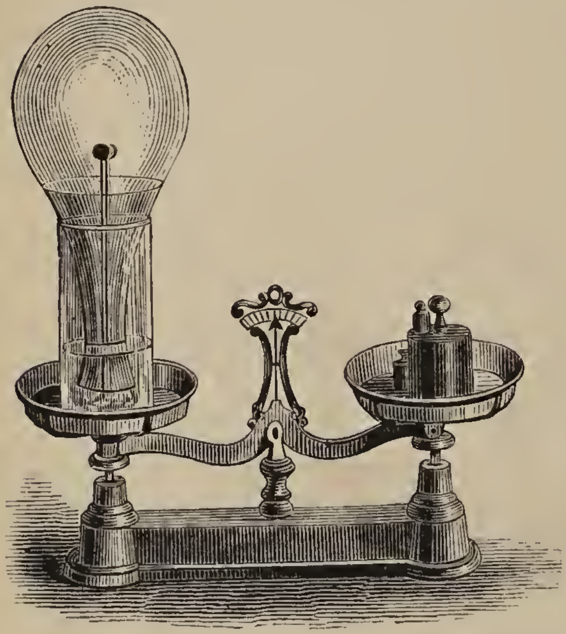


Fig. 12.

What has actually happened is this : The phosphorus has combined with oxygen, which is one of the constituents of the air in the flask. The compound, phosphorous oxide, is absorbed by the water, and remains unseen. This new substance contains all the oxygen and all the phosphorus which seem to have been lost.

In all cases of chemical combination, the compound weighs exactly the same as the constituents used up in producing it.

17. No Loss in Chemical Decomposition. — In chemical decomposition, also, the quantity of matter remains unchanged. The red mercuric oxide is decomposed by heat: let us get the facts by experiment with this substance. We must make the experiment so that both constituents shall be preserved in the apparatus. (Fig. 13.)

The Experiment. — Put about three grams of the red powder into a small tube of hard glass, which is provided with a bent tube reaching just through the cork of a flask of about two hundred cubic centimeters capacity. Let this flask be joined to another of equal size, by means of a bent tube which reaches almost to its bottom, but only just through the cork of the other. The first flask is to be nearly or quite full of water, and *must be air-tight* at the cork; while the second is empty and loosely corked. Place the whole apparatus, which should be perfectly dry outside, upon the scales, as shown in the figure, and accurately balance it by weights. Then heat the tube, and decompose the red oxide. The powder will waste away; globules of metallic mercury will collect on the side of the tube above the heat; while the oxygen will pass into the first flask, and drive the water over into the second.

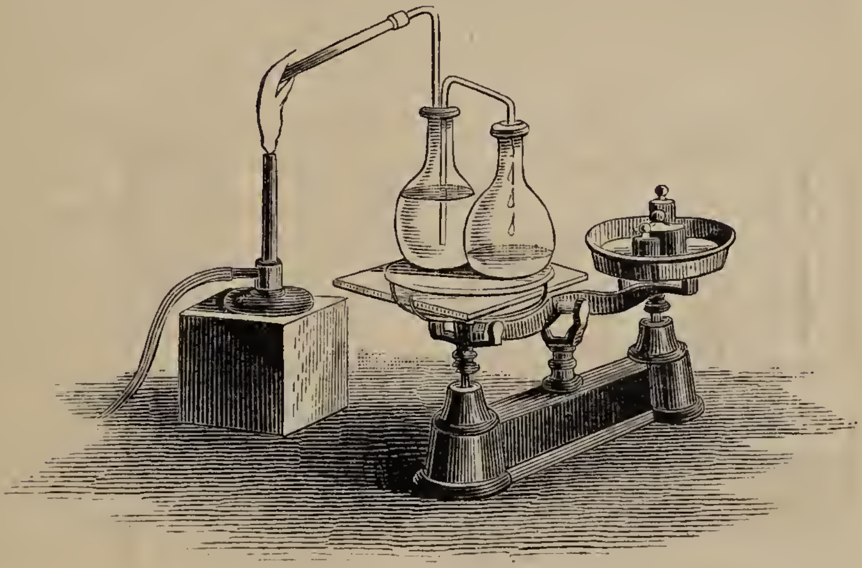


Fig. 13.

Leave the apparatus at rest for a time until the tube has become cold, and it will be found that the balance is undisturbed. The whole weight is just the same as at first, showing that no loss of matter has occurred during the decomposition.

General Results. — Many accurate experiments have proved beyond doubt, that matter is never destroyed nor created by chemical action. During all the chemical changes through which substances may go, the *quantity of matter remains unaltered.* This is the foundation fact in the science of chemistry.

18. Weighing. — In the study of the chemical action of the elements, and of the properties of the compounds they form, the determination of the weights of substances is an operation of supreme importance. The chemist must do it with the utmost precision. The instrument he employs is the *chemical balance* (Fig. 14).

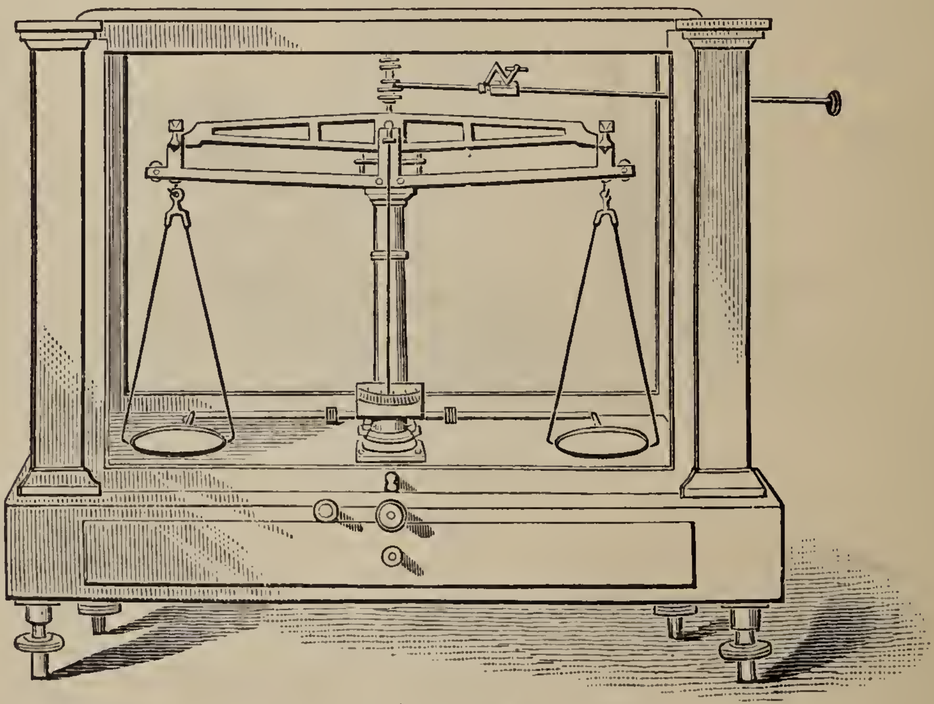


Fig. 14.

The workmanship in this instrument is of the finest kind, and the adjustments are made with the nicest skill. Heavy bodies are not to be weighed in the chemical balance, but the weight of light ones can be found with marvelous accuracy. With two hundred grams in each scale-pan, the instrument shown in the figure will turn by the addition of the twentieth of a milligram to either side. One-twentieth of a milligram is about one thirteen-hundredth of a grain.

The French or metric system of weights is very generally employed in chemistry. The student should be familiar with it. In this system

10 milligrams = 1 centigram.
10 centigrams = 1 decigram.
10 decigrams = 1 GRAM. = 15.43+grs.
10 grams = 1 decagram.
10 decagrams = 1 hectogram.
10 hectograms = 1 *kilogram* = 2.2+lbs. av.
10 kilograms = 1 myriagram.

19. Measuring. — But when the substance is a gas, weighing is a difficult operation, and the quantity is usually found by volume instead of by weight. For this purpose glass tubes of various sizes are accurately graduated (Fig. 11). The divisions along the tube show the volume inside in cubic inches or cubic centimeters, and fractions of these units.

Liquids also are more easily measured than weighed, and for many purposes their volumes are taken instead of their weight. In Fig. 15 some forms of graduated vessels for liquids are represented.

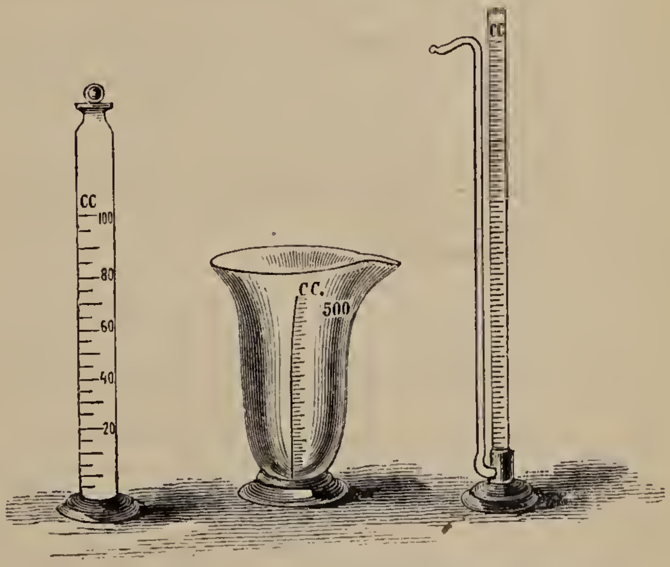


Fig. 15.

The cubic centimeter is the volume of one gram weight of pure water at 4° C. The number of cubic centimeters, therefore, represents also the weight of the water in grams.

As to any other liquid, if the number of cubic centimeters is multiplied by the specific gravity, the product will represent the weight of that liquid in grams.

The cubic centimeter is also called milliliter. It and the liter are the two denominations commonly used. The full table is as follows: —

10 milliliters = 1 centiliter.
10 centiliters = 1 deciliter.
10 deciliters = 1 LITER, = .22 gal.
10 liters = 1 decaliter.
10 decaliters = 1 hectoliter.
10 hectoliters = 1 kiloliter.
10 kiloliters = 1 myrioliter.

The operations of weighing and measuring require many precautions, and the exercise of much skill, to insure the nicest accuracy. We will not here enter upon the minute details.

REVIEW.

I. — SUMMARY OF PRINCIPLES.

20. By chemical attraction substances are made to combine. They disappear altogether, but new substances are seen in their stead.

Moreover the products of chemical action are often invisible gases.

Nevertheless all the matter of the original substance exists in the products of the action.

During all chemical changes the quantity of matter remains unaltered.

The quantity of matter in a body is found by taking its weight, or in the case of gases, and often of liquids also, by taking its volume.

Hence the operations of weighing and measuring are fundamental and most important operations in chemistry. The science is built upon the data given by them.

The French or metric system of weights and measures is most largely employed in chemistry.

The milligram, the gram, and the kilogram are the three most useful units of weight.

The cubic centimeter and the liter are the most useful units of volume. 1,000 milligrams = 1 gram; 1,000 grams = 1 kilogram; 1,000 cubic centimeters = 1 liter.

II. — EXERCISES.

By what means may we find out whether any gain or loss occurs in chemical change?

What are the results of experiment?

Of what importance is the balance in chemistry?

What system of weights is generally used?

Give the table. What is the value of the gram in English measure? Of the kilogram?

How many milligrams in one ounce Troy?

How many grains are equal to 500 m. g.?

How are the quantities of gases and liquids found?

What is a cubic centimeter?

What is the weight of 100 cc. of alcohol, whose specific gravity is .82?

What is the English value of the liter?

Give the table of volumes in the metric system.

SECTION IV.

ANALYSIS AND SYNTHESIS.

21. Analysis. — Any process by which a compound may be separated into its constituents, and its composition determined, is called ANALYSIS. If in the process only the names of the constituents are found, the analysis is *qualitative;* if their proportions are found, the analysis is *quantitative.*

By Electricity. — Many substances may be decomposed, and their composition found, by the action of electricity. We have already learned that such a process is called ELEC-

TROLYSIS. The most interesting example of electrolysis is that of water. Let us repeat the experiment, and study it more in detail.

Two platinum strips are inserted in a jar of water; and over them are inverted two long and slender tubes, previously filled with water. (See Fig. 16.) The wires of a galvanic battery are then inserted in the screw-cups, *s s*. Instantly a torrent of gas-bubbles rises in each tube, and will continue to do so until the tubes are filled. These two gases are the constituents of water. As the experiment goes on, it will be noticed that one tube is being filled faster than the other; indeed, being of equal size, one fills *twice as fast* as the other. If the gases are tested, that which is most rapidly set free is found to be hydrogen, the other oxygen.

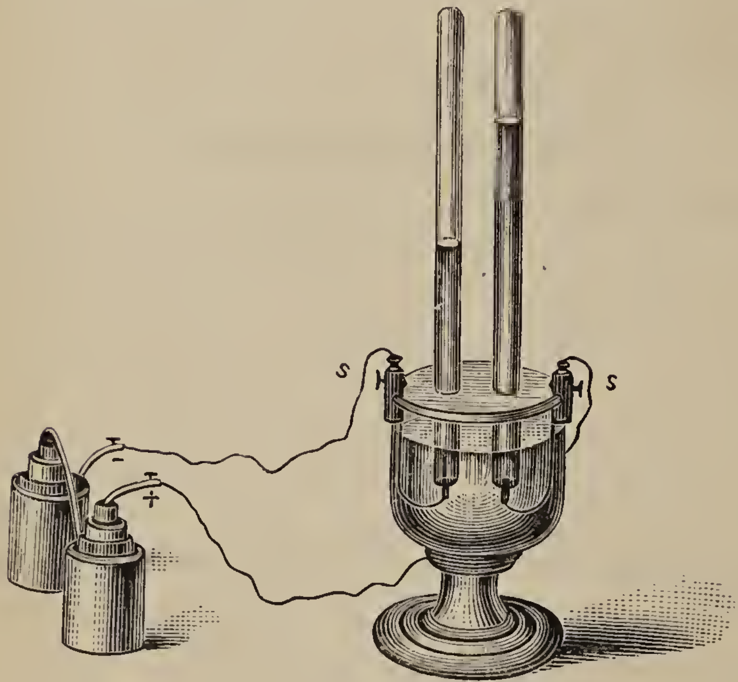


Fig. 16.

The experiment teaches that water is composed of hydrogen and oxygen, in the proportion *by volume* of two of hydrogen to one of oxygen.

Proportions by Weight. — This composition by volume is shown directly by the tubes; but, if we desire to know the proportions of the gases by weight, we must make a calculation.

The specific gravity of oxygen is 16. By this is meant that any volume of oxygen weighs just sixteen times as much as an *equal* volume of hydrogen. If, then, there were equal volumes of the two gases in water, their proportions by weight would be as 1 : 16. Since there are two volumes of hydrogen, the proportions are as 2 : 16. Eighteen parts of water (16 + 2) contain sixteen parts of oxygen and two parts of hydrogen. This being the ratio, it is easy to find just what part of any quantity of water is hydrogen, and what part oxygen. For example: one hundred of water must contain $\frac{100}{18}\times 2 = 11.11$ of hydrogen, and $\frac{100}{18}\times 16 = 88.89$ of oxygen.

Percentage Composition. — The composition of a compound is usually given by the hundred, and it is then called the PERCENTAGE COMPOSITION. The results in the case of water are written in this way: —

WATER.

Hydrogen	11.11
Oxygen	88.89
	100.00

By the Prism. — When the light of a burning substance is passed through a prism, it is decomposed; and the appearance of the spectrum depends upon the nature of the burning substance. Hence the constituents of a compound may be told by the appearance of its spectrum. This method of detecting the presence of substances is called SPECTRUM ANALYSIS.

For example: when the spectrum of burning sodium is viewed through a telescope, a bright yellow line, of surprising brilliancy, is seen in the yellow part of the spectrum. Or, if potassium is burned, a single crimson line and another of blue, both of great beauty, will be seen in opposite ends of the spectrum. Nor will the appearance of either set be changed by the presence of the other; for, if a mixture of sodium and potassium is burned, the observer will see both sets in the same spectrum, their place, size, and brightness the same as when each was formed alone. The spectra of many elements have been carefully studied; they can be easily recognized by one who has previously made their acquaintance.

By Chemical Action. — The methods of analysis just described are limited in application. The general method is by the chemical action of bodies upon each other. More than a hint of this method would be out of place here. But suppose, for example, that a solution of unknown substances in water is to be analyzed. A few drops of hydrochloric acid may be added. If a white, solid substance is formed, it shows the presence of either silver, mercury, or lead. To find out which of these metals is present, a little ammonia is added: if the solid disappears again, the metal is silver; if it simply turns black, mercury is present; but if it remains unchanged, the metal is lead. If the acid fails to produce the white solid, or PRECIPITATE, as any solid formed in this way from solution, by the action of chemicals, is called, then these three metals are counted absent, and some other agent is used by which another group of metals may be detected.

So far the analysis is only *qualitative.* To obtain *quantitative* results, the balance must be used.

The original substance is first weighed, and finally the precipitate, clean and pure and dry. These two weights enable the chemist to calculate the weight of the metal in one hundred parts of the original compound.

22. Synthesis. — The composition of a compound may be found by causing its constituents to combine, and noticing the proportions required. This process is called SYNTHESIS. The synthesis of water may illustrate: it may be made in an apparatus called a EUDIOMETER.

This instrument is a glass tube, open at one end, and carefully graduated, having two metallic wires passing through the glass near the closed end, and almost touching each other inside. To use it, measured quantities of hydrogen and oxygen, say fifty volumes of each, are put into it. The tube should still be much less than full of the mixed gases, and should be firmly held with its open end in the cistern. (Fig. 17.) If now electricity from a Leyden-jar or a Ruhmkorff-coil is passed through the mixture, the two gases combine with violent explosion; water will be formed by their union; and water or mercury from the cistern will rise into the tube to fill the space they occupied. The quantity of gas left will be found to be twenty-five volumes, and, if tested, will be found to be oxygen.

It is evident that fifty volumes of hydrogen have taken twenty-five volumes of oxygen to form water; and the experiment therefore teaches that water is composed of hydrogen and oxygen in the proportions by volume of two of hydrogen to one of oxygen.

By analysis or by synthesis, the chemist is able to find out just what all kinds of matter are composed of, and even the exact proportions by weight or by volume, in which their constituents are united.

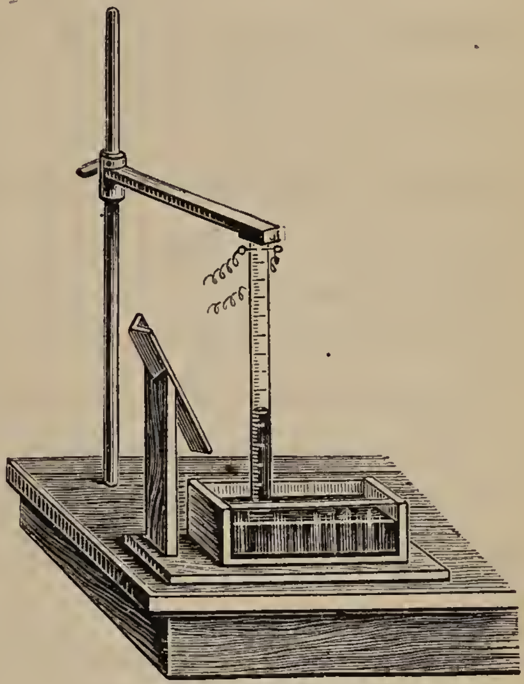


Fig. 17.

REVIEW.

I.—SUMMARY OF PRINCIPLES.

23. The composition of a compound may be found by analysis or by synthesis.

Analysis is any process by which a compound is separated into constituents, and its composition determined.

Synthesis is any process by which the constituents of a compound may be made to combine and produce it.

Qualitative analysis is an analysis in which only the names of the constituents are found.

Quantitative analysis is an analysis in which the proportions by volume or by weight of the constituents are determined.

By the percentage composition of a compound, we mean the proportional parts of all its constituents in one hundred parts of the compound.

In one hundred ounces, grams, or other units, by weight, of water, there are 88.89 ounces, grams, or other units of oxygen, and 11.11 of the same units of hydrogen.

By volume, however, water consists of two volumes hydrogen to one volume oxygen.

A precipitate is any solid substance produced by the chemical action of other substances in solution.

In the solution of a compound of either silver, mercury, or lead, hydrochloric acid will produce a white precipitate, and in these only.

Hydrochloric acid is said, therefore, to be a Test for the presence of these metals.

Conversely, any compound of these metals is a test for hydrochloric acid. If a solution contains the slightest quantity of this acid, a drop or two of silver nitrate will turn it white.

II. — EXERCISES.

Define analysis. Define qualitative analysis. Define quantitative analysis. Describe the electrolysis of water. What does this analysis of water show?

Define specific gravity. What is the specific gravity of oxygen? Knowing that water consists of one volume oxygen and two volumes hydrogen, how would you calculate the proportions by weight? Define percentage composition. What is the percentage composition of water?

Define spectrum analysis. What is a more general method of analysis? Define precipitate. How would you tell which of the metals, silver, mercury, or lead, is contained in a substance? How would you find out how much the original substance contained?

Define synthesis. Describe the eudiometer. What does the synthesis of water teach?

SECTION V.

THE LAWS OF COMBINATION.

24. The Law of Constant Proportions. — The first great law discovered by the use of the balance and the measuring tube is that of the invariable proportions of the constituents in any chemical compound. Whether we find the composition of water by analysis or by synthesis; whether we operate upon water from the clouds, from the spring, or from any other source, — we find it to consist of the same elements, and in the same proportions by weight or volume. The same may be said of every other known compound. The law states, that —

The same compound is always made up of the same constituents, combined in the same definite proportions. This is known as the law of "definite" or of "constant proportions."

25. Composition of Hydrochloric Acid. — We may study the combination of hydrogen and chlorine by experiment, and determine their constant proportions in hydrochloric acid.

The Experiment. — Let us provide a small cistern, partly full of a strong solution of common salt. Let us fill two graduated tubes (Fig. 18) also with this brine, and invert them in the cistern. Into one of these we may pass up 20 cubic centimeters of hydrogen and 40 cubic centimeters of chlorine. Into the other we will pass 20 cubic centimeters of chlorine and 40 cubic centimeters of hydrogen. All this having been done quickly, we at once place the whole apparatus for five minutes in direct sunlight. Now, chlorine is a greenish-yellow gas, which will extinguish the flame of a match; while hydrogen is a colorless gas, which takes fire instantly on contact with the flame.

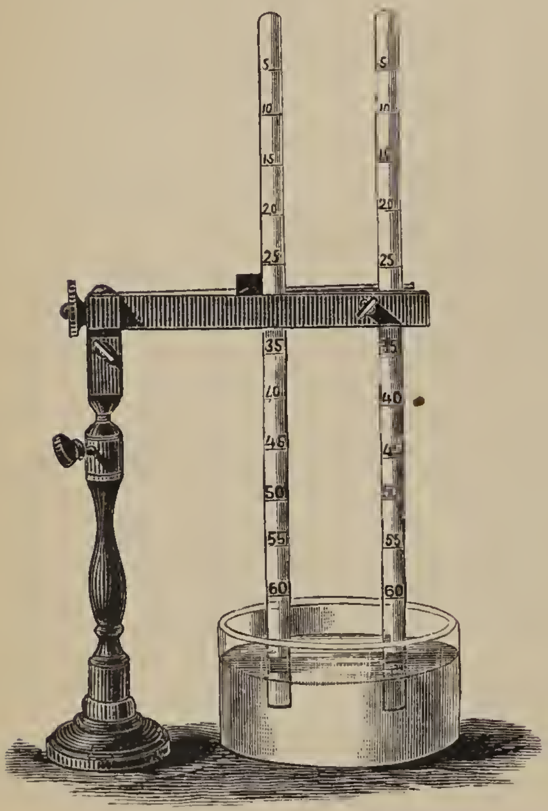


Fig. 18.

Finally it will be seen that only 20 cubic centimeters of gas remain in each tube. That in the first tube is greenish-yellow, and if tested will extinguish a flame: it is chlorine. That in the second tube is colorless, and if tested will take fire: it is hydrogen.

The Result by Volume. — The 20 cubic centimeters of hydrogen in the first, combined with 20 cubic centimeters of chlorine. The 20 cubic centimeters of chlorine in the second, combined with 20 cubic centimeters of hydrogen. The *proportions* by volume, in both cases, are 1 : 1. We can not make these two gases combine in any other proportions. The composition of hydrochloric acid is, invariably, *hydrogen and chlorine in equal volumes.*

The Result by Weight. — Now, the weights of equal volumes of these gases are as 1 to 35.5: hence the composition of hydrochloric acid by weight is 1 of hydrogen to 35.5 of chlorine. These proportions are invariable.

Hydrochloric acid *always* consists of these same elements, combined in these same proportions.

26. Combining Weights. — So definite and constant are the weights of the elements in combination, that a number is attached to each, which is called its COMBINING WEIGHT. It represents the *smallest relative proportion, by weight*, in which the element can unite with others to form compounds.

For example, it is found that any given weight of hydrogen must have at least

35.5 times as much chlorine to form a compound.
80 " " bromine " "
127 " " iodine " "

These numbers are the combining weights of chlorine, bromine, and iodine, *because* they are the weights of the smallest quantities of these elements which can combine with a unit weight of hydrogen.

Again, if we examine some other compounds of chlorine, we find that

$$35.5 \text{ of chlorine must have} \begin{cases} 39 \text{ of potassium,} \\ 23 \text{ of sodium,} \\ 108 \text{ of silver.} \end{cases}$$

These numbers are the combining weights of potassium, sodium, and silver, because they are the smallest relative quantities of these elements which can enter into combination.

The Unit of Combining Weights. — *Hydrogen* is the *standard* of comparison for combining weights. Any given weight of hydrogen may be regarded as the unit. 35.5 times that is the smallest weight of chlorine, and 108 times that weight is the smallest weight of silver, which can combine without excess of either. Thus every element has a numerical value in chemical change. These combining weights are to be found in the table of elements on p. 53.

27. The Law of Multiple Proportions. — Each element has but one combining weight, and yet the same two elements may form more than a single compound.

Some Facts. — Of oxygen and nitrogen, for example, no less than five compounds are known to exist. Examine their composition, and another law will be brought to light.

	(1)	(2)	(3)	(4)	(5)
Nitrogen	28	14	28	14	28
Oxygen	16	16	48	32	80

Now 16 is the combining weight of oxygen; and we find in this series that when the quantity of oxygen is not 16 it is some *multiple* of 16.

Again, 14 is the combining weight of nitrogen; and we also find, that when the quantity of nitrogen is not 14 it is a *multiple* of 14.

The Law. — This illustrates the law, that *if one element unites with another, in more than one proportion, these proportions will all be multiples of its combining weight.*

If we attempt to combine these elements in any other proportions, the result will be that all of one of them will be used up to form one or more of these five compounds, while the excess of the other will be left uncombined.

28. The Combining Weight of a Compound. — A compound must contain the same quantity of matter as its constituents: this follows from the indestructibility of matter. Water contains two parts of hydrogen and sixteen parts of oxygen. Its numerical value must therefore be $16+2=18$. Moreover, water combines with many other substances, and its combining weight is found to be 18.

This illustrates the law, that *the combining weight of a compound is the sum of the combining proportions of its constituents.*

Combining Proportions. — It is important to notice that the term "combining proportion," just now used, does not mean combining weight in all cases. The combining weight of hydrogen is 1: the quantity which combines with 16 of oxygen, however, is 2; and it is this which enters into the combining weight of water. We have applied the term "combining weight" to the *smallest* proportion of any substance which may enter into combination: we apply the term "combining proportion" to the *relative weight actually* *existing in the compound.* Thus the combining weight of oxygen is 16: its combining proportion is sometimes 16, and sometimes 32 or 48 or 80, or some other multiple of 16.

29. Dalton's Theory. — John Dalton, in the year 1808, declared his belief that the combining weights of the elements are simply the relative weights of the very smallest particles of them which can enter into chemical action.

We may divide a body into a multitude of pieces. Indeed, we may grind it to the finest dust, and it may seem that there is no end to its divisibility. But Dalton's theory assumes that there is a point beyond which we cannot carry the division, and that when we reach this point we have minute particles whose weights are not alike in different elements. These ultimate particles of the elements are called ATOMS. The chemist defines the atom to be *the smallest portion of an element which can take part in a chemical action.*

According to this theory, an atom of oxygen weighs sixteen times as much as an atom of hydrogen; and, in the same way, the combining weights of chlorine, sodium, silver, and of other elements, tell us how many times the atoms of these bodies are heavier than an atom of hydrogen. Hence the terms "combining weights" and "atomic weights" mean the same thing.

Dalton's theory assumes also that chemical action is an action among *atoms*, that chemical attraction acts upon the *atoms* of elements, and brings them together to produce the compound.

This theory is known as the ATOMIC THEORY.

Reason for Accepting this Theory. — The atomic theory is generally believed; for, while it is impossible to get at these atoms to weigh them separately, to handle them, or to see them, still it seems impossible to discover any good reason for the laws of combination, unless they do exist. On the other hand, if we admit the existence of atoms, it is perfectly clear that the weights of the elements in a compound must be definite and constant. That *theory* in science is accepted, which gives a clear and complete explanation of all the *facts* which it embraces. This the atomic theory seems to do, and this is the evidence of its truth.

30. The Molecule. — When hydrogen and chlorine combine to form hydrochloric acid, each *atom* of chlorine seizes an *atom* of hydrogen. In a tube full of the compound, we have simply a multitude of these pairs of atoms. The smallest quantity of hydrochloric acid which can be, must, according to this theory, consist of these two atoms. So, common salt is a compound of chlorine and sodium, and the smallest piece of it which can exist as salt, must contain two atoms, one of chlorine and one of sodium. These ultimate particles of a compound are called MOLECULES.

It is believed that the atoms of the elements, also, when they are in the free state, that is, not in combination, are linked together in groups, generally of two: and so the chemist defines the molecule to be *the smallest particle of an element or compound which can exist in the free state.*

31. Combination by Volume. — We have learned already that it is more easy to measure gases than to weigh them. But the question arises, How much by volume of the gases must be taken in order to produce compounds without leaving any part of either constituent uncombined? This question is answered by the law of Gay Lussac, discovered by him in 1808. It states that "*the weights of the combining volumes of the gaseous elements bear a simple relation to their atomic weights.*"

To find this Relation. — This simple relation is found by the balance. By weighing equal volumes of these elements it is found that —

1 liter of oxygen weighs 16 times as much as a liter of hydrogen.

1 liter of nitrogen weighs 14 times as much as a liter of hydrogen.

1 liter of chlorine weighs 35.5 times as much as a liter of hydrogen.

Here we see that *these weights of equal volumes are the same as the atomic weights of the elements.* This is true of nearly all the elements when in the form of gas or vapor.

If, then, we *bring such elements together in equal volumes, or in some multiple of the equal volume*, it will be the same in effect as bringing them together in the proportions of their combining weights or in some multiple of their combining weights: no excess of either can remain after combination has taken place.

The Volume of the Compound. — Experiments show further that the simple relation extends to the volume of the compound also. The following are some of the results of exact experiment: —

1 vol. chlorine and 1 vol. hydrogen form 2 vols. hydrochloric acid.
1 vol. bromine and 1 vol. hydrogen form 2 vols. hydrobromic acid.
1 vol. iodine and 1 vol. hydrogen form 2 vols. hydriodic acid.
1 vol. oxygen and 2 vols. hydrogen form 2 vols. steam.
1 vol. nitrogen and 3 vols. hydrogen form 2 vols. ammonia.

We here find that two volumes of the compound are formed in every case; and we discover the curious fact, that, whether the sum of the constituent volumes be 2 or 3 or 4, the volume of the compound is only 2. It is found in other cases also, that, *whatever the number of volumes which enter into combination, the volume of the resulting compound is* 2.

Avogadro's Law. — The simplest explanation we can find of these facts is based upon two assumptions, viz.: —

Every molecule of each of these elementary gases contains two atoms.

Equal volumes of all gases, simple and compound, contain an equal number of molecules.

Let us study the case of steam in the above table, remembering that, according to the atomic theory, two atoms of hydrogen and one atom of oxygen form one molecule of steam.

If each molecule of oxygen contains two atoms of oxygen, and only one is needed in a molecule of steam, there will be just twice as many molecules of steam produced as there are of oxygen required to form it. And then if *equal volumes* of steam and oxygen contain the *same number of molecules*, twice as many molecules of steam must fill just twice the volume of the oxygen. Hence the combination of *one* volume of oxygen with sufficient hydrogen *must* produce *two* volumes of steam. The same kind of reasoning explains all other cases, and hence it is believed that the assumptions on which the explanations are founded are true.

The supposition that "*Equal volumes of all gases, simple and compound, contain equal numbers of molecules*," was first made by Avogadro in 1811. It is known as AVOGADRO'S LAW, and is generally accepted, both in chemistry and in physics.

REVIEW.

I.—SUMMARY OF PRINCIPLES.

32. Chemical changes occur in obedience to certain laws, known as the laws of combination.

The "Law of Constant Proportions" states that the same compound is always made up of the same constituents, combined in the same definite proportions by weight.

Water always consists of hydrogen and oxygen in the proportions, by weight, of two of hydrogen and sixteen of oxygen.

Hydrochloric acid always consists of hydrogen and chlorine, in the proportions, by weight, of one of hydrogen to 35.5 of chlorine.

The combining weights of substances are the smallest *relative* proportions, by weight, in which they enter into combination.

Hydrogen is the standard of combining weights; any given weight of this gas may be considered the UNIT.

The "Law of Multiple Proportions" states, that if one substance unites with another, in more than one proportion, these proportions are all multiples of its combining weight.

The combining weight of a compound is the sum of the combining proportions of its constituents.

The combining weight of water is eighteen, of hydrochloric acid 36.5.

These laws of combination have been established by experiment. Our knowledge of them is quite independent of any theory. Still the theory proposed by Dalton does explain them.

This "Atomic Theory" assumes that every element consists of ATOMS, or minute particles, the smallest that can take part in any chemical action. It assumes further, that any compound is formed by the union of the atoms of its elements, and that the combining weights of the elements are the relative weights of their atoms.

A MOLECULE is the smallest particle of an element, or of a compound, which can exist in the free state.

Chemical change consists in the breaking up of old molecules, and the re-arrangement of their atoms to form new ones.

A molecule of hydrochloric acid contains one atom of hydrogen and one of chlorine. A molecule of water contains two atoms of hydrogen and one of oxygen. A molecule of ammonia contains four atoms, three of hydrogen and one of nitrogen.

A molecule of an element generally contains two atoms.

Combining volumes are the relative proportions by volume in which gaseous or volatile substances unite.

The weights of the combining volumes of the elementary gases bear a simple ratio to their atomic weights. This is Gay Lussac's Law.

Equal volumes of all gases, elementary or compound, at the same temperature and pressure, contain an equal number of molecules. This is Avogadro's Law.

According to this law, the molecules of all gases or vapors are of the same size. That is, the *molecular volumes* of all gases, elementary or compound, are equal: the numerical value is two.

Combining volume, combining or atomic weight, specific gravity, and molecular weight, are all referred to hydrogen as a standard of comparison.

Let us represent the names of a few elements by their initial letters, and under each put its numerical values. We shall have

	H	O	N	P	Hg.	
	1 —	1 —	1 —	$\frac{1}{2}$ —	2	= combining vols.
Weights of these =	1 —	16 —	14 —	31 —	200	= atomic weights.
Weights of *equal* vols. =	1 —	16 —	14 —	62 —	100	= specific gravities.
Mol. of H = 2 atoms.	2 —	32 —	28 —	124 —	200	= molecular weights.
	2 —	2 —	2 —	2 —	2	= molecular vols.

From which we discover that—

Atomic weight = specific gravity × combining volume:

Molecular weight = 2 × specific gravity.

II.—EXERCISES.

What is the composition of water? What is the composition of hydrochloric acid, by volume? By weight? Is it always the same? Is the same thing true of other substances? State the law of constant proportions.

Define combining weight. What is the unit of measure for combining weights? The combining weight of chlorine is 35.5: what is meant by this statement?

Ten grams of hydrogen will require how much chlorine to convert it all into hydrochloric acid?

What would be the result if ten grams hydrogen, mixed with 360 grams chlorine, are exposed to sunlight?

If 100 cubic centimeters of oxygen are mixed with 100 cubic centimeters of hydrogen, what will remain after passing the electric spark through the mixture?

If 100 grams of oxygen are to be converted into water, how much hydrogen must be used?

How many compounds of oxygen and nitrogen are known? What law do their compositions illustrate? State this "law of multiple proportions."

What is the combining weight of water? State the law which this illustrates.

Define atom. Define molecule. What is the atomic theory? On what ground is the atomic theory accepted?

State Gay Lussac's Law. Compare the weights of one liter of each, — oxygen, nitrogen, and hydrogen. What relation between these weights and the atomic weights of these elements? Define combining volume. What is the combining volume of a gaseous or volatile compound?

State Avogadro's Law.

How much oxygen is required to unite with 10 liters of hydrogen?

How many liters of water-vapor will be formed by the union of 2,000 cubic centimeters of hydrogen and 1,000 cubic centimeters of oxygen?

Are the combining volumes of the *elements* the relative volumes of their atoms. or of their molecules?

SECTION VI.

CHEMICAL NOTATION AND NOMENCLATURE.

33. Symbols. — Instead of writing the names of elements in full, chemists have agreed to use a set of symbols to represent them. These symbols are the capital initial letters of the names (often of the Latin names), or, in case two elements have the same initial, this capital with a small letter added. Thus O stands for oxygen; N for nitrogen; H for hydrogen. Of carbon, cobalt, and copper (Latin *cuprum*), the symbols are C, Co, and Cu. The symbol not only represents the name of the element: it also represents just *one atom* of it. The symbols of all the elements are given in the table on p. 53.

34. Formulas of Compounds. — Instead of writing the names of compound bodies in full, a system of formulas has been adopted. The formula of a compound is made up of the symbols of its constituents, with figures placed a little below to show the number of combining weights. Instead of writing the statement that water consists of hydrogen and oxygen, two combining weights of the one to one of the other, we may with less trouble write the formula, $H_2 O$, which teaches the same thing. That nitric acid consists of one part of hydrogen, one of nitrogen, and three of oxygen, is shown by the simple expression, $H N O_3$. When no figure is used, one is understood. The formula not only represents the name: it also represents just *one molecule* of the compound.

What they teach. — These formulas contain much valuable information about the compounds they represent.

They teach: —

1st, The names of the constituents, by the letters they contain.

2d, The number of combining weights of each, by the figures they contain.

3d, The number of combining volumes of gaseous or volatile constituents, also by the figures they contain.

4th, The combining *proportion* of each, for this is equal to the combining weight multiplied by the figures used.

5th, The number of atoms of each element in the molecule of the compound.

Illustration. — The formula of potassium chlorate, for example, is $K Cl O_3$, and it answers all of the following questions: —

1st, What are the constituents of this compound? Potassium, chlorine, and oxygen.

2d, How many combining weights of each? One of potassium, one of chlorine, three of oxygen.

3d, How many volumes of each? One of potassium, one of chlorine, three of oxygen.

4th, What combining *proportion* of each? K = 39.1, Cl = 35.5, O = 48.

5th, How many atoms in one of its molecules? One of potassium, one of chlorine, and three of oxygen.

CHEMICAL NOMENCLATURE.

35. The Nomenclature. — Not until the year 1787 was any attempt made to reduce the language of chemistry to a system. But the number of compounds to be described increased to such an extent that even the strongest memory could not hope to keep their meaningless names. To avoid this difficulty, a system was proposed by Lavoisier, by which the name of a substance should indicate its composition. The simplicity of the system, and the accuracy with which it expressed chemical theories, secured its universal adoption; but the theories themselves having now, in great part, been rejected, new names have been introduced, and it can not yet be said that the system is permanent. Nevertheless, certain specific rules are observed. The object is to make the name of a compound show what elements it is composed of, and, as nearly as may be, also the relative proportions in which they are present.

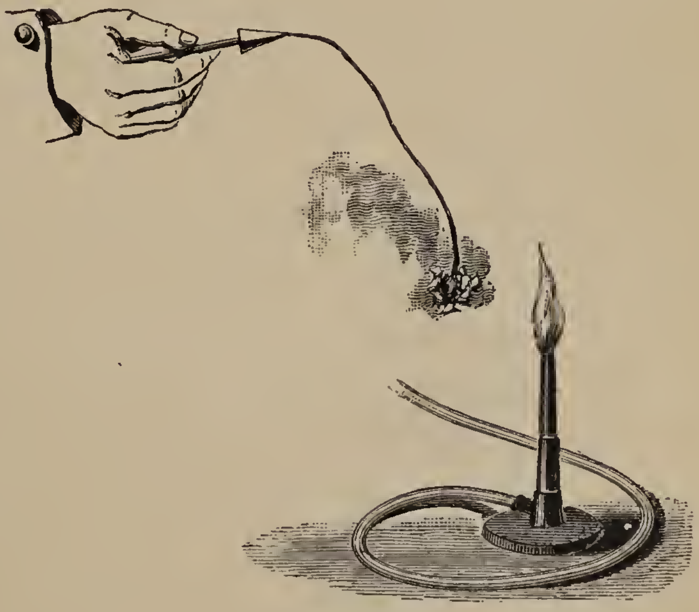


Fig. 19.

36. Binary Compounds. — When only two elements unite, they form what is called a Binary Compound. Among the most numerous and important are the binary compounds of oxygen.

For example, let a piece of magnesium, which can be had in the form of wire or ribbon, be held in a pair of forceps, while the other end is heated in the flame of a lamp. The metal first becomes red-hot, and then takes fire (Fig. 19). It burns with a vivid light, and throws off large volumes of white smoke. This white smoke is the compound of oxygen with magnesium. It is called Magnesium Oxide.

In like manner all the binary compounds of oxygen are called Oxides, while the other part of the name signifies the element with which the oxygen is united.

Prefixes. — It often happens that oxygen forms several binaries with the same other element, by combining in multiple proportions. This is true with the metal manganese. Three distinct oxides of manganese are known. Their formulas are: —

$Mn\,O$ . . .	1 atom of O in the molecule.
$Mn\,O_2$ . . .	2 atoms " " "
$Mn_2\,O_3$ . . .	3 " " " "

To show that one atom of oxygen is present in the molecule, the first is called manganese *mon*oxide. The two atoms of oxygen in the second are indicated by the name manganese *di*oxide; while the proportions three to two atoms in the molecule of the third are shown by the name manganese *sesqui*oxide.

In general, we may say that *the number of combining weights, or of atoms, in the molecule, is shown by a prefix to that part of the name which represents the element.*

The prefix *mon* may be omitted; the term *oxide* alone indicating the monoxide. And often the prefix *per* is given in the name of the oxide which has the largest proportion of oxygen, without regard to the number of atoms.

Terminations. — In case the same element forms *two* oxides, it is customary to use the terminations *ic* and *ous* in naming them. Thus sulphur and oxygen form

$S\,O_3$	sulphur*ic* oxide.
$S\,O_2$	sulphur*ous* oxide.

And the two compounds of mercury and oxygen are called mercuric oxide, $Hg\,O$, and mercurous oxide, $Hg_2\,O$.

Still these same substances may be named by the use of prefixes; and then, in the case of sulphur, we have, for the first, sulphur trioxide, and, for the second, sulphur dioxide. In very many other cases, also, the same compound may have two or more different names.

Other Binary Compounds. — The very same rules for naming these oxygen compounds apply to the naming of the binaries of many other elements, such as chlorine and sulphur. Thus, for example: —

Compounds of	are called	such as	formula
Chlorine,	chlorides,	sodium chloride,	$Na\,Cl$.
Bromine,	bromides,	potassium bromide,	$K\,Br$.
Iodine,	iodides,	lead iodide,	$Pb\,I_2$.
Sulphur,	sulphides,	iron sulphide,	$Fe\,S$.

The same prefixes, and the terminations *ic* and *ous*, are used for exactly the same purposes in naming these as in naming the oxides.

37. Acids, and their Composition. — If we place a piece of dry phosphorus upon a little support, resting on a dry plate, touch it with a hot wire, and quickly invert over it a dry bell-jar (Fig. 20), the burning phosphorus will throw off large volumes of milk-white vapors, much of which will finally condense into snow-like flakes upon the sides of the glass.

In this experiment the oxygen of the air, and the phosphorus, unite to form the new white substance, — the phosphorous pentoxide, whose formula is $P_2\,O_5$ — a binary compound.

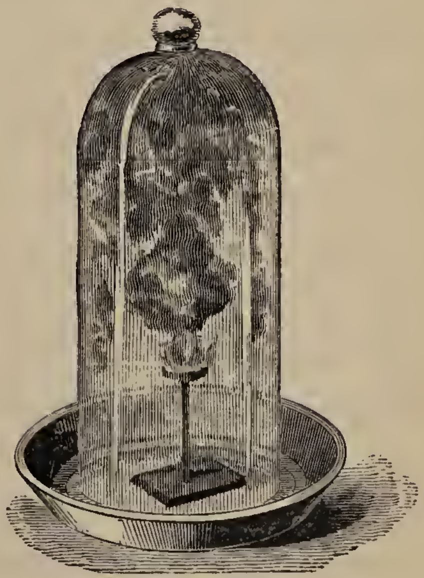


Fig. 20.

Next, let the jar be lifted, and a little water poured into it.

The water combines with the white oxide, which disappears with a hissing sound. The fluid becomes *intensely sour* to the taste; and if a little of this solution be added to a solution of blue litmus, the *blue color* will be instantly *changed to red.*

This power to redden blue litmus, or other vegetable colors, is characteristic of an all-important class of compounds called Acids. The water and the oxide combined to form Phosphoric Acid.

Composition of Acids. — The composition of this phosphoric acid is shown by its formula: $H_3 P O_4$. Its molecule contains three atoms of hydrogen, one of phosphorus, and four of oxygen. Below are the formulas of several other acids for inspection: —

Sulphurous acid,	$H_2 S O_3$.	Hydrochloric acid,	$H Cl$.
Sulphuric acid,	$H_2 S O_4$.	Hydrobromic acid,	$H Br$.
Nitric acid,	$H N O_3$.	Hydriodic acid,	$H I$.
Carbonic acid,	$H_2 C O_3$.	Hydrosulphuric acid,	$H_2 S$.

A glance at these formulas reveals the fact that *hydrogen is a constituent in every one* of these acids.

From all these facts we may gather this description: viz., *an acid is a hydrogen compound, which is generally sour to the taste, and which will redden vegetable colors.*

They are of Two Classes. — Another glance at the above formulas reveals the fact that some acids contain oxygen, and that others do not. Those which contain oxygen are called Oxygen Acids, or Oxacids; while those which do not are called Hydracids. The oxygen acids comprise far the larger number of this class of bodies.

38. Names of the Oxygen Acids. — The terminations *ic* and *ous* are used in the names of the oxygen acids, signifying the same thing as when used in the names of binaries. That is, the *ic* indicates the larger proportion of oxygen in the molecule. Thus, for example, $H N O_3$ is nitric acid, while $H\,N\,O_2$ is nitrous acid. Two acids of sulphur are distinguished in the same way.

The prefix *hypo* is also used, and always means *less*. *Hypo*nitrous acid, for example, is one which contains a still smaller proportion of oxygen than the nitrous acid: its formula is $H\,N\,O$. The prefix *per* is also sometimes used to form the name of the acid which contains the largest proportion of oxygen.

Names of the Hydracids. — The names of the hydracids also end in *ic*, but they are especially characterized by the prefix *hydro*. Find this prefix in the names of acids in the list already given. These four are the most important acids of this class, and the name shows the constituents in each.

The hydrogen acids are really binary compounds, and we may name them as such if we choose. Instead of hydrochloric acid, we may speak of $H\,Cl$ as hydric chloride.

39. Salts. — Notice the following experiment. Into a bottle put a few clippings of zinc, and upon them pour a quantity of hydrochloric acid. A vigorous boiling quickly begins; hydrogen gas escapes, and may be set on fire with a match; the zinc slowly disappears, and finally a quiet liquid remains, which would be quite clear but for some black residue from the impurities of the zinc employed. By evaporating the clear liquid we may obtain a white solid: this proves to be zinc chloride.

The explanation is this: The zinc has decomposed the acid, taking the place of the hydrogen which was in combination with the chlorine. Before the action we had hydric chloride and zinc: after the action we have zinc chloride and hydrogen.

Atoms of zinc take the place of *atoms* of hydrogen in the *molecules* of the acid. This zinc chloride is not only a binary compound: it is also called a SALT. If sodium takes the place of hydrogen in the same acid, the *salt*, sodium chloride (*common salt*), is formed.

Again: if the two combining weights of hydrogen in sulphuric acid are both replaced by sodium, a salt, sodium sulphate, will be formed; or, if only one of them is replaced by sodium, the remaining compound is still a salt. From these illustrations we gather this description: *A salt is a compound formed by substituting a metal for either the whole or a part of the hydrogen in an acid.*

40. Names of Salts. — The salts formed from hydracids are named by the method already given for neutral binary compounds. These salts were formerly called *haloid salts.*

Names of Oxy-Salts. — Of salts formed from oxacids, the name is made by changing the ending of the name of the acid from which they are derived, from *ic* to *ate*, or from *ous* to *ite.*

A few examples will make this method clear.

Salts from	are	such as
Sulphuric acid,	sulphates,	sodium sulphate.
Sulphurous acid,	sulphites,	sodium sulphite.
Nitric acid,	nitrates,	potassium nitrate.
Nitrous acid,	nitrites,	potassium nitrite.
Chloric acid,	chlorates,	potassium chlorate.
Chlorous acid,	chlorites,	sodium chlorite.
Hypochlorous acid,	hypochlorites,	sodium hypochlorite.

In case only a part of the hydrogen of the acid is displaced, the presence of the remainder is indicated in the name by the prefix *hydro.* When, for example, all the hydrogen of sulphuric acid $H_2 S O_4$ is replaced by sodium, the salt $Na_2 S O_4$ is called *sodium sulphate;* but, if only one of the two combining weights of hydrogen is displaced, the presence of the metal and of the remaining hydrogen is shown by the name *hydro-sodium sulphate* $H Na S O_4$.

In the same way potassium and sulphuric acid may form either potassium sulphate or hydro-potassium sulphate.

The acids are sometimes called HYDROGEN SALTS. Just as $K N O_3$ is called potassium nitrate, so $H N O_3$ may be called hydrogen nitrate instead of nitric acid.

41. Bases. — Let a small piece of the metal potassium be dropped upon the surface of water: it instantly takes fire (Fig. 21), runs swiftly about over the surface of the water, until finally it is all wasted away. If now we prepare some reddened litmus, by adding to blue litmus a few drops of acid, and afterward add some of the water on which the potassium burned, we find that the blue color of the litmus is quickly restored.

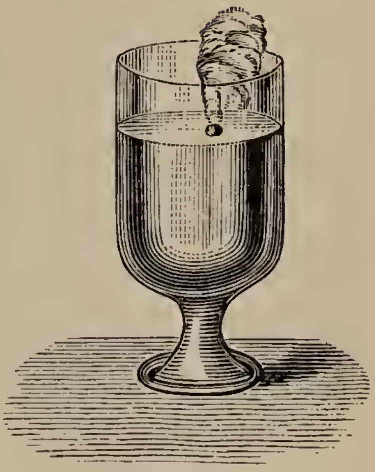


Fig. 21.

The power to neutralize an acid, shown by the restoration of the blue color to reddened litmus, is characteristic of another large and important class of compounds. They are called BASES.

Hydrates. — This whole class of bodies also receives the name of HYDRATES. A hydrate is a compound of hydrogen and oxygen, with a third element, usually a metal. It is often caustic to the taste, and restores the blue color of reddened litmus. A hydrate receives the name of the metal which it contains. We have, accordingly, *sodium* hydrate, $Na H O$; *potassium* hydrate, $K H O$; silver hydrate, $Ag H O$; and lead hydrate, $Pb H_2 O_2$, or, as it is oftener written, $Pb (HO)_2$.

REVIEW.

I. — SUMMARY OF PRINCIPLES.

42. The symbol of an element is the capital initial of its name, sometimes accompanied by a small letter; as, C for carbon, and Co for cobalt.

The symbol represents not only the *name*, but also a single *atom*, of the element.

The formula of a compound consists of the symbols of its elements, with small figures attached, to show the number of combining weights of each.

The formula represents a single molecule of the compound.

Among the many classes of compounds are binaries, acids, salts, and hydrates.

The name of a binary compound is known by its terminating in *ide*. Besides this its name combines the names of both constituents, and prefixes are used to show how many atoms of the constituents are present.

Acids are hydrogen compounds, which are sour, which are able to redden vegetable colors, and which will exchange hydrogen for a metal in chemical action.

Acids are called oxacids if they contain oxygen, and hydracids if they do not.

The oxacids are named after the element which is combined with their hydrogen and oxygen, the terminations *ic* and *ous* being used to indicate larger and smaller proportions of oxygen. The prefix *hypo* is also used to indicate a smaller proportion of oxygen, and *per* to show a larger proportion of the same element.

The prefix *hydro* characterizes the name of a hydracid. These acids may also be named as if they were simply binary compounds.

Salts are described as substances which may be produced by substituting atoms of metal for atoms of hydrogen in an acid.

They are named after the acids from which they are derived, by changing the ending of the name of the acid from *ic* to *ate*, and from *ous* to *ite*. The name of the metal is then prefixed.

Acids are sometimes regarded as hydrogen salts, and are named as such. Instead of sulphuric acid, we may use the name hydrogen sulphate.

Hydrates or Bases are described as compounds of hydrogen and oxygen with a metal having the power to neutralize an acid. They are usually caustic to the taste, and restore the blue color of reddened litmus.

The hydrates are named after the metals which they contain.

II. — EXERCISES.

Define symbol. Formula.

How much of an element does a symbol represent?

How much of a compound does a formula represent?

How many atoms in a molecule of alcohol, its formula being $C_2 H_6 O$?

What is the symbol of Hydrogen? Oxygen? Nitrogen?

How would you represent a molecule of each?

Define binary compound.

What are oxides? How is any oxide named?

What is a Monoxide? A Dioxide? A Trioxide? A Sesquioxide? A Peroxide?

For what purpose are the terminations *ic* and *ous* used?

What are Iodides? Chlorides? Bromides? Sulphides?

Mercury combines with oxygen; what shall we call the compound? *Ans.* Mercuric oxide.

But there is another oxide of this metal, containing a less *proportion* of oxygen; what shall it be named?

What are the elements of cupric oxide?

What are the elements of cuprous oxide?

What are the elements and their proportions in chromium trioxide?

Name the compound of potassium and iodine.

Name the compound of lead and sulphur.

Name the compound of lead and iodine.

Name the compound of copper and chlorine.

What are the elements in arsenic sulphide?

What are the elements in zinc sulphide?

What is the difference between cupric chloride and cuprous chloride?

What is an acid? Name and distinguish the two classes.

What terminations are used in naming the oxygen acids? What do they indicate?

What prefixes are used? What do they indicate?

What are the constituents of phosphoric acid?

Ans. Hydrogen, oxygen, and phosphorus.

By what part of the name is each one of these elements suggested?

What are the constituents of bromic acid?

Name the elements in sulphurous acid.

Name the elements in hyposulphurous acid.

What difference in the composition of the last two acids named, is indicated by their names?

What difference in composition is indicated by the names iodic acid and per-iodic acid?

What acid will be formed by the union of hydrogen, oxygen, and bromine?

Name the acid which contains oxygen, manganese, and hydrogen. *Ans.* Manganic acid.

What other acid with the same elements, but with a larger proportion of oxygen?

What are the elements in hydrofluoric acid?

Name the elements in hydriodic acid.

Name the elements in hydrosulphuric acid.

What is the difference in composition of hydrosulphuric acid and sulphuric acid?

Name the acid compound of hydrogen and iodine.

What is a salt? Haloid salts? What are oxy-salts?

How are the haloid salts named?

How are the oxy-salts named?

Name the salt from nitric acid and potassium.

Name the salt from copper and sulphuric acid.

Name the salt from hypochlorous acid and sodium.

Name the acid and the metal from which calcium carbonate may be derived.

Name the acid and the metal from which calcium hyposulphite may be derived.

What are the *elements* in ferrous sulphate?

What are the elements in barium sulphite?

Of what elements is magnesium carbonate composed?

What acid is required with copper to form copper nitrite?

What are the constituents of potassium chlorate?

Name the metal in the aluminium silicate.

Name the metal and the acid in magnesium citrate.

Name the elements in calcium phosphate.

What are the constituents of hydrogen sulphate?

What other name may be given to the same compound?

Name nitric acid as a salt. Chloric acid. Acetic acid.

Define base. What other name is given to this class?

How are the hydrates named?

What group of atoms occurs in the molecule of every hydrate? *Ans.* H O.

What are the elements of calcium hydrate?

What are the elements of magnesium hydrate?

What are the elements of copper hydrate?

Name the elements of silver hydrate.

Name the hydrate containing barium. What other elements does it contain?

CHAPTER II.

THE NON-METALLIC ELEMENTS.

SECTION I.

GENERAL DESCRIPTION.

43. The Number of the Elements. — The elements are those simple substances which, up to the present time, have not been decomposed. The existence of sixty-four has been established, and very recent researches seem to show that the existence of six or seven more is extremely probable.

These few elements are the alphabet of chemistry. They are united in pairs, or in greater numbers, to form the solid, liquid, and gaseous matter of the entire globe.

Not even all of this small number are found abundantly. The greater part of the whole mass of our earth is made up of less than half a score of these elements, while some of the others are of such rare occurrence as to be of little interest except to the chemist.

44. The following table contains the sixty-four elements, arranged in alphabetical order, with their symbols and atomic weights, for future reference.

To the chemist in his nice researches, the most exact values of atomic weights, which it is possible to obtain, are important; to the student, and for purposes of illustration, an approximate value, the nearest whole number, is usually sufficient.

The exact value for sodium, for example, is 22.99, but for all practical purposes the atomic weight of this element is 23.

NAMES.	Symbols.	Atomic Weight.	NAMES.	Symbols.	Atomic Weight.
Aluminium .	Al	27.3	Mercury . .	Hg	199.8
Antimony .	Sb	122.0	Molybdenum,	Mo	95.6
Arsenic . .	As	74.9	Nickel . . .	Ni	58.6
Barium . .	Ba	136.8	Niobium . .	Nb	94.0
Beyllium . .	Be	9.0	Nitrogen . .	N	14.01
Bismuth . .	Bi	210.0	Osmium . .	Os	198.6
Boron . . .	B	11.0	Oxygen . .	O	15.96
Bromine . .	Br	79.75	Palladium .	Pd	106.2
Cadmium . .	Cd	111.6	Phosphorus .	P	30.96
Cæsium . .	Cs	133.0	Platinum . .	Pt	196.7
Calcium . .	Ca	39.9	Potassium .	K	39.04
Carbon . . .	C	11.97	Rhodium . .	Rh	104.1
Cerium . . .	Ce	141.2	Rubidium . .	Rb	85.2
Chlorine . .	Cl	35.37	Ruthenium .	Ru	103.5
Chromium .	Cr	52.4	Selenium . .	Se	78.0
Cobalt . . .	Co	58.6	Silver . . .	Ag	107.66
Copper . . .	Cu	63.0	Silicon . . .	Si	28.0
Didymium .	Di	147.0	Sodium . .	Na	22.99
Erbium . .	Er	169.0	Strontium . .	Sr	87.2
Fluorine . .	F	19.1	Sulphur . .	S	31.98
Gallium . .	Ga	. . . .	Tantalum . .	Ta	182.0
Gold . . .	Au	196.2	Tellurium .	Te	128.0
Hydrogen . .	H	1.0	Thallium . .	Tl	203.6
Indium . .	In	113.4	Thorium . .	Th	231.5
Iodine . . .	I	126.53	Tin	Sn	117.8
Iridium . .	Ir	196.7	Titanium . .	Ti	48.0
Iron	Fe	55.9	Tungsten . .	W	184.0
Lanthanum .	La	139.0	Uranium . .	U	240.0
Lead . . .	Pb	206.4	Vanadium .	V	51.2
Lithium . .	Li	7.01	Ytrium . . .	Y	93.0
Magnesium .	Mg	23.94	Zinc	Zn	64.9
Manganese .	Mn	54.8	Zirconium .	Zr	90.0

It is well to remember that the chemist defines an element to be a substance which has *never yet been* decomposed. He does not claim to know that these sixty-four kinds of matter are really the ultimate simples of material bodies. Indeed, the science of *Spectrum Analysis*, in the hands of Mr. Lockyer, is giving some reason for supposing that they are in reality compounds of still more elementary forms. It is not impossible that more powerful analyses may yet decompose them.

45. Metals and Non-Metals. — The elements are usually divided into two classes, — the *metals*, and the *non-metals* or *metalloids*. The first includes such substances as iron, gold, silver, and copper; the second, such as sulphur, oxygen, nitrogen, and hydrogen. This division is purely arbitrary. It is not found in nature, but is made for the convenience of study. No line of division can distinctly mark the separation between the metals and the non-metals. Several elements, of which arsenic is one, have neither distinctly metallic nor non-metallic properties.

46. Classification. — The non-metals are divided into four groups. This division is founded on the relation between the *number of their atoms* which may enter a molecule together. This relation is seen by comparing the formulas of

Hydrochloric acid	$Cl\,H$,
Water	$O\,H_2$,
Ammonia	$N\,H_3$,
Marsh-gas	$C\,H_4$,

for in them we see that one atom of chlorine can hold but *one* atom of hydrogen in the molecule, while one atom of oxygen is able to hold *two*, one of nitrogen *three*, and one of carbon *four*.

Now, an atom of chlorine is never known to combine with *more than one of hydrogen*. We must consider it *equivalent* to one of hydrogen. So we say that an atom of oxygen is *equivalent* to two atoms of hydrogen, that one of nitrogen is *equivalent* to three, and one of carbon is *equivalent* to four of hydrogen.

To express this different equivalence, —

Chlorine	is called a	univalent	element	or a	monad.
Oxygen	"	bivalent	"	"	dyad.
Nitrogen	"	trivalent	"	"	triad.
Carbon	"	quadrivalent	"	"	tetrad.

The Groups. — But chlorine is not the only element whose atom is equivalent to one of hydrogen: there are bromine, iodine, and fluorine besides. And these four are grouped together, and called the UNIVALENT GROUP of non-metals.

Others are like oxygen, and with it form the BIVALENT GROUP: they are sulphur, selenium, and tellurium. In the same way we have the TRIVALENT GROUP, consisting of nitrogen, phosphorus, arsenic, and boron; and the QUADRIVALENT GROUP, containing only carbon and silicon. These four groups include all the non-metals.

Quantivalence. — Thus the atoms of the different elements have different values in chemical change. The atom of hydrogen is the *unit* by which to measure these *atomic values.* What is the value of a chlorine atom? One unit, because it is able to bind only one atom of hydrogen in a molecule. For a similar reason, the atom of any bivalent substance has a value of two units, of a trivalent substance three units, and of a quadrivalent substance four units.

This atomic value of a substance is called its QUANTIVALENCE. We accordingly say that the quantivalence of chlorine is 1, that of oxygen 2, of carbon 4. And we may define quantivalence to be *the combining power of a substance as measured by the number of hydrogen atoms, or their equivalent, which one of its atoms can bind together in a molecule.*

47. Graphic Symbols and Formulas. — The quantivalence of an element is represented by Roman numerals, or by primes or dashes attached to its symbol. For example: —

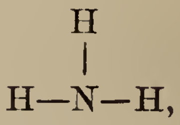
1 univalent atom of	chlorine . . .	Cl^{i} or	Cl' or	$Cl-$.
1 bivalent "	oxygen . . .	O^{ii} or	O'' or	$-O-$.
1 trivalent "	nitrogen . . .	N^{iii} or	N''' or	$-\overset{\vert}{N}-$.
1 quadrivalent "	carbon . . .	C^{iv} or	C'''' or	$-\underset{\vert}{\overset{\vert}{C}}-$.

The symbol of an element with its quantivalence shown by dashes attached to it is a Graphic Symbol.

The formula for a compound may then be made to show whether all the unit values of its elements are satisfied, and in what way. For example, in the molecule of water an atom of oxygen holds two of hydrogen. Put the symbols of hydrogen and oxygen together in this way, H—O—H, and we see that each of the two unit values of the oxygen is balanced by a unit value of hydrogen. A similar formula for

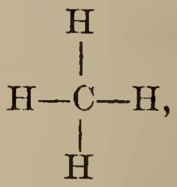
Hydrochloric acid is written . . . . H—Cl,

Ammonia is written . . . .



$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & — & N & — & H, \end{array}$$

Marsh-gas is written . . . . . .



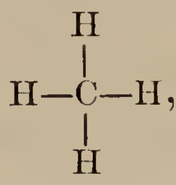
$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & — & C & — & H, \\ & & | & & \\ & & H & & \end{array}$$

in which we can see how every unit of quantivalence of Cl, of N, and of C, is balanced by a unit of hydrogen. Such formulas are called Graphic Formulas.

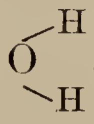
Unsaturated Molecules. — There is a compound of carbon and nitrogen called cyanogen. Its molecule contains one atom of each element, and its common formula is therefore C N. Its graphic formula is N≡C—. This graphic formula shows at a glance that three of the four units of the carbon atom are balanced by the three units of the nitrogen atom, and the fourth is unbalanced or free. A molecule like this, in which there is one or more unbalanced units of quantivalence, is called an Unsaturated molecule; while a molecule like that of water, in which all the units of quantivalence are balanced, is said to be Saturated.

The quantivalence of a *molecule* is the number of its unbalanced units. Thus the quantivalence of cyanogen, N≡C—, is 1, while that of water, H—O—H, is 0.

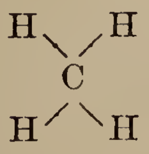
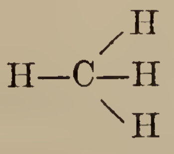
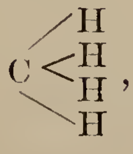
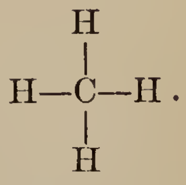
Constitutional Formula. — Now, these graphic formulas tell us something about the *way in which the atoms are put together* to make a molecule. For instance, in the formula for water, H—O—H, it appears that the atom of oxygen binds the two atoms of hydrogen to itself. The hydrogen atoms do not seem to be bound to one another at all. Nor can they be. Let us try to make the molecule by binding them together, and we have H—H and —O—, but the quantivalence of the hydrogen is already exhausted: there is not a single unit left free by which it can bind the oxygen.

So in the molecule of marsh-gas,  $H-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-H$, the carbon atom must be the *nucleus* about which the hydrogen atoms are arranged.

Since these graphic formulas show what is believed to be the way in which the atoms are grouped to constitute a molecule, they are called Constitutional Formulas.

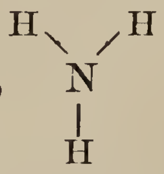
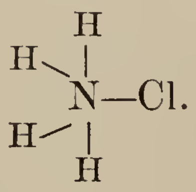
Caution. — We must not for a moment suppose that these formulas show the shapes of the molecules, or the geometrical arrangement of the atoms in them. We are not to infer that the molecule of water is a *line* of atoms H—O—H: it may be a *triangle*, , or it may have some other form. We *know nothing* of its shape.

The molecule of marsh-gas may be represented by

 ,  ,  , as well as by  $H-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-H$.

What the formula does mean to teach is, that all the atoms of hydrogen are bound by chemical attraction to the single atom of carbon, and not to one another.

The chemist studies the grouping of the atoms in the molecules, as the astronomer studies the grouping of the planets and the stars in cosmic systems, and represents the results in these constitutional formulas.

48. Variable Quantivalence. — The quantivalence of an atom is not always the same. In different compounds the same atom may have different values. For example, in the molecule of ammonia, , nitrogen is a trivalent atom, but in a molecule of ammonium chloride, NH_4Cl, it is pentivalent,  Its value is increased from 3 to 5.

Whenever the quantivalence of an atom changes, it changes by *two units*, as in the example just given. If the quantivalence of an atom is *even* at one time, it is *always* even. If the quantivalence of an atom is *odd* in one compound, it is odd in all compounds. Hence the elements are divided into two classes; one class containing all those whose quantivalence is represented by an even number, and the other class containing all those whose quantivalence is represented by an odd number. The members of the first class are called ARTIADS, and those of the second are called PERISSADS. An artiad may have a value represented by any even number, as 2, 4, 6, or 8, and a perissad may have a value represented by any odd number, as 1, 3, 5, or 7.

REVIEW.

I. — SUMMARY OF PRINCIPLES.

49. An element is a substance which has never yet been decomposed.

There are sixty-four elements whose existence is at present well established.

The elements are, for convenience of study, divided into two classes, metals and non-metals.

The quantivalence of an element is the chemical value of its atom, measured by the number of hydrogen atoms, or their equivalent, which it can hold in combination.

A univalent element is one whose atom is equivalent to one atom of hydrogen.

A bivalent element is one whose atom is equivalent to two atoms of hydrogen.

A trivalent element is one whose atom is equivalent to three atoms of hydrogen.

A quadrivalent element is one whose atom is equivalent to four atoms of hydrogen.

Among the non-metals we find no elements whose quantivalence is not one or another of these four values. Hence the non-metals are placed in four groups: the Univalent, the Bivalent, the Trivalent, and the Quadrivalent.

The quantivalence of an element is represented to the eye sometimes by a Roman numeral, sometimes by primes, sometimes by dashes attached to its symbol.

The quantivalence of an element is not always the same: it varies, however, only by the addition or subtraction of two units.

The formula for a compound is a combination of the symbols of its elements to represent the composition of a molecule.

A formula which shows nothing but what is actually known by experiment, the names and proportions of the elements, is an EMPIRICAL FORMULA.

A formula which represents any theory in regard to the arrangement of the atoms in the molecule is a RATIONAL FORMULA.

A formula which represents the quantivalence of the atoms, and shows how the atoms are grouped in the molecule, is a CONSTITUTIONAL FORMULA.

II.—EXERCISES.

Define element. How many elements are at present recognized? Are they known to be absolutely simple substances? Into what classes are they usually divided?

What relation gives a better classification? Illustrate the equivalence of Cl, O, N, and C atoms. Define the terms univalent, bivalent, trivalent, and quadrivalent. Define quantivalence.

How is the quantivalence of an atom represented? What do P''', Br', S'', $N\equiv$, $-O-$, and $-\overset{|}{\underset{|}{C}}-$ represent?

What is a graphic formula? Write the graphic formula for water, for hydrochlóric acid, for ammonia, and for marsh-gas.

When is a molecule said to be saturated? Unsaturated? Write the graphic formula for cyanogen. What is the quantivalence of cyanogen? Describe the quantivalence of any molecule.

What is a constitutional formula? What caution are we to observe in the use of constitutional formulas?

What is an empirical formula? What is a rational formula?

SECTION II.

HYDROGEN.

Fig. 22.

50. Preparation. — Pure water consists of hydrogen and oxygen, and there are many ways in which the hydrogen may be obtained from it. Many of the metals have power to take the oxygen alone from water. If a piece of sodium be dropped upon water, it melts into a globule, floats, and runs briskly around over the surface, taking the oxygen to itself, and letting the hydrogen go free. If the sodium be confined in a net of wire-gauze, it may be brought under the mouth of a test-tube previously filled with water. (Fig. 22.) The hydrogen will then be collected in the tube.

A more Practical Way. — Into a bottle, B, Fig. 23, are put some fragments of zinc. Through the well-fitting cork pass two tubes, one a funnel-shaped tube reaching down into the bottle, the other a bent tube reaching over to the cistern of water. If now a mixture of water and sulphuric acid is poured through the funnel until the lower end of its tube is covered, hydrogen will flow rapidly through the bent tube, and may be collected in jars upon the shelf of the cistern. None should be collected, however, until all the air in the apparatus has been driven out.

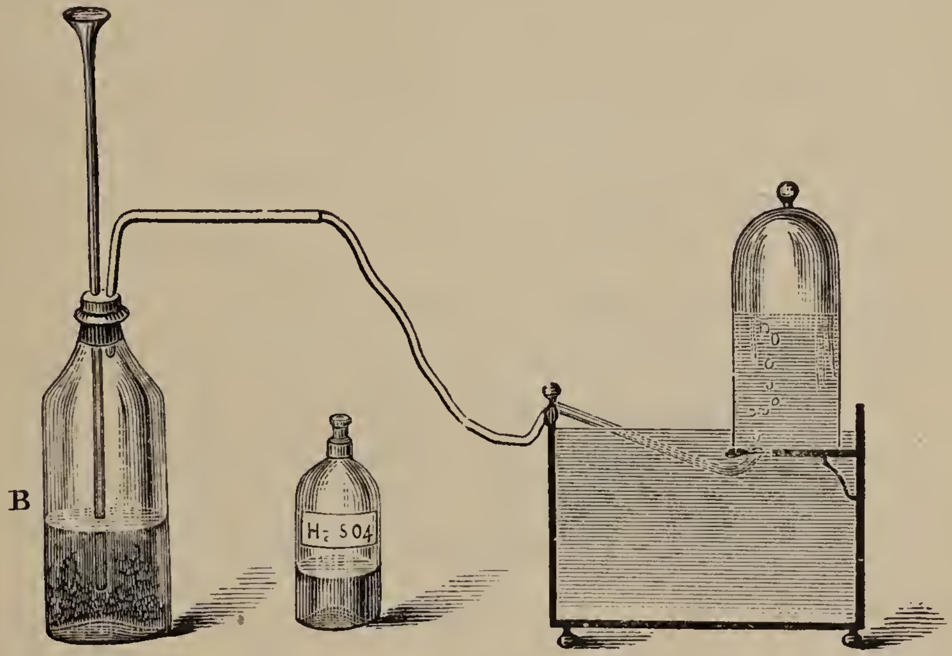


Fig. 23.

Explanation. — The molecule of sulphuric acid contains two atoms of hydrogen, one of sulphur, and four of oxygen. *An atom of zinc takes the place of the two atoms of hydrogen:* they are driven away, while a new molecule containing the atom of zinc and those of sulphur and oxygen is formed. This new molecule is zinc sulphate.

51. Reactions. — We may represent this chemical change to the eye by formulas and algebraic signs. Thus: —

$$\underset{\text{Sulph. acid.}}{H_2SO_4} \quad + \quad \underset{\text{Zinc.}}{Zn} \quad = \quad \underset{\text{Zinc sulphate.}}{ZnSO_4} \quad + \quad \underset{\text{Hydrogen.}}{H_2}.$$

The formulas for the substances which were put together in the experiment are joined by the sign of addition in this expression, and made the first member of an equation; while the new substances made by the chemical action are represented by their formulas in the second member.

Chemical changes are very generally called REACTIONS, and are represented by chemical equations like the one above.

Reaction of Sodium and Water.—For another example let us *write the reaction* when hydrogen is obtained by the action of sodium. In the experiment we put the metal in contact with water; let us add their formulas for the first member of our equation, and the formulas for the substances produced for the second.

$$\underset{\text{Water.}}{H_2O} \quad + \quad \underset{\text{Sodium.}}{Na} \quad = \quad \underset{\text{Sodium Hydrate.}}{NaHO} \quad + \quad \underset{\text{Hydrogen.}}{H}.$$

This equation shows us that one atom of sodium takes the place of one atom of hydrogen in the molecule of water, and produces one molecule of sodium hydrate, and sets free one atom of hydrogen.

No Loss nor Gain.— Not an atom is lost, nor an atom gained, in these exchanges. The second member must show the same number of atoms of every element found in the first, but differently arranged to make the formulas of the new compounds. The precision of exchanges is, beyond comparison, perfect. To illustrate this, let us write the same equation with the combining weights of the elements (see table, p. 53).

$$\begin{array}{ccccc}
H_2O & + Na = & NaHO & + H. \\
(2 \times 1 + 16) & + 23 = & (23 + 1 + 16) & + 1. \\
18 & + 23 = & 40 & + 1.
\end{array}$$

The two members of this equation are exactly equal.

This symbolic language of chemistry is of the greatest value. It shows, at a glance, an amount of information which, if spread out in ordinary language, would often be tedious or obscure, and reveals relations which in the ordinary language would be unseen.

The equation above tells us that every eighteen grams of water will require exactly twenty-three grams of sodium to completely decompose it, and that exactly forty grams of sodium hydrate will be produced, and just one gram of hydrogen will be set free.

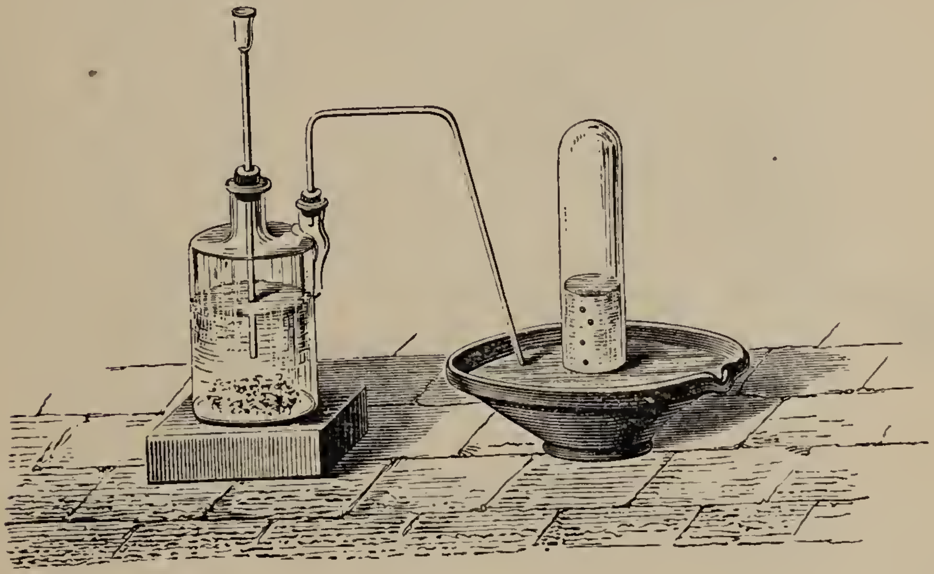


Fig. 24.

Reaction of Zinc and Hydrochloric Acid. — When we use zinc and hydrochloric acid as we may in the hydrogen generator (Fig. 24), the reaction may be written

$$\begin{array}{l} Zn + 2\,HCl = ZnCl_2 + H_2, \\ 65 + 73 = 136 + 2; \end{array}$$

which shows at a glance that in order to obtain two grams of hydrogen we must use sixty-five grams of zinc and seventy-three of hydrochloric acid.

52. Physical Properties. — Hydrogen is a colorless gas, and when pure has neither taste nor odor. It is the lightest of known substances; its specific gravity is only .0692 (air = 1). It is eleven thousand times lighter than water.

Let us take two small glass vessels, and fill one with hydrogen, the other with air. Then suppose we bring their mouths together closely, so that the one containing hydrogen shall be above the other (Fig. 25). Next turn them quickly over, to bring the air-vessel to the top. A moment afterward we may remove the lower jar, and then touch the mouth of the upper one with a lighted match: a sharp explosion tells us that the hydrogen is there. It has risen through the air to the top, just as oil will rise through water to its surface.

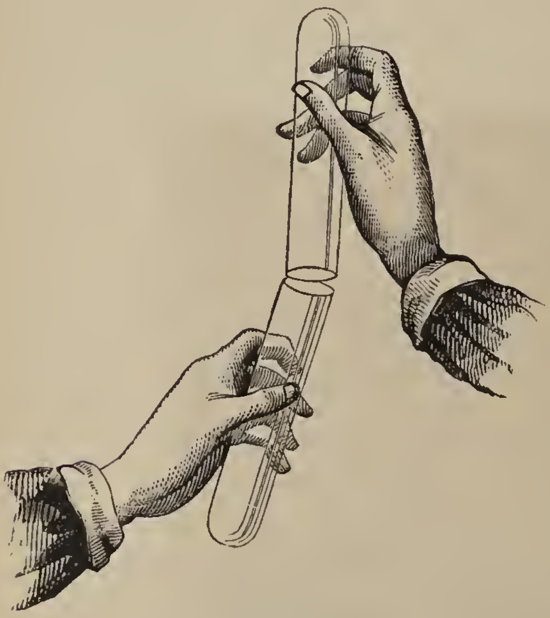


Fig. 25.

Soap-bubbles blown with cold air will fall to the floor, but if blown with hydrogen they will rise quickly to the ceiling (Fig. 26).

This gas is slightly soluble in water, one hundred cubic inches of which will absorb one and a half of hydrogen.

Hydrogen is capable of being absorbed by many metals when heated to temperatures more or less elevated. The most remarkable absorption is by the metal palladium, which, at ordinary temperatures, will take up more than nine hundred times its own volume. Such absorption of gases by metals is called Occlusion.

Hydrogen, by intense cold and enormous pressure, may be reduced to the liquid form. This was first done by two experimenters, M. Cailletet and M. Pictet, working independently, but succeeding at the same time, both in December, 1877. M. Cailletet brought the gas under a pressure of nearly three hundred atmospheres, and then suddenly released it. By the sudden expansion of a portion of the gas the temperature of the rest was enormously reduced, and the hydrogen took the form of mist.

M. Pictet subjected the gas to a temperature of –140 C., and a pressure of six hundred and fifty atmospheres. On opening a valve the hydrogen came out in a jet of steel-blue color. But after a little while, on opening the valve a second time, and although *the pressure was still three hundred and fifteen atmospheres, no hydrogen came out.* The fact that no gas issued when under such pressure showed that the hydrogen was not only liquefied but actually reduced to the solid form: it was frozen.

Fig. 26.

53. Chemical Properties. — Hydrogen is a combustible gas. If a small jar, filled with it, is carefully lifted from the cistern, and a lighted taper is quickly pushed up into it (Fig. 27), the gas takes fire with a slight explosion.

We may study the flame by means of the apparatus shown in Fig. 28. The hydrogen generator is connected with a second bottle by a bent tube reaching down nearly to the bottom. This bottle contains some water, and is provided with a tall tube whose upper end is drawn to a small opening. Let the acid be poured through the funnel upon the zinc, until a rapid evolution of the gas is produced, and *wait until the gas has driven the air completely from the apparatus.* Then touch the jet of escaping hydrogen with a match-flame: it will take fire, and continue to burn as long as the gas is supplied.

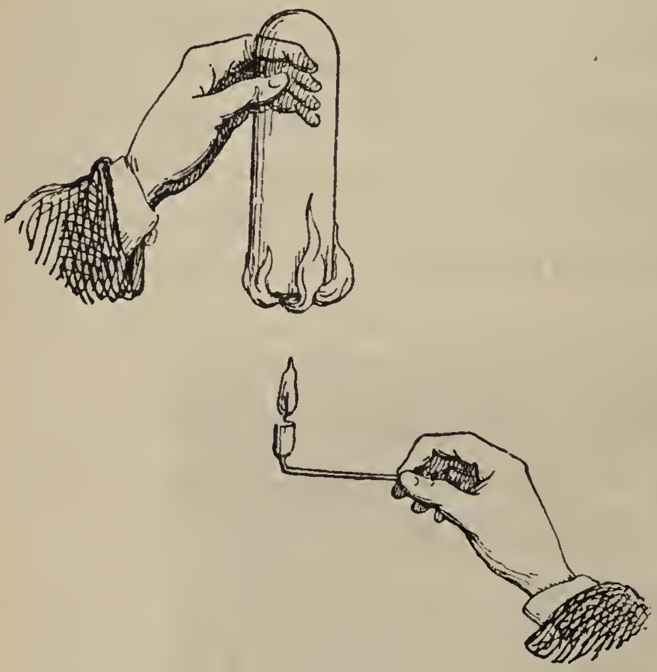


Fig. 27.

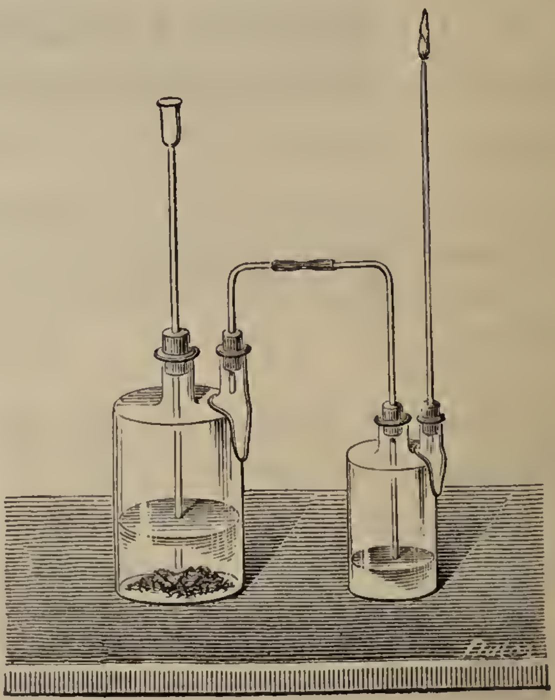


Fig. 28.

The flame of pure hydrogen yields almost no light, but its heat is intense. One gram of hydrogen will yield heat enough to raise 34,462 grams of water from 0° to 1° C.

The mixture of hydrogen with air is explosive. When the mixture is made with oxygen, in the proportions of two volumes hydrogen to one volume oxygen, the explosion is deafening. If it takes place in free air, it is not dangerous; but, if the mixture is confined in a tight bottle, the fragments of the bottle will be driven with violence as by a charge of gunpowder.

The combustibility and the explosibility of hydrogen are due to the strong chemical attraction between it and oxygen. It also enters into combination with the other non-metals with the exception of boron; but its compounds with the metals are rare. By some chemists it is regarded as itself a metal, and has been called Hydrogenium. This view is based upon the general resemblance of the chemical actions of hydrogen and the metals.

The molecule of hydrogen is supposed to contain two atoms. Its symbol is H, which represents one atom, but its molecule is represented by H—H, or H_2.

54. Occurrence. — Hydrogen rarely occurs in nature free; it has been found among the gases given off in volcanic eruptions; but in compounds it is a very abundant element. One-ninth, by weight, of all the water of the globe, is hydrogen; and besides this it is an important element in all animal and vegetable bodies.

55. Hydrogen as a Standard. — This element has been adopted as the standard of atomic weight, of equivalence, of density of gases, and of molecular weight.

The weight of its atom is 1 = unit of atomic and molecular weight.
Its atom value or quantivalence is . . 1 = " quantivalence.
The weight of a unit volume is . . . 1 = " density or specific gravity.
Its molecular weight = 2 atomic weights is . . . 2
The weight of 1 liter is . . . . 0.08936 gram.

REVIEW.

I. — SUMMARY OF PRINCIPLES.

56. Hydrogen may be obtained from water by the action of sodium or of potassium at ordinary temperature.

It may also be obtained from water by iron and some other metals at a red heat.

It may also be obtained from water by electrolysis.

It is usually obtained by the action of zinc upon dilute sulphuric or hydrochloric acid.

Hydrogen is the lightest known substance: it is 14.4 times lighter than atmospheric air.

With air it is combustible and explosive. It will not support combustion nor animal life.

Its symbol is H, and its atomic weight is 1. Its molecule is represented by H—H, and its molecular weight is 2. Its density is 1, and its quantivalence is 1.

Chemical changes are called reactions. Reactions are written in the form of equations. The formulas of the substances entering into the action are connected by the sign of addition to form the first member, while the formulas of the substances produced by the action, joined by the same sign, form the second member, of the equation.

The molecular weights of the various substances represented in the equation show the relative quantities of them which take part in the reaction.

II. — EXERCISES.

Describe the preparation of hydrogen by the use of sodium. By the use of zinc. Give the explanation.

What is a chemical reaction? How are reactions written? Write the reaction of sodium and water in the preparation of hydrogen. What does the equation show? Write the equation with the numerical values of the substances. What does this numerical equation show?

In the equation we find that eighteen grams of water will yield just one gram of hydrogen: then how much water would be needed to yield twenty grams of hydrogen?

How much sodium would be needed with it?

How much sodium hydrate would be formed?

How much water would be required to yield seventy-five grams of hydrogen? *Ans.* 1,350 grams.

How many liters of hydrogen will be liberated from one hundred grams of water by sodium?

Ans. 100 grams H_2O yield $\frac{100}{18}$ of 1 gram $H = 5.555+$.

A liter of H weighs 0.08936 gram; hence $\frac{5.555}{0.08936}$ = the number of liters $= 62.16+$.

Write the reaction of zinc and hydrochloric acid. Change it to a numerical equation.

How much zinc must be used to prepare seventy-five grams of hydrogen?

How many liters of hydrogen will be set free by using one hundred grams of zinc?

How much hydrochloric acid will be required to liberate one hundred liters of hydrogen?

Ans. $0.08936 \times 100 = 8.936$ grams, the weight of 100 liters H; to liberate 2 grams H, requires 73 of H Cl (see reaction); hence $\frac{8.936}{2} \times 73 = 326.164$ grams H Cl will be required for 100 liters of H.

How many grams of H_2SO_4 must be decomposed by zinc to yield one hundred liters of H?

Give the physical properties of hydrogen. What is occlusion?

Why is hydrogen combustible? How much heat is evolved by the combustion of one gram of hydrogen?

Where is hydrogen found in nature? Give the symbol of hydrogen. Atomic weight. Quantivalence. Density. Molecular weight. How is the molecule of hydrogen represented in symbols?

SECTION III.

THE UNIVALENT NON-METALS.

I. — CHLORINE.

57. Preparation from Hydrochloric Acid. — The chlorine in hydrochloric acid may be set free by the action of manganese dioxide. For this purpose the acid, with about one-third of its weight of the dioxide, is put into a large flask (Fig. 29), provided with a tight cork and suitable tubes as represented in the cut. By heating the mixture gently, the chlorine is given off in abundance; passes over through the tube to the bottom of the jar, A, in which it gradually rises, until the jar is filled.

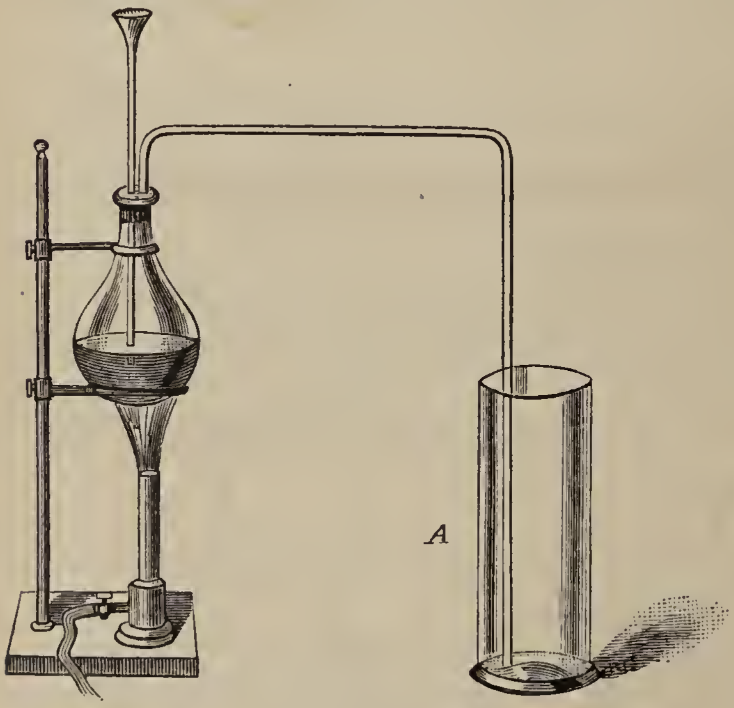


Fig. 29.

The reaction is as follows: —

$$MnO_2 + 4\ HCl = 2\ H_2O + MnCl_2 + 2\ Cl.$$

This equation is read in this way: One molecule of manganese dioxide with four molecules of hydrochloric acid, yield two molecules of water, one of manganese dichloride, and two atoms (one molecule) of chlorine.

The oxygen of the dioxide, combining with the hydrogen of the acid, forms water; while a part of the chlorine, taking the manganese, forms manganic chloride, and the rest of the gas is set free.

Preparation by Use of Common Salt. — In place of hydrochloric acid we may use sodium chloride and sulphuric acid: these with the manganese dioxide, gently heated, evolve the gas abundantly. The following equation shows the chemical changes which occur: —

$$\underset{\text{Sodium Chloride.}}{2\,NaCl} + \underset{\text{Sulphuric Acid.}}{3\,H_2SO_4} + \underset{\text{Manganese Dioxide.}}{MnO_2} = \underset{\text{Hydro-Sodium Sulphate.}}{2\,NaHSO_4} + \underset{\text{Manganese Sulphate.}}{MnSO_4} + \underset{\text{Chlorine.}}{2\,Cl} + \underset{\text{Water.}}{2\,H_2O.}$$

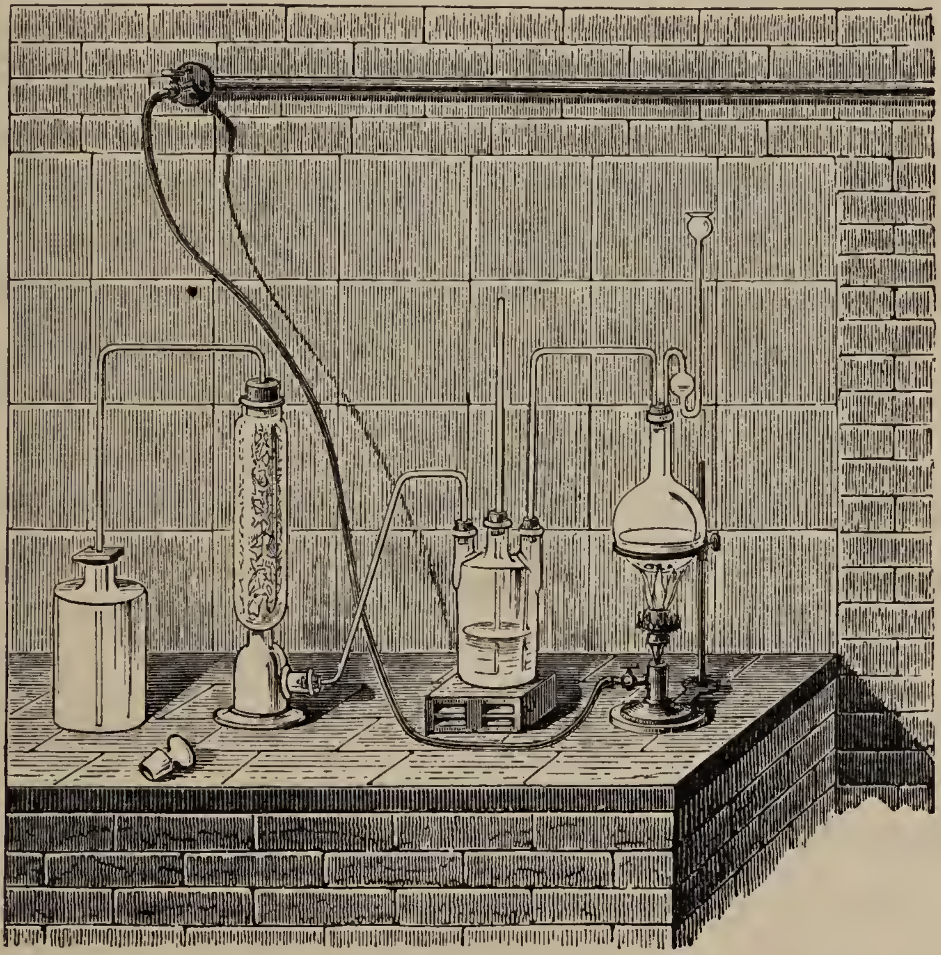


Fig. 30.

To Obtain the Gas Pure and Dry. — To purify and dry the gas, obtained by either of these methods, it is passed from the flask first through water contained in a wash-bottle, — the three-necked bottle in Fig. 30, — and then through a vessel — the tall jar in the figure — containing calcium chloride. The water washes out the impurities, and the chloride absorbs the moisture. The pure, dry chlorine may then be collected in bottles or jars.

58. Physical Properties. — Chlorine is a transparent gas, having a greenish-yellow color and a characteristic suffocating smell. When inhaled, it powerfully irritates the air-passages, and produces coughing, even when mixed with large portions of air. If breathed pure it even causes death.

Chlorine is nearly two and a half (2.45) times heavier than an equal volume of air. Its density, hydrogen being the standard, is 35.37.

Chlorine is very soluble in cold water. Water, at ordinary temperature, dissolves about twice its own bulk. It dissolves less and less as the temperature rises, until at 100° C. it dissolves none.

The solution of chlorine is useful in the laboratory; and it is prepared by passing the gas from the generator through water, as shown in Fig. 31. The gas is washed in the small bottle, and dissolved in those which follow.

Under a pressure of four atmospheres, chlorine becomes a yellow liquid at ordinary temperature. If cooled to −34° C., the same effect is produced without any extra pressure.

59. Chemical Character. — Chlorine combines readily with hydrogen. The application of either heat or light will cause them to unite. When a mixture of the two is placed in the direct rays of the sun, the combination is rapid, sometimes with explosion. Diffuse light causes the combination gradually; while, if mixed and kept in the dark, no chemical action takes place.

So strong is the chemical force, or affinity, between these elements, that chlorine will decompose many compounds of hydrogen. An instructive experiment illustrates this. A lighted wax-taper, plunged into a jar of chlorine, is extinguished; but, curiously enough, it is at once relighted, burning afterward with a dark red flame, and giving off a dense black smoke. The explanation is this: the white flame in the air is due to the action of oxygen; and since there is none of this element in the jar, the flame dies. But the wax contains hydrogen; and the chlorine decomposes the wax, and combines with this element so vigorously as to produce a dull red flame.

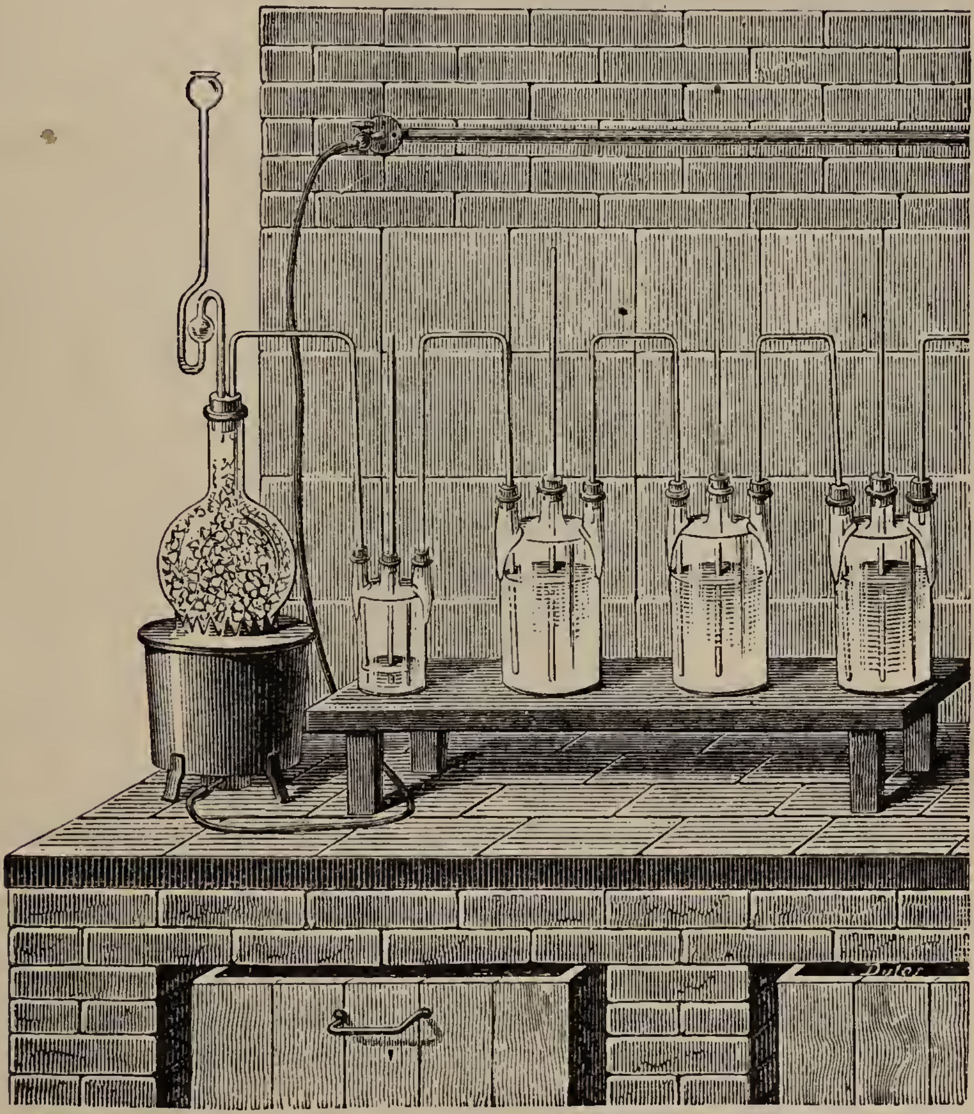


Fig. 31.

Application of this Affinity. — Chlorine is largely used in the arts of bleaching and disinfecting. But *dry* chlorine will not bleach: it must be moist. Its power to destroy colors and odors is due to its attraction for hydrogen.

In bleaching, the chlorine decomposes the water present, and combines with its hydrogen. The oxygen of the water then attacks the coloring matter, decomposing it and destroying the color.

Chlorine and the Metals. — Chlorine combines, also, with most metals readily. Powdered antimony sprinkled into a bottle of chlorine takes fire (Fig. 32), and falls to the bottom in a shower of sparks. Metals which, like gold, resist the action of oxygen and the acids, are attacked by chlorine. Neither nitric acid nor hydrochloric acid alone will act upon gold, but when mixed they form what is called *aqua-regia*, in which there is free chlorine; by this mixture gold trichloride is speedily formed.

Fig. 32.

60. Occurrence. — Chlorine is not found free in nature, but in combination it is one of the most abundant elements. Sodium chloride (common salt, $NaCl$) is distributed throughout the air, the soil, the rocks, and the sea. More than half $\left(\frac{35.5}{58.5}\right)$ the weight of this substance is chlorine; and calculating from the amount of salt in sea-water, we may find that something more than five gallons of this gas is contained in the salt of one gallon of sea-water. Other chlorides exist in the water of the sea: potassium and magnesium chlorides are the most abundant. Chlorides are found also in the bodies of plants and animals.

II. — BROMINE, IODINE, AND FLUORINE.

61. Bromine. — Bromine is found in small quantities combined with metals in the waters of the sea. It is a dark brown-red liquid, having a strong affinity for hydrogen and the metals. Its compounds are used to some extent in medicine, and quite largely in photography.

62. Iodine. — Iodine, like chlorine and bromine, is not found free in nature, but combined with metals it is found in sea-water, in sea-plants, and in some mineral springs.

It is a blue-black crystalline solid, very volatile, giving off a superb violet-colored vapor when warmed. It has a strong affinity for hydrogen and the metals. Its compounds are used in photography, and are highly prized in medicine.

63. Fluorine. — Fluorine is found combined with metals in certain minerals — fluor-spar (calcium fluoride, $Ca\ Fl_2$) being the most common. It is obtained free only with the greatest difficulty, and on this account its physical properties are imperfectly known. It seems to be a gas, having a violent affinity for hydrogen, for the metals, and, indeed, for many other elements.

III. — THE GROUP.

64. Comparison of their Physical Properties. — Leaving fluorine out of the account, we notice that chlorine is a gas at ordinary temperature, bromine a liquid, and iodine a solid. At a little higher temperature all three are gaseous. In the colors of these gases we notice a curious gradation. That of bromine represents one end of the spectrum, dark-red; that of chlorine represents the middle part, greenish-yellow; and that of iodine represents the other end, a beautiful violet.

Comparison of their Chemical Properties. — These four elements closely resemble each other in their chemical properties. They combine with the same substances, and generally in the same proportions. For hydrogen and the metals they all have strong attraction ; with oxygen they (with the possible exception of fluorine) unite, although with a feeble force ; with nitrogen they form explosive compounds. They are all univalent, and the single compound which each one forms with hydrogen is an acid.

The following formulas illustrate the analogous composition of the hydrogen compounds of these elements : —

HCl	Hydrochloric acid.
HBr	Hydrobromic acid.
HI	Hydriodic acid.
HF	Hydrofluoric acid.

IV. — Hydrochloric Acid.

65. Hydrochloric Acid is a Gas. — The hydrochloric acid found in commerce is a liquid, but a simple experiment will show that this liquid is the solution of a gas in water.

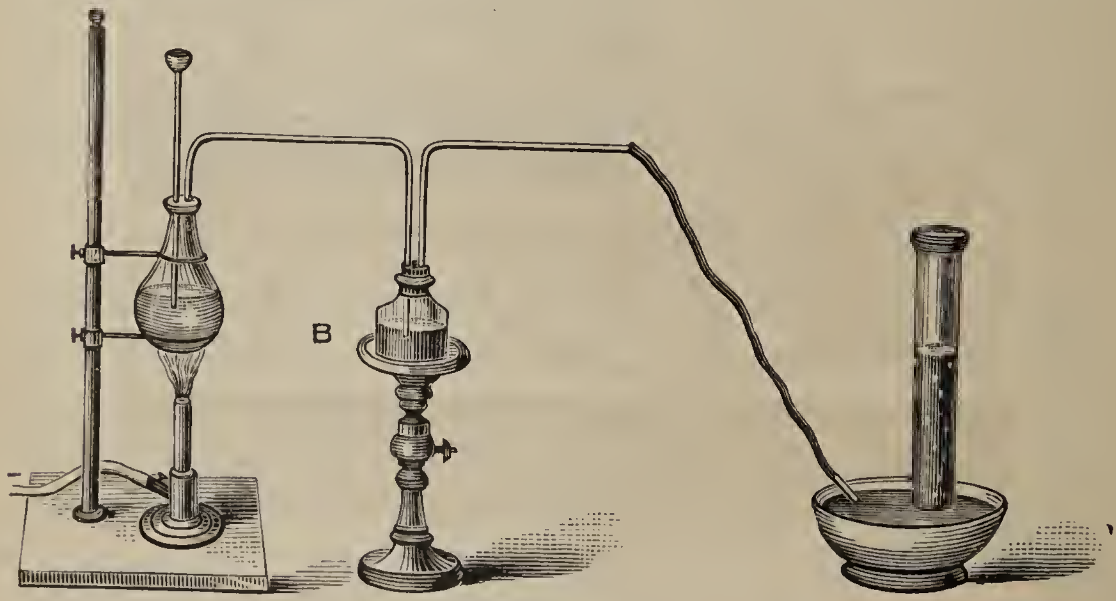


Fig. 33.

Some of the liquid is put into a flask (Fig. 33), and heated : a colorless gas is by this means driven over through the bent tube, and, being dried while going through sulphuric acid in the bottle B, finally enters a jar, previously filled with mercury, and inverted over a small cistern of the same fluid. If, when the jar is full of gas, it be taken from the mercury and its open mouth inserted in water, the gas will be dissolved, the water rising into the jar at the same time with surprising swiftness. Now, the solution thus obtained is found to be weak hydrochloric acid, and we hence learn that the real acid is a gas.

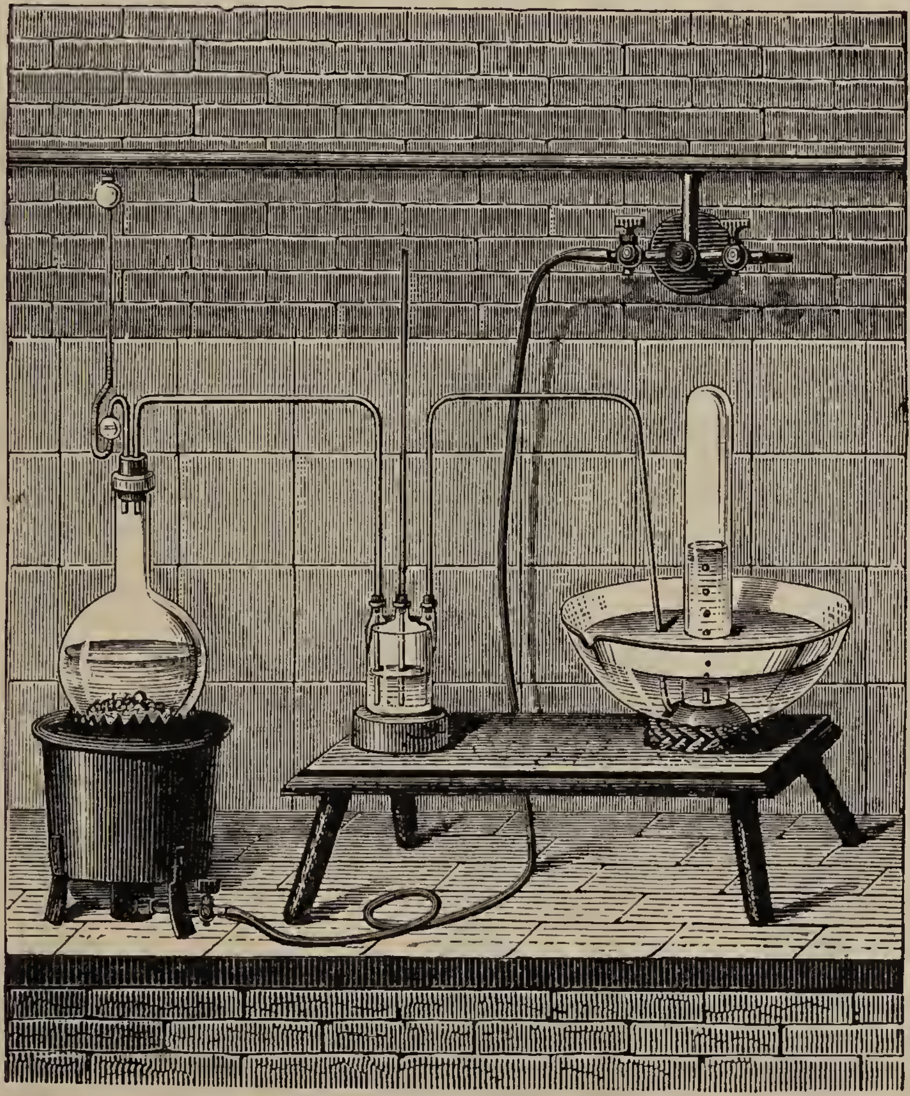


Fig. 34.

66. Preparation. — For some purposes the gas may be obtained in small quantity, by the process just described.

It may, however, be made by the reaction of common salt with sulphuric acid.

For this purpose six parts by weight of salt are put into a large flask, and eleven parts of strong acid are slowly poured through a funnel-tube upon it. The gas is rapidly set free, and is purified by passing through a little water in a wash-bottle; after which it may be collected over mercury, as shown in Fig. 34, or by displacement like chlorine, or it may be dissolved by passing it into water.

The Reaction. — The salt furnishes the chlorine, and the sulphuric acid furnishes the hydrogen, to form the hydrochloric acid. The chemical change is as follows: —

$$NaCl + H_2SO_4 = NaHSO_4 + HCl.$$

A molecule of sodium chloride, with a molecule of sulphuric acid, yields a molecule of hydro-sodium sulphate and one molecule of hydrochloric acid.

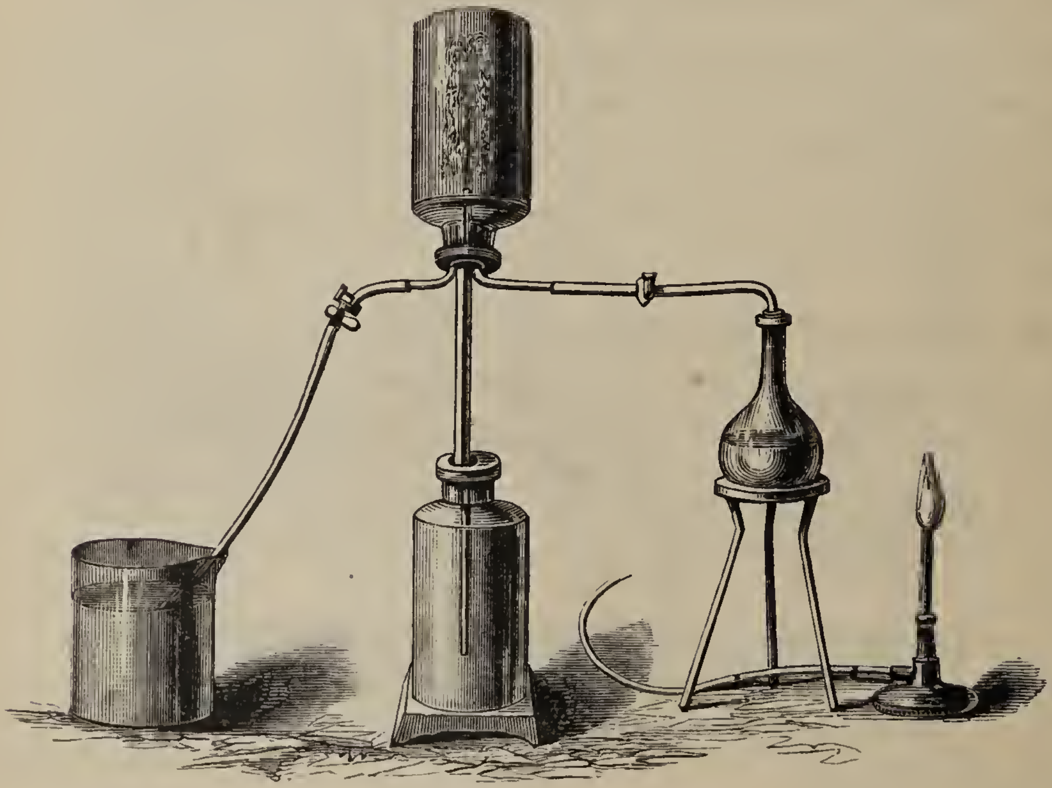


Fig. 35.

67. Physical Properties. — Hydrochloric acid is a colorless gas, a little heavier than air, and 18.2 times heavier than hydrogen.

One of its most remarkable properties is its solubility in water; and it is this solution which is found in commerce and used in the laboratory, as hydrochloric acid. The great solubility of this gas may be demonstrated by experiment with apparatus shown in Fig. 35. Some commercial acid is heated in the small flask. The acid gas is rapidly evolved. It passes into the upper bottle, and drives the air before it, out through the side tube, which reaches over to the surface of water in a beaker. When the air has been all driven out, this side tube is closed by a spring-clamp. The gas then passes down the vertical tube, which reaches from the upper bottle to the bottom of another below containing water colored blue with litmus. The flame is then withdrawn from the flask, and the other clamp immediately closed. At this moment the upper bottle and the vertical tube are filled with the acid gas. The water then quickly absorbs the gas, and ascends the vertical tube. It enters the upper bottle in the form of a fountain, and continues to rise as long as any gas remains to be absorbed. At 15° C., water dissolves four hundred and fifty times its own volume of hydrochloric gas. On contact with air the colorless gas takes moisture, and yields dense white fumes.

68. Chemical Properties. — Hydrochloric acid is an energetic acid. It reddens litmus, and it readily yields its hydrogen when acted on by metals. For example: —

HCl	$+$	Na'	$=$	$NaCl$	$+$	$H.$
Acid.		Sodium.		Sodium Chloride.		Hydrogen.
$2HCl$	$+$	Zn''	$=$	$ZnCl_2$	$+$	$2H.$
Acid.		Zinc.		Zinc Chloride.		Hydrogen.

The metals take the place of the hydrogen of the acid, and produce *chlorides*.

The gas is wholly irrespirable. It will neither burn nor support combustion.

The molecular weight of hydrochloric acid is 36.5 ($H = 1$, $Cl = 35.5$), and its molecular volume is two. Hence its density is $\frac{35.5}{2} = 18.25$.

69. Uses. — In the manufacture of many commercial products, such, for example, as gelatine, "bleaching-powder," and many chemicals, hydrochloric acid is used in very great quantities. In the laboratory of the chemist it is an indispensable reagent.

REVIEW.

I. — SUMMARY OF PRINCIPLES.

70. Chlorine is obtained by gently heating a mixture of hydrochloric acid and manganese dioxide.

It is a heavy greenish-yellow gas, with a penetrating and characteristic odor. It is soluble in one-half its volume of cold water.

Its most characteristic attractions are for hydrogen and the metals. With hydrogen it combines in equal volumes, and forms hydrochloric acid; with the metals it forms a large class of compounds called chlorides.

Its affinity for hydrogen is taken advantage of in the process of bleaching cotton fabrics. Only moist chlorine will remove color or destroy bad odors.

The gas is not found in nature free, but its compounds are very abundant. The most abundant chloride is common salt, — the sodium chloride ($NaCl$).

Chlorine, Bromine, Iodine, and Fluorine form a well-marked chemical group. They are all univalent elements. They enter into combination with the same elements, and their compounds have similar composition.

Hydrochloric acid is the only known compound of chlorine and hydrogen. It was formerly called MURIATIC ACID.

It is manufactured in large quantities for many commercial purposes. Common salt and sulphuric acid are the materials employed. When heated together they evolve large volumes of the gas, which is dissolved in water to constitute the commercial acid.

The commercial acid is usually highly colored yellow by the presence of organic impurities. The pure article is perfectly colorless.

Hydro-sodium sulphate ($Na\,H\,S\,O_4$) or sodium sulphate ($Na_2\,S\,O_4$) is formed at the same time with the acid: the first, if one molecule of salt to one of sulphuric acid are the proportions used; and the second, if two molecules of salt are used to one of the acid, thus:—

$$2\,Na\,Cl + H_2\,S\,O_4 = Na_2\,S\,O_4 + 2\,H\,Cl.$$

This sodium sulphate ($Na_2\,S\,O_4$) is used in the manufacture of sodium carbonate. It is the well-known *Glauber's Salts.*

Hydrochloric acid changes many metals into chlorides. The chemical action consists in the substitution of an atom of the metal for the hydrogen, in one or more molecules of the acid.

The substitution will be for hydrogen in one molecule of acid if the metal is univalent, in two molecules of the acid if the metal is bivalent, and in three if the metal is trivalent.

These chlorides are all very soluble in water, except three. Silver chloride ($Ag\,Cl$), mercurous chloride ($Hg_2\,Cl_2$), and lead chloride ($Pb\,Cl_2$), are the three exceptions.

A mixture of hydrochloric acid and nitric acid is called AQUA REGIA. Neither of these acids alone can attack gold, but aqua-regia quickly converts gold into a chloride ($Au\,Cl_3$). Aqua-regia owes this power to dissolve gold to the *free chlorine* which it contains.

II.—EXERCISES.

Describe the preparation of chlorine.

Give the reaction. Take the needed atomic weights from the table of elements, p. 53. and write the numerical equation in the reaction.

How many grams of Cl may be obtained by using fifty grams of hydrochloric acid? *Ans.* 24.315.

How many liters in this weight of chlorine: its density being 35.5?

Ans. 0.08936 × 35.5 = weight of 1 liter of Cl.

$$\frac{24.315 \text{ grams}}{\text{weight of 1 liter}} = \text{number of liters of Cl.}$$

What volume of chlorine will fifty grams of $Mn\,O_2$ liberate from H Cl when the two are heated together?

What is the weight of 2,500 cubic centimeters of chlorine gas?

What are the physical properties of chlorine? Illustrate the chemical attraction of chlorine and hydrogen. To what useful purposes is this property of chlorine applied? Explain the bleaching action of chlorine.

Illustrate the chemical attraction of chlorine and the metals. What class of compounds does chlorine form? What is aqua-regia?

In what form is chlorine abundant in nature?

Give a brief description of bromine.

Give a brief description of iodine.

Give a brief description of fluorine.

Compare the physical properties of chlorine, bromine, and iodine. Compare their chemical properties. Give the formulas and names of their hydrogen compounds. What is the quantivalence of the members of this group?

What is the liquid hydrochloric acid of commerce? How may the gas itself be obtained? Write the reaction when salt is used with sulphuric acid.

What are some of the properties of this gas?

Describe its solubility.

What chemical change occurs when sodium and hydrochloric acid are brought together?

What chemical change when zinc and hydrochloric acid are brought together?

What class of substances are formed by the action of metals on this acid?

What are the uses of this acid?

SECTION IV.

THE BIVALENT NON-METALS.

I. — OXYGEN.

71. Preparation. — Oxygen may be obtained by heating potassium chlorate. This is the best and usual method. Potassium chlorate is a white solid, about 39.2 per cent of its weight being oxygen. It gives up all this oxygen when heated. For this purpose it is finely powdered, mixed with about an equal weight of black oxide of manganese, and put into a flask, F (Fig. 36). A bent tube reaches through the cork of the flask and over into the water of the cistern; and upon the shelf of the cistern there stands an inverted jar filled with water. Now, when the flask is heated, the chlorate is decomposed, and oxygen, being set free, passes through the bent tube, and bubbles out of the water. If the end of the tube is brought under the mouth of the jar, the oxygen will rise into the jar, which in a little time will be filled. The reaction is as follows: —

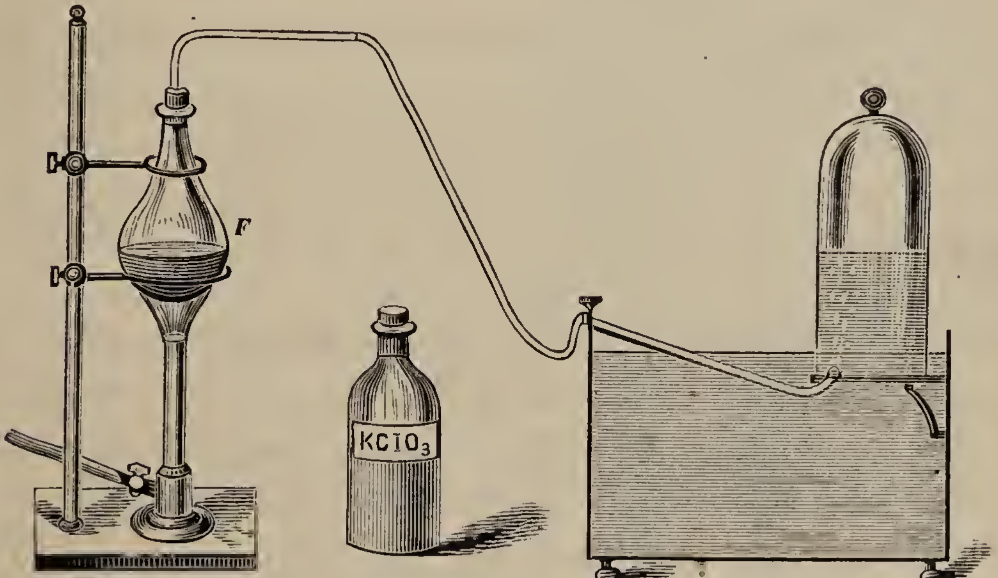


Fig. 36.

$$\begin{array}{cccccc} KClO_3 & = & KCl & + & O_3. \\ \text{Potassium chlorate} & = & \text{Potassium chloride} & + & \text{Oxygen.} \\ 122.6 & = & 74.6 & + & 48. \end{array}$$

The manganese dioxide is used in order to obtain the gas at a lower temperature. It takes no part in the chemical action, and, in the end, is found mixed with the potassium chloride in the flask.

72. Its Physical Properties. — Oxygen is a colorless and transparent gas; without odor or taste; a little heavier than air, its specific gravity being 1.1056 (air = 1), or 16 when hydrogen is the unit.

This gas is slightly soluble in water, one hundred volumes of water absorbing about three volumes of oxygen at ordinary temperatures.

When submitted to a pressure of three hundred atmospheres, and a cold of −140° C., if the pressure be suddenly removed oxygen is reduced to the liquid form. This liquefaction was accomplished by Pictet in December, 1877.

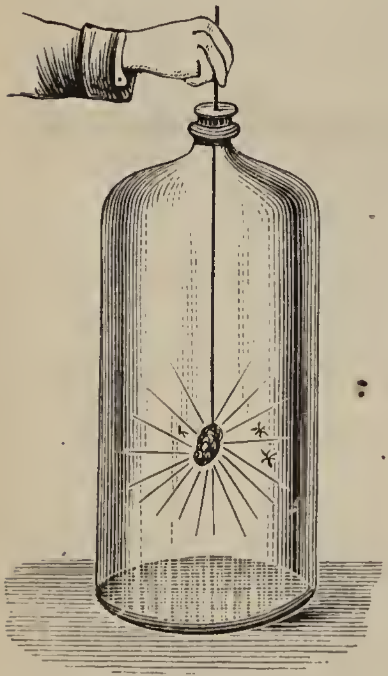


Fig. 37.

73. Its Chemical Properties. — A lighted taper lowered into a jar of oxygen burns with surprising brilliancy. A glowing spark upon the wick is all that is needed: the oxygen instantly, and with a slight explosion, kindles it into a vivid flame.

Into a jar of oxygen, hang, by means of a wire, a piece of charcoal bark on which there is a spark of fire (Fig. 37). Quickly the charcoal bursts into a beautiful and vigorous combustion. The carbon combines with the oxygen, and produces carbon dioxide, thus: —

$$\begin{array}{ccccc} C & + & O_2 & = & CO_2 \\ \text{Carbon} & + & \text{Oxygen} & = & \text{Carbon dioxide.} \end{array}$$

Experiments like these illustrate the fact that bodies which burn in air will burn with greater vigor in oxygen.

Again: let an iron wire or a steel watch-spring be tipped with a bit of wood. Set fire to the wood, and at once plunge it into a jar of oxygen. Quickly the metal takes fire, and burns (Fig. 38), the iron with a steady and beautiful light, or the steel with a multitude of star-like sparks. Iron sesquioxide ($Fe_2 O_3$) is produced by this chemical action. Experiments such as this illustrate the fact that substances which do not burn in air may burn with great rapidity in oxygen.

This gas not only supports combustion, it is also necessary to the life of animals. It is in the air, and animals breathe it; it goes into their blood, and purifies it. By being mixed with nitrogen, its violent action is toned down so that the most delicate organ may not only withstand it, but be invigorated by its presence.


Fig. 38.

Oxygen unites with all other elements with the exception of fluorine, and its compounds are very numerous and abundant.

Its combining weight is 16. Its molecule is supposed to contain two atoms; so that while the symbol of oxygen is O, its molecule is represented by $O = O$ or O_2, and its molecular weight is therefore 32, and molecular volume 2.

74. Occurrence. — Oxygen is the most abundant element in nature: minerals, plants, and animals alike contain large quantities of it. One-fifth part of the air by weight is oxygen, eight-ninths of all the water on the globe, and about one-half of all the solid rocks. Besides this about four-fifths of the weight of vegetable bodies is oxygen, and about three-fourths of that of animals. It is, perhaps, not too much to say, that one-half of all the matter of the world, as far as it has been examined, is oxygen. And yet, when freed from its prisons in solid and liquid bodies, oxygen is a gas, invisible as air and but little heavier.

75. Ozone. — Let strips of unsized paper be soaked in a solution of potassium iodide mixed with starch. If these strips be hung in a jar of air, no action or change will occur; but if a few drops of ether be added, and a hot glass rod be put into the jar (Fig. 39), the paper will very soon become colored blue. The oxygen, which at first could not attack the iodide, has been changed by the ether and the heat so that it can. This more *active* form of oxygen is called OZONE.

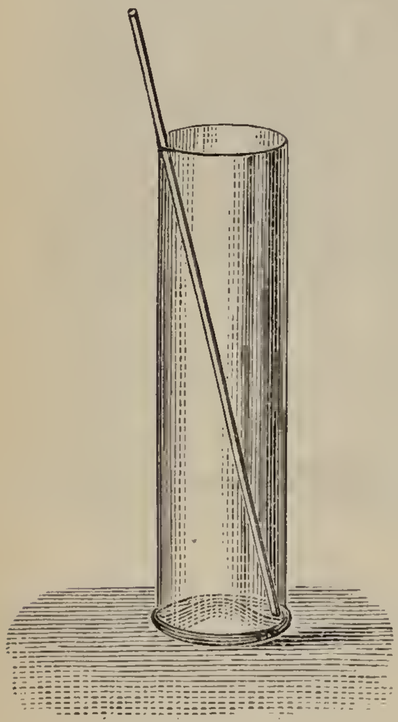


Fig. 39.

That ozone is nothing but oxygen, may be proved. By passing electric sparks through pure oxygen, ozone is formed. Now, electricity is not a kind of matter, and hence can add nothing to, nor can it take any thing from, the element oxygen. Ozone is therefore but another form of oxygen. Moreover, if left alone in the jar, ozone will in time return to the form of oxygen.

Vigorous as is the chemical action of ordinary oxygen, it is still more vigorous in the form of ozone. This is the chief difference in the action of these two forms of oxygen.

Ozone by Slow Combustion of Phosphorus. — Let a stick of phosphorus be laid on a plate with water deep enough to half immerse it. Then let a glass jar be inverted over it. White fumes arise from the phosphorus, due to its combination with oxygen, and, during this *slow combustion* of the phosphorus, a portion of the oxygen of the air is changed to ozone. Its presence may be shown by the starch paper.

By Electricity. — A silent discharge of electricity through air is able to convert a considerable portion of its oxygen into ozone. Indeed, a stronger ozone mixture can be obtained in this way than in any other. There is no way known in which oxygen can be wholly changed into ozone.

Allotropism. — Ozone is called an allotropic form of oxygen. Whenever the same element exists in two forms, they are said to be allotropic. And the property of an element in virtue of which it can show different properties under different conditions, is called ALLOTROPISM.

76. Density of Ozone. — Whenever oxygen is changed into ozone, it is reduced in volume; and when the ozone is changed back to oxygen again the original volume is restored. This proves that ozone is denser than oxygen. Experiments indicate that three volumes of oxygen form two volumes of ozone. If so, ozone is one and a half times heavier than oxygen, and its density is 24 $(H = 1)$.

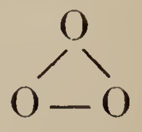
Explanation. — This can be best explained by supposing that the molecule of ozone contains three atoms, while that of oxygen contains only two. We should then have

$$\begin{array}{ccc} 3\,O_2 & = & 2\,O_3. \\ \text{Oxygen.} & & \text{Ozone.} \end{array}$$

That is, three molecules of oxygen produce two of ozone.

Thus the molecule of oxygen is . . . . $O = O$,

and the molecule of ozone is . . . .



It is believed that it is because the molecule of ozone contains this extra atom of oxygen, that ozone is able to convert substances into oxides more vigorously than can oxygen itself. This extra atom is easily given up to unite with other substances, with silver for example, while the two remaining atoms form a molecule of oxygen.

77. Occurrence. — Ozone occurs in the atmosphere. The quantity varies from time to time. It is often more abundant after a thunder-storm, being formed by the action of atmospheric electricity. It is a powerful agent in cleansing the atmosphere from organic impurities: it decomposes them, and converts their constituents into oxides.

II. — Oxygen and Hydrogen.

78. The Combustion of Hydrogen. — We have seen that hydrogen is a very combustible gas. Now, the product of its combustion is Water. This may be demonstrated by an experiment represented in Fig. 40. Hydrogen is set free in the bottle seen at the left in the picture. It passes

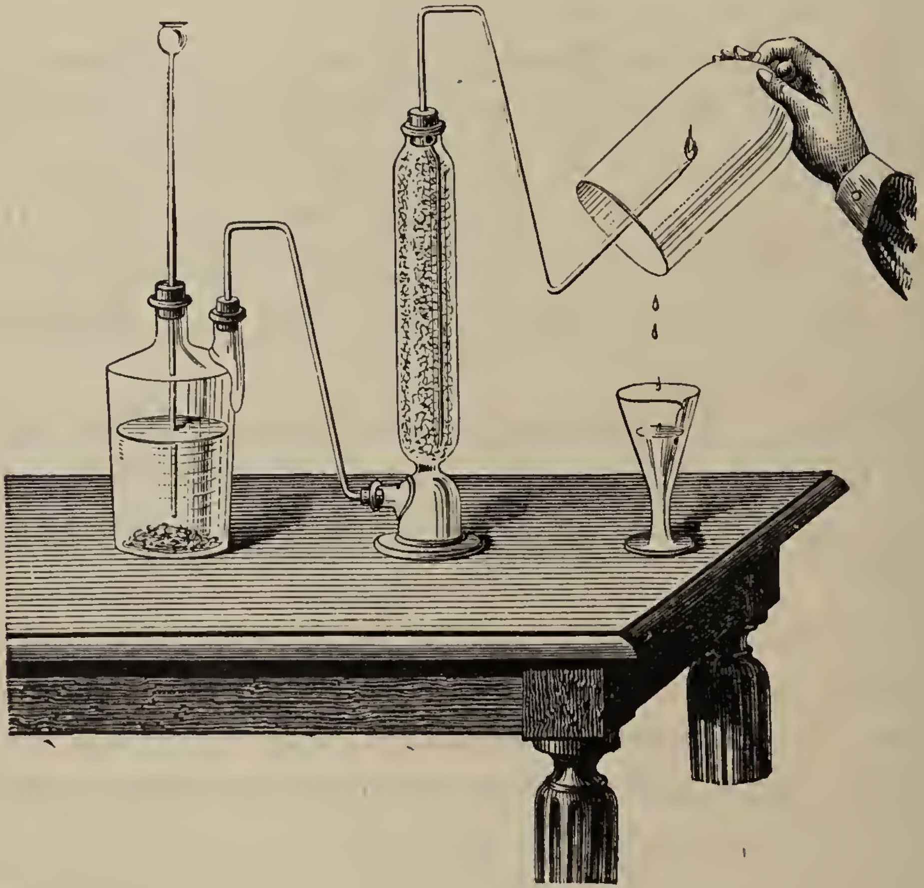


Fig. 40.

over through the bent tube to the bottom of a tall upright jar. This jar contains some *calcium chloride*, a substance which has a very strong attraction for water, but none for hydrogen. The gas, passing out at the top of the jar, will be *thoroughly dry*. Finally the hydrogen issues from the other end of the bent tube reaching from the top of the drying-jar, and is set on fire. A cold glass vessel is then brought down over the flame of burning hydrogen.

What is the result? Instantly the walls of the cold vessel are dimmed with dew! The dew rapidly accumulates; and, if the jar is kept cold, it collects into little globules. The little globules, grown larger and larger, at length trickle down the sloping sides of the vessel, and slowly drop from its mouth.

79. Properties of Water. — Water is a colorless, tasteless liquid. It evaporates at all temperatures, keeping the atmosphere charged with its invisible vapor, which, when condensed, produces fogs and mists, dews and rain. It boils at 100° C. (212° F.). At 15° C. (59° F.) water is 819 times heavier than an equal volume of air. Its greatest density is at a temperature of 4° C. (39° F.).

Freezing. — The freezing point of water is 0° C. (32° F.), and on freezing it expands nearly $\frac{1}{11}$ of its bulk. Ice is therefore lighter than water (.917) when equal volumes are compared.

Ice is *crystallized* water. In solid masses the crystalline form is obscured, but it is distinctly seen when the ice begins to form. It is beautifully exhibited in the frost figures often found upon the window-pane in winter, and also in the snowflakes when just fallen. (Fig. 41.)

80. Water as a Solvent. — A great many substances are soluble in water. A substance is said to be dissolved in water when its particles are so completely separated and scattered through the liquid as to be invisible. Its cohesion is entirely overcome by the stronger force of adhesion between its particles and those of water. A substance that may thus disappear is said to be *soluble;* and the fluid which contains it is called a SOLUTION, while that which dissolves it is called a SOLVENT. Salt is soluble in water: the brine is a solution, the water a solvent.

Fig. 41.

But it is well known that only a certain quantity of salt can be dissolved in water: all added beyond that, will remain undissolved. It is so with other solids: water can only dissolve a limited quantity. A solution in which the fluid has all it can hold of a soluble body is said to be SATURATED.

Effect of Heat. — The solvent power of water may be increased by heat. To this, however, there are some exception. Common salt, for example, will dissolve about equally in hot and cold water, and lime better in cold than in hot water.

Solubility of Gases. — Many gases, also, are soluble in water. Oxygen and nitrogen are examples: small quantities of both these gases, taken from the air, may be found in all natural waters. The presence of this dissolved air in water may be proved easily. We only need to place the liquid in a glass flask, and gently warm it. Little bubbles of gas will shortly be seen clinging to the bottom and sides of the vessel. More and more numerous and larger and larger they become, until, as the heat increases, they break away, and escape at the surface of the water. These bubbles are bubbles of air. It was in solution in the water, and the heat has driven it out. It may be caught in another vessel (Fig. 42) if desired.

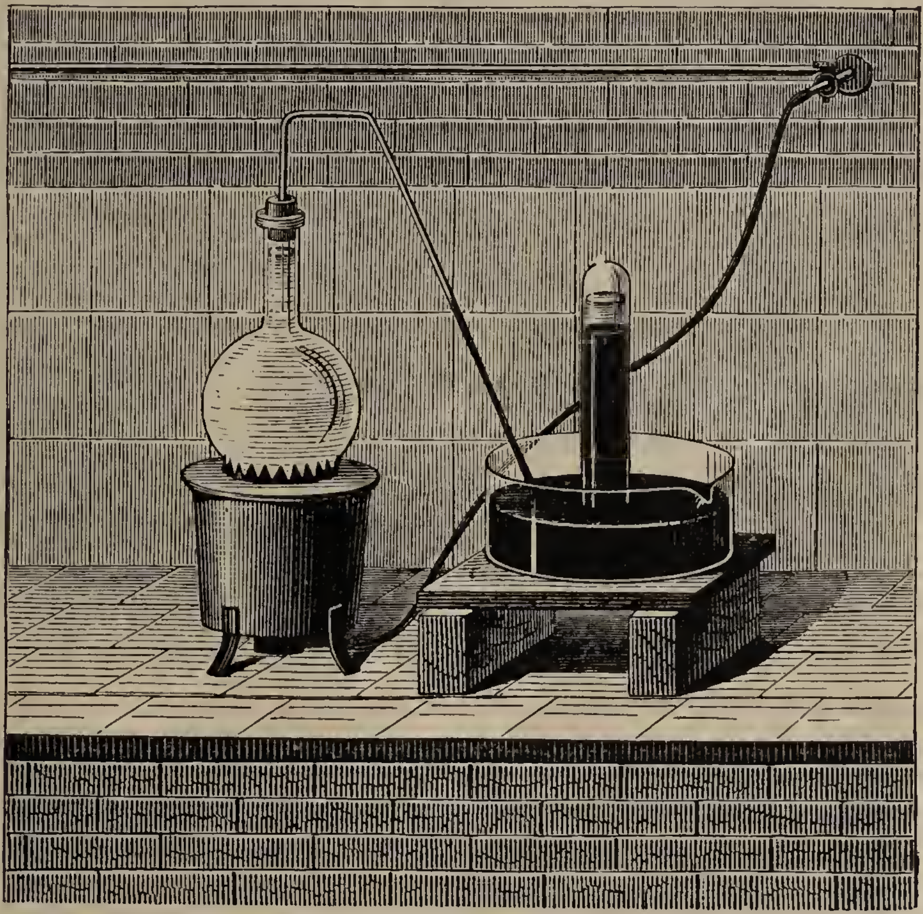


Fig. 42.

The experiment illustrates the fact that we may expel certain gaseous impurities from water by boiling.

Like solids, each gas has its own degree of solubility. Oxygen, for example, is more soluble than nitrogen, while carbon dioxide is much more soluble than either. Of this gas, water will dissolve about its own volume. Others are still more soluble: of ammonia gas, water at 0° C. will dissolve 1,148 times its volume.

81. Natural Waters. — Because water can dissolve so many other bodies, we may not expect to find pure water upon the earth. Coming from the clouds, it dissolves the gases of the atmosphere; sinking into the soil, it dissolves the minerals it meets; and flowing over rocks, it dissolves their materials. Hence the waters of all seas, lakes, rivers, and wells, are, without exception, impure; the kind and quantity of their impurities depending on the material over which they have passed.

Mineral Springs. — Salt springs are those whose waters have dissolved out the salt from the soil through which they have flowed. In another place some other substances may exist in the soil or rocks; and the water, dissolving these, forms other kinds of mineral springs.

The water may, for example, contain some salt of iron in solution, in sufficient quantity to give it an astringent taste and other special properties. In this case it is called CHALYBEATE WATER.

The Saltness of the Sea. — In a similar way we may account for the saltness of the sea. Common salt is scattered in small quantities through most soils, and is dissolved by water which trickles through them. This water collects in rivers, and finds its way to the sea. From the sea, water can only escape by evaporation. But, by this process, only pure water passes away: the salt is left behind. Age after age this work goes on, and large quantities of salt have thus already been accumulated in the sea.

Salt lakes are found, such as the Great Salt Lake of Utah. They are lakes without an outlet; so that, while the water may escape by evaporation, there is no escape for the salt.

Such Solid Impurities increase the Weight of the Water. — A very pretty experiment shows that brine is heavier than fresh water. In the first place, a small vial is partly filled with a colored liquid, so that when corked it will just sink in fresh water. Into a tall jar (Fig. 43), put fresh water several inches deep, and into this the little vial. Now by means of a long funnel-tube a saturated solution of salt may, with care, be poured to the bottom of the jar without mixing with the fresh water above. The fresh water will be lifted, floating on the brine, while the vial, also lifted, marks the dividing line between them.

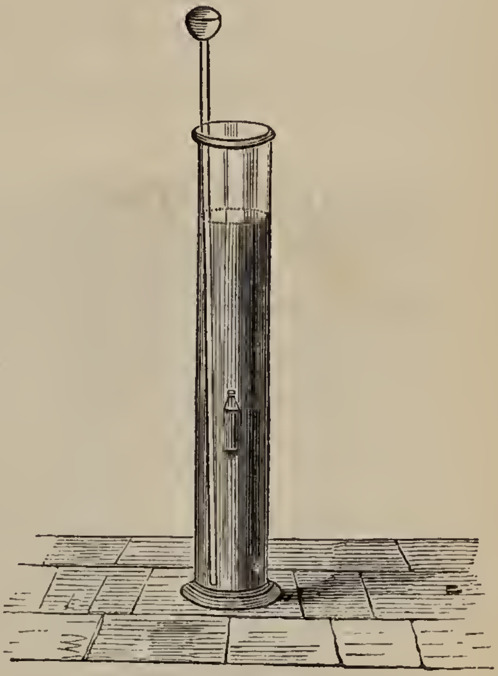


Fig. 43.

82. To Purify Water. — Water very often contains sediment, — solid matter not in solution, but held suspended in the liquid. By long standing this solid matter is likely to fall to the bottom. Nevertheless it is often desirable to separate it more perfectly or more quickly.

Fig. 44.

Filtration. — When a solid in fine division is mixed with a liquid, it may be separated by passing the liquid through some porous substance which will not let the particles of the solid pass. The process is called Filtration. The porous body through which the fluid passes is called a Filter, and the clear liquid which issues is called the FILTRATE. A filter of common form is made of unsized paper. Cut in the shape of a circle, it is folded so as to fit into a funnel (Fig. 44). The turbid fluid poured upon it filters through; and the clear liquid is caught in the vessel below, while the sediment is left upon the filter.

On a large scale water is freed from its suspended matter by passing it through filtering beds of sand or gravel.

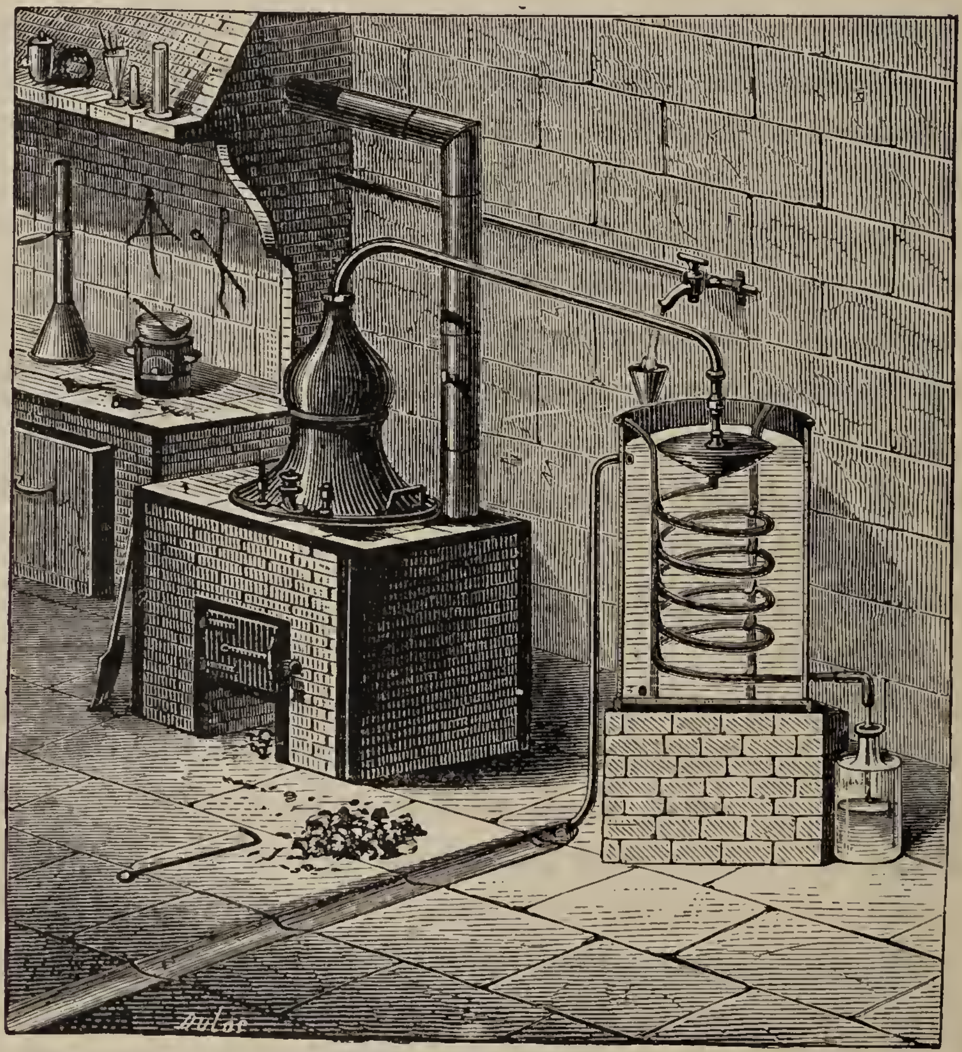


Fig. 45.

Distillation. — Impurities dissolved in water may be either fixed or volatile matter. When the water is boiled, only the gaseous impurities will pass away with the steam. The solid impurities will be left behind in the boiler, and if the steam be collected and cooled back into water, the fluid will be free from them. This process of boiling and then condensing the vapor is called DISTILLATION.

In Fig. 45, one form of a STILL, as the apparatus for distilling liquids is called, is represented. It consists of the boiler and the worm. The boiler is a close vessel standing over a furnace. The worm is a pipe, usually of tin, bent into a coil, and placed in a vessel which is kept full of cold water. The boiler and worm are connected by a steam-pipe.

The steam passing over from the boiler enters the worm, where, condensed by the low temperature, it becomes water again, and may be caught as it flows from the bottom.

To Remove Volatile Impurities. — By boiling water in an open vessel, such gases as air and carbon dioxide are driven away. But ammonia and organic substances, which may be present, cannot be removed so easily. It is necessary to put potassium permanganate and potassium hydrate with the distilled water, and then distill it a second time. These substances decompose the organic matter; and if the first portion of the water which comes over in this second distillation be thrown away, another may be collected which is likely to be *pure distilled water*.

By Freezing. — If the quantity of solid impurities in solution be small, water may be purified from them by freezing. The water only will become ice, while the impurities will be left in the fluid which remains unfrozen. On re-melting the ice a purer water is obtained.

But where the quantity of salt in the water is large, there is a different result. The fluid must be cooled to a much lower degree, and then a portion of the salt enters into combination with the water to form the crystals which are deposited. In the case of common salt the crystals begin to form at $-7°$ C., and they are composed of salt and water in proportion represented by

$$NaCl + 2H_2O.$$

On remelting these crystals, salt-water is obtained. Many other salts, in the same way, combine with ice; and these new compounds have been called CRYOHYDRATES.

83. Hard and Soft Water. — Water which contains compounds of lime or magnesia in solution is familiarly called HARD WATER. The magnesia may be present as a sulphate, and the lime as a carbonate, which is soluble in water containing carbon dioxide.

A water is hard when it requires much soap to make a lather. Now, soap is a *salt*, which contains a fatty acid combined with a base; and these salts of lime and magnesia decompose the soap, and combine with its acid constituent. The new compound is insoluble in water, and adheres to the surface of whatever is being washed.

To Soften Hard Water. — If the water is hard, because of the presence of lime carbonate, it may be softened by boiling; for, by boiling, the carbon dioxide is driven away, and the carbonate, no longer soluble, will settle as a precipitate. Or, by adding some "milk of lime" (calcium hydrate, CaH_2O_2, and water), the free carbon dioxide may be removed, and the water softened. If water is hardened by a sulphate, it can not be softened in these ways: its hardness is *permanent*.

84. Hydrogen Dioxide. — A second compound of hydrogen and oxygen, and the only one beside water, is H_2O_2, called HYDROGEN DIOXIDE. It is prepared by the action of dilute sulphuric acid on pure barium dioxide; thus, —

$$BaO_2 + H_2SO_4 = BaSO_4 + H_2O_2.$$

Barium dioxide + Sulphuric acid = Barium sulphate + Hydrogen dioxide.

The dioxide, H_2O_2, remains dissolved in water, but by carefully evaporating the water the dioxide may be obtained.

85. Properties. — Hydrogen dioxide is a colorless, oily liquid, almost one and a half (1.452) times heavier than water. It has no smell, but it has an astringent taste. If in contact with the skin, it raises a white blister. It destroys vegetable colors.

This compound is very unstable, decomposing spontaneously at 15° C., and at 100° C. so rapidly as sometimes to produce explosion. The products are water and oxygen. The presence of finely divided silver or gold decomposes it with great violence, while, curiously enough, the metal itself remains unchanged.

86. Hydroxyl. — Hydroxyl is the name given to *half a molecule* of the hydrogen dioxide: the formula for the dioxide being H_2O_2, that of hydroxyl is HO.

This group of atoms, HO, is found in a large number of compounds. Thus we find it in sodium hydrate, $NaHO$; in potassium hydrate, KHO; in calcium hydrate, $Ca(HO)_2$; and in alcohol, C_2H_5HO.

It is believed that in a great many chemical changes this group remains unbroken. The hydrogen and the oxygen atoms cling together, and go together into combination, and stay together in the new molecule. The group is univalent. This is shown by its constitutional formula, H—O—.

III. — OXYGEN AND CHLORINE.

87. Oxides and Acids of Chlorine. — The chemical attraction between chlorine and oxygen is very feeble. Still there are three oxides of chlorine known, — the monoxide Cl_2O, the trioxide Cl_2O_3, and the peroxide ClO_2.

The first two of these oxides combine with water to form acids; thus: —

$$Cl_2O + H_2O = 2\,HClO,$$
Chlorine monoxide + Water = Hypochlorous acid.
$$Cl_2O_3 + H_2O = 2\,HClO_2,$$
Chlorine trioxide + Water = Chlorous acid.

In addition to these, there are two other acids of chlorine:

Chloric acid	$HClO_3$,
Perchloric acid	$HClO_4$.

88. Chloric Acid. — In this series of seven compounds, the chloric acid is the most important. Even this one is important, only because it yields a valuable class of salts, — the chlorates. Potassium chlorate is an example.

Hypochlorous Acid. — Next in importance is the hypochlorous acid, which is of interest chiefly because it is a constituent in "bleaching-powder." This useful substance contains calcium hypochlorite, $Ca\,(Cl\,O)_2$, and its bleaching power depends on this substance. The addition of hydrochloric or sulphuric acid liberates its chlorine, and this chlorine then attacks the coloring matter.

All these compounds of chlorine and oxygen are very unstable bodies, decomposing easily, and some of them spontaneously, often with explosive violence.

REVIEW.

I. — SUMMARY OF PRINCIPLES.

89. Oxygen may be obtained in many ways. It is usually obtained by heating potassium chlorate, $K\,Cl\,O_3$.

It is produced by heating mercuric oxide, $Hg\,O$. Also by the electrolysis of water.

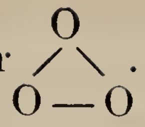
Oxygen is colorless, tasteless, odorless, a little heavier than air, and slightly soluble in water.

It enters into combination with all the elements, with the possible exception of fluorine.

Substances which burn in air burn with increased vigor in oxygen, and many substances which will not burn in air will burn when heated in this gas.

The combining weight of oxygen is 16: its molecular weight is 32: its molecular formula is O_2 or $O=O$.

Oxygen exists in an allotropic form, called Ozone. In this condition it attacks and combines with many things for which, in its ordinary state, it has no affinity.

The molecular formula for ozone is O_3, or . Having three atoms in the molecule instead of two, its density is $\frac{3}{2}$, or 1.5, that of oxygen.

The compounds of oxygen are very numerous. Whether produced by common oxygen or by ozone, their composition is the same: they are called oxides.

Oxygen is the most abundant element in nature.

Oxygen and hydrogen form two compounds, viz.: Water or Hydrogen Oxide, H_2O, and Hydrogen Dioxide, H_2O_2.

Water is the sole product of the combustion of pure hydrogen.

Water consists of hydrogen and oxygen in the proportion of 1 : 8 by weight. By volume the proportions are two of hydrogen to one of oxygen.

The graphic formula for water is H—O—H. An atom of oxygen links together two atoms of hydrogen to form a molecule.

The power of water as a solvent is very great. Heat generally increases the solubility of solid bodies in water; it diminishes the solubility of gases.

Rain-water, caught after much has fallen, is the purest form of water to be found in nature.

All spring-water contains soluble salts and gases. If these substances are present in sufficient quantity to impart special properties to the water, it is called mineral water.

Water may be purified by filtration or by distillation: by filtration it may be freed from sediment; by distillation it may be freed from all solids and most of its gaseous impurities.

The presence of salts of lime and magnesia renders water unable to dissolve soap readily; such waters are called hard waters.

Hydrogen dioxide is a colorless liquid, nearly one and a half times heavier than water. It readily decomposes into oxygen and water.

Its graphic formula is H—O—O—H.

Hydroxyl is a compound of one atom each of hydrogen and oxygen. Its graphic formula is H—O—. One of the bonds of the oxygen atom is free. In other words, the molecule is univalent.

This group, H O, is found in a large number of compounds.

Oxygen and chlorine form three oxides of chlorine. There are also four acids with chlorine.

All these compounds are very unstable.

Chloric acid is important because it forms a large and useful class of compounds, the Chlorates.

Hypochlorous acid is important because some of its salts are bleaching agents. The calcic hypochlorite is the bleaching agent in " bleaching powder," often called " chloride of lime."

II. — EXERCISES.

Describe the preparation of oxygen. Write the reaction.

What weight of oxygen will be furnished by one thousand grams of potassium chlorate?

How much by volume, the density of oxygen being 16?

What weight of potassium chlorate must be used to obtain ten liters of oxygen?

What volume of oxygen may be obtained from one hundred grams of the mercuric oxide, Hg O?

What are the physical properties of oxygen? What is the chemical action when carbon burns in oxygen? When iron burns in oxygen? What is the relation of oxygen to animal life? What class of compounds does oxygen form?

Give the atomic weight of oxygen. Its atomic volume. Its molecular weight. Its molecular volume. Its density.

Where does oxygen occur in nature?

How may ozone be produced? By what test may it be recognized? In chemical character how does it differ from oxygen?

Define allotropism. What evidence that ozone is an allotropic form of oxygen? What is the density of ozone? What explanation is given?

What is the product when hydrogen is burned?

What is the composition of water by volume? By weight? What is its freezing point? Its boiling point? What is the temperature of water at its greatest density?

Define solution. Solvent. What is a saturated solution?

What is the effect of heat upon solubility?

What are mineral waters?

Define filtration. Evaporation. Distillation. Filtrate. Distillate. Describe the process of distillation.

What are cryohydrates?

What are "hard waters"? How may hard water be softened?

What is the composition of hydrogen dioxide? How may it be obtained? What are its properties? What are the products of its decomposition?

What is hydroxyl? In what is this group of atoms found? What is its quantivalence?

Name the oxides of chlorine. How many acids of chlorine are known? Which of these are important? What is bleaching-powder?

IV. — SULPHUR.

90. Preparation. — By far the greater part of sulphur in commerce has been obtained from the "native sulphur," such as is found in Sicily, simply by the application of heat. The sulphur is vaporized, and, passing over into large cold chambers, the vapor is again condensed. In this way sulphur in the finest powder, known in commerce as FLOWERS OF SULPHUR, is obtained.

When smaller chambers are used, their walls soon become so heated by the hot vapors, that the sulphur is kept in a melted state. This liquid, drawn off into molds and cooled, forms what is known as ROLL BRIMSTONE.

91. Properties. — Sulphur is a solid at ordinary temperatures. It is tasteless and without odor.

The effects of heat on this element are very unusual. Let a small quantity of sulphur be laid on a piece of writing-paper, and held over the flame of a candle; in a little time it melts to a clear, yellow liquid. The melting occurs at about 115° C. Again, put more of it into a test-tube, and heat it gradually. After melting, it remains a clear, limpid liquid up to about 132° C., and then begins to get thick and dark-colored. At about 250° C. it is so viscid that it can scarcely be poured from the tube; but, as the heat increases, it becomes less viscid again, and finally, at about 450° C. it boils.

If, when heated to almost the boiling point, the mobile liquid is poured into cold water, it becomes curiously unlike common sulphur; it cools into a dark-colored solid, with a considerable degree of *elasticity*. In this condition it is called PLASTIC SULPHUR. On standing, it soon becomes yellow and brittle again.

92. Crystalline Forms. — Sulphur may be obtained in crystals, either by melting it or by dissolving it.

If we fill a test-tube one-half full of sulphur, and melt it, and then let it cool slowly, we may soon see the formation of crystals, which grow from the cool sides of the tube toward the center of the liquid. These crystals will multiply until they crowd one another more and more so that the whole finally becomes a compact mass.

If we desire to keep the crystals separate, we may melt the sulphur in a beaker, let it cool slowly until a good crop has formed, then pierce the crust on top, and pour off the remaining liquid. Fine transparent needle-shaped crystals of sulphur will line the walls of the beaker.

These experiments illustrate the production of crystals by the *method of fusion*. Crystals may also be obtained by the *method of solution*.

Crystals by Solution. — Sulphur is not at all soluble in water, but it dissolves easily in carbon disulphide. The carbon disulphide is very volatile: it rapidly evaporates on exposure to air, and deposits the sulphur in the form of crystals. These crystals are not needle-shaped like those obtained by fusion: they are rhombic octahedrons instead.

Sulphur is not the only substance which is *able to crystallize in two distinct forms.* Such substances are said to be DIMORPHOUS.

93. Chemical Properties. — Sulphur has a wide range of attractions. It unites with oxygen and with hydrogen to form important acids. With the metals it forms a numerous class of compounds called SULPHIDES. These were formerly called *sulphurets.*

94. Sulphur in Nature. — In some volcanic districts sulphur is found free. The mines of sulphur on the island of Sicily contain the sulphur mixed with earthy matter, and much more rarely in the form of very pure and beautiful crystals.

Sulphur is found in combination with metals in the earth almost everywhere. Iron pyrites, so common and familiar, is a disulphide ($Fe\ S_2$). In combination with hydrogen it is found in the water of what are called sulphur-springs. And, besides all this, sulphur is an important element in many animal and vegetable substances.

95. Its Uses. — Sulphur is very largely used in the arts. It is a constituent of gunpowder and of sulphuric acid. It is used in the manufacture of friction-matches, in medicine also; and in its elastic state, produced under the influence of heat, it is used in taking casts of coins and medals.

V. — SULPHUR AND HYDROGEN.

96. Sulphuretted Hydrogen. — Two compounds of sulphur and hydrogen are known. They are represented by the formulas H_2S and H_2S_2. The first of these, hydrogen sulphide, or, as it has long been called, sulphuretted hydrogen (H_2S), is an indispensable substance in the laboratory, and it is that which is dissolved in the waters of the so-called sulphur-springs, giving to them their odor, taste, and medicinal qualities.

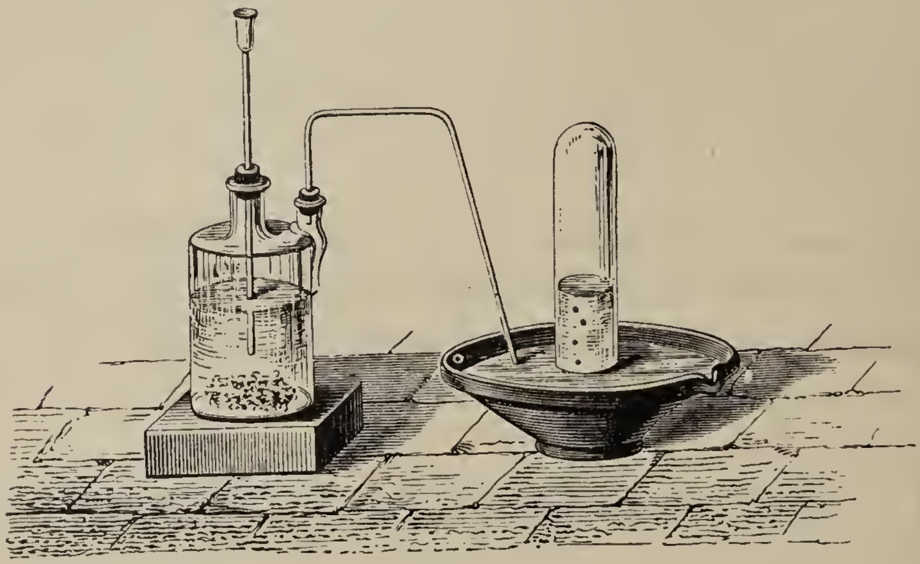


Fig. 46.

97. Preparation. — Fragments of ferrous sulphide (FeS) are placed in a bottle (Fig. 46), which is provided with a tightly fitting cork and two tubes, one a funnel-tube reaching nearly to the bottom, the other a bent tube reaching from the cork over into the water of a cistern. Dilute sulphuric acid is poured through the funnel-tube. Chemical action at once begins, and the sulphuretted hydrogen is set free.

$$\begin{array}{ccccccc} FeS & + & H_2SO_4 & = & FeSO_4 & + & H_2S, \\ \text{Ferrous sulphide} & + & \text{Sulphuric acid} & = & \text{Ferrous sulphate} & + & \text{Sulphuretted hydrogen.} \end{array}$$

98. Properties. — Hydrogen sulphide is a colorless gas, having a powerful and unpleasant odor, resembling the odor of rotten eggs. Very small proportions of the gas mixed with air render it unpleasant; while the gas itself, when inhaled, acts as a powerful poison. Very dilute chlorine is an antidote to this poison. Should an antidote ever be required, then moisten a towel with dilute acetic acid, sprinkle a few grains of bleaching-powder upon it, fold the powder inside, and apply the outside to the nose. The acid liberates the chlorine, and the chlorine attacks the hydrogen sulphide.

Reaction with Metals. — Sulphuretted hydrogen is very readily decomposed by many metallic compounds, whose metals are in this way converted into sulphides. Thus, —

$$Cu\,SO_4 \quad + \quad H_2S \quad = \quad H_2SO_4 \quad + \quad Cu\,S,$$

Copper sulphate + Hydrogen sulphide = Sulphuric acid + Copper sulphide.

This power to convert metals into sulphides renders hydrogen sulphide an indispensable reagent in the laboratory.

Solubility. — This gas is freely soluble in cold water. The solution is easily prepared by passing the gas through successive bottles of water, arranged as represented in Fig. 47. The sulphide still retains its odor and its chemical power, and the solution can be used for chemical purposes instead of the gas.

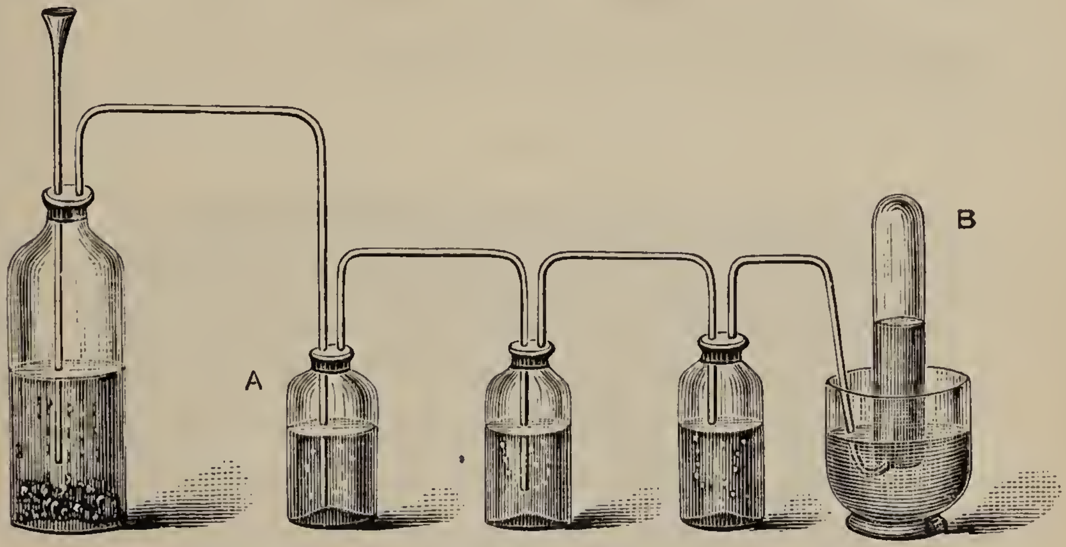


Fig. 47.

VI. — Sulphur and Oxygen.

99. Oxides and Acids. — Sulphur and oxygen form two oxides : —

Sulphurous oxide	SO_2,
Sulphuric oxide	SO_3.

By combining with the elements of water, these yield two very important acids: —

Sulphurous acid	H_2SO_3,
Sulphuric acid	H_2SO_4.

Oxides which form acids by uniting with water are often called ANHYDRIDES. The SO_2 may be called sulphurous anhydride, and the SO_3, sulphuric anhydride.

$$\begin{array}{ccccc} SO_2 & + & H_2O & = & H_2SO_3, \\ \text{Sulphurous anhydride} & + & \text{Water} & = & \text{Sulphurous acid.} \end{array}$$

In addition to the two acids just named, six others are known. One of these is thiosulphuric acid: it was formerly known as hyposulphurous acid. Its salts, then known as hyposulphites, are now called thiosulphates instead. Sodium thiosulphate is an example, and a useful substance.

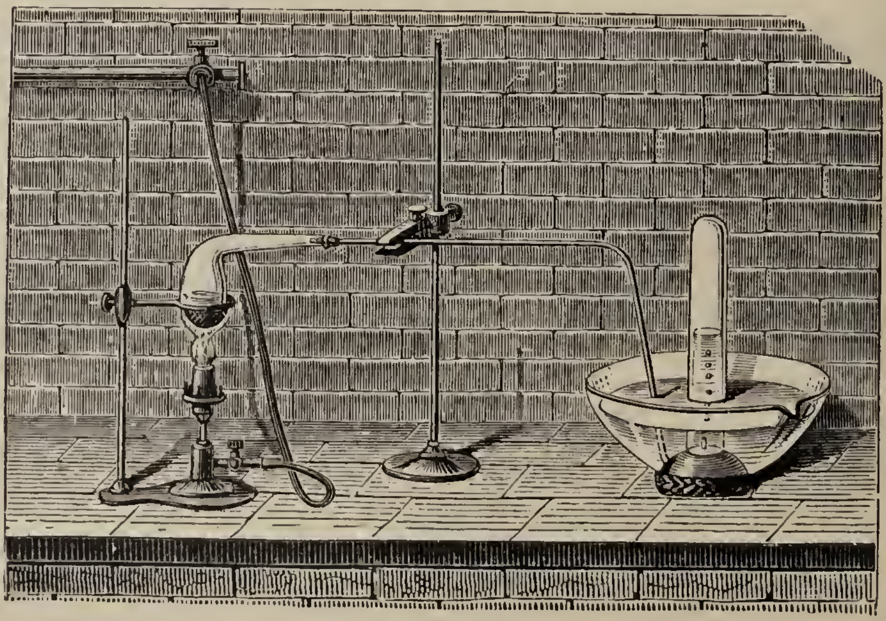


Fig. 48.

100. Sulphurous Oxide. — Set fire to a piece of sulphur in a deflagrating spoon, and thrust it into a bottle of air (Fig. 3). The sulphur burns, combining with the oxygen, and forms sulphurous oxide.

To obtain this gas in a purer state, we may treat copper with sulphuric acid. The operation may be conducted in an apparatus shown in Fig. 48. The copper and acid are brought together in a retort, and heated. The gas may be collected over mercury. Three new substances are produced: viz., copper sulphate, water, and sulphurous oxide, thus: —

$$\underset{\text{Copper}}{Cu} + \underset{\text{Sulphuric acid}}{2H_2SO_4} = \underset{\text{Copper sulphate}}{CuSO_4} + \underset{\text{Water}}{H_2O} + \underset{\text{Sulphurous oxide.}}{SO_2},$$

101. Properties. — Sulphurous oxide is a colorless gas, with a very stifling odor. It is very soluble in water, but when dissolved it is no longer sulphurous oxide: it has combined with water, and has become sulphurous acid. This is easily demonstrated by pouring a little water into a bottle in which sulphur has been burned, and then adding a little of this liquid to a beaker of blue litmus solution. The reddening of the litmus, which immediately occurs, declares the acid character of the solution.

102. Bleaching. — Both the gas and its solution have power to remove colors. If, for example, a brightly tinted flower be held over burning sulphur, its color disappears.

This action is employed in the arts for bleaching silk, woolen, and straw goods. For example, if a milliner wishes to whiten her straw, she may hang it in a small chamber, which may be nothing more than an inverted barrel, and then burn some sulphur under it. The sulphurous oxide enters the straw, and discharges its color.

On a large scale the slightly moistened articles are hung in large chambers in which sulphur-fires are kindled.

103. A Reducing Agent. — Sulphurous oxide takes oxygen out of many other substances, and *reduces* them to simpler forms of combination. If, for example, we fill a bottle with the gas, by burning in it some sulphur, and then thrust into it a glass rod, or, better, a shaving of wood, moistened with strong nitric acid, dark red fumes will soon make their appearance. These red fumes show that a compound of nitrogen and oxygen, NO_2, is set free.

The chemical action may be written as follows: —

$$2HNO_3 + SO_2 = H_2SO_4 + NO_2,$$

Nitric acid + Sulphurous oxide = Sulphuric acid + Nitrogen peroxide.

A substance which readily takes oxygen out of others is called a Reducing Agent. Sulphurous oxide is valuable to the chemist as a reducing agent.

Nitric acid, on the other hand, readily imparts oxygen to other substances. In the above illustration it gives oxygen to the SO_2, and *oxidizes* it to SO_3, which then combines with $2H_2O$ to produce the H_2SO_4. A substance which readily parts with oxygen to other substances, as nitric acid does, is called an Oxidizing Agent. Nitric acid is valuable to the chemist as an oxidizing agent.

104. Preparation of Sulphuric Acid. — Sulphuric acid is made by oxidizing sulphurous oxide in presence of water. The oxidizing agent employed is nitric acid, or a compound of nitrogen and oxygen such as the nitrogen peroxide, NO_2.

On a large scale the sulphurous oxide is obtained by burning sulphur, or iron pyrites (FeS_2) which yield sulphur when heated. This sulphurous oxide is carried over into immense chambers, which are made of sheet-lead. As it enters the first lead chamber, it meets with nitric acid, which falls through it in fine streams, and also with steam, which is blown in from a boiler. The sulphurous oxide is changed into sulphuric acid, which condenses on the walls of the chamber, and trickles down to the floor, where it remains dissolved in water.

The nitric acid is thus reduced to nitrogen monoxide, NO; but this unites at once with the oxygen of the air in the chamber, and becomes nitrogen peroxide, NO_2.

This nitrogen peroxide gives half its oxygen to change another portion of SO_2 to SO_3, which, in presence of steam, becomes H_2SO_4, and is itself again reduced to NO.

Again, the NO becomes NO_2 by taking oxygen from air; and, again, the NO_2 oxidizes another portion of SO_2. In this way the NO is a carriage which conveys oxygen from the air to the sulphurous oxide, continually.

The fluid which collects on the floors of the lead chambers is very dilute sulphuric acid, having a specific gravity of 1.55.

Oil of Vitriol. — The dilute acid is made stronger by drawing it off into leaden pans, and heating it. The water passes off as steam, and the acid remains. In this way the acid is concentrated until it contains only twenty-two per cent of water, or has a specific gravity of 1.71. In this form it is called OIL OF VITRIOL. Put up in large bottles, incased in wood, called *carboys*, it is sold in great quantities.

Oil of vitriol always contains lead sulphate taken from the lead pans. To avoid this impurity the concentration is often made in vessels of glass or of platinum. Even then the acid may contain other impurities, especially arsenic, which comes from the pyrites employed. To obtain pure sulphuric acid, the commercial acid is distilled in glass retorts.

105. Properties. — Pure sulphuric acid is a colorless, very heavy, and oily liquid. The usual brown color of the "oil of vitriol" is due to a little organic matter which it contains.

Its attraction for water is one of the most remarkable characters of this acid. Left in an open bottle, it speedily takes moisture from the atmosphere, and increases in volume.

The combination of the acid with liquid water is a source of heat. Let a little water be put into a beaker, and about four times as much strong acid added. The beaker will be found too hot to be handled with convenience; and if a test-tube containing alcohol, or even a little water, be placed in the mixture, its contents will boil vigorously.

Many organic bodies are decomposed by sulphuric acid.

They are composed of carbon, hydrogen, and oxygen; and the acid seizes their hydrogen and oxygen in the proportions which form water, leaving the carbon behind. A piece of white pine wood is almost instantly blackened in this way by contact with the acid, and paper is also charred as if by fire. Let about 50 centimeters of the strong acid be mixed with the same volume of a thick sirup of sugar in a tall jar: the chemical action is violent, and the sugar is converted into a bulky mass of porous charcoal.

106. Uses. — Sulphuric acid is regarded as the most useful acid known. Its manufacture is one of the important branches of industry. Probably more than a million tons is made yearly in Great Britain and the United States.

It is used in the manufacture of almost all other chemicals, in the preparation of fertilizers, in the art of bleaching, in dyeing, and in calico-printing.

107. The Sulphates. — By action on the metals, sulphuric acid forms a large class of compounds, called SULPHATES. Both atoms of hydrogen in its molecule (H_2SO_4), or one alone, may be displaced by atoms of metal, and hence *two classes* of sulphates exist. Thus with sodium we may have Na_2SO_4, or $NaHSO_4$. The first is the NORMAL sodium sulphate; the second is the ACID sodium sulphate. A normal salt is one in which all the replaceable hydrogen of the acid is removed. An acid salt is one in which only a part of the replaceable hydrogen of the acid is removed.

The hydrogen of an acid which can be replaced by metals is called BASIC hydrogen. And sulphuric acid is said to be a DIBASIC acid, because its molecule contains two atoms of hydrogen, for which metallic atoms may be substituted.

Most of the sulphates are freely soluble in water. Lead sulphate and a few others are but little soluble, while barium sulphate is quite insoluble.

108. Tests. — If a little barium chloride is added to a solution containing a sulphate, or sulphuric acid itself, the insoluble barium sulphate will be invariably produced. It will appear as a white precipitate, which will not dissolve in hydrochloric acid. The appearance of this precipitate declares the presence of the acid, or its compounds, in the suspected liquid.

Free sulphuric acid may be detected by its charring organic bodies. Vinegar has been known to contain this acid. To detect the adulteration, add a little white sugar, and then evaporate the vinegar to dryness at a gentle heat. If a black residue remains, it indicates the presence of the acid.

VII. — SELENIUM AND TELLURIUM.

109. Selenium. — Selenium is a solid element found in small quantities, usually in combination, as selenides of metals. Copper, lead, and silver selenides are examples. The element exists in two allotropic forms. In one it is soluble in carbon disulphide ($C S_2$), in the other it is not.

The relations of selenium to electricity are most curious and important. When cooled suddenly from a melted state, it is very brilliant, dark-colored, and a *non-conductor* of electricity. When cooled slowly from fusion, it has a dull lead-colored surface, is crystalline, and is a *conductor of* electricity. Its power to conduct is increased by heat. So exceedingly sensitive is it, that even the heat of a candle-flame, at a distance, will diminish its resistance to the passage of the current. Even when the radiation from the candle passes through a layer of water, it still affects the selenium, indicating that the element is affected not only by heat but by light alone, since water transmits the light but absorbs heat.

110. Application. — The rapidity with which selenium will respond to the touch of light is remarkable. Let the light be flashed on and off in quick succession, and the resistance of the selenium changes with equal rapidity.

This principle has been applied by Graham Bell in his *photophone.* Light is thrown upon a mirror which reflects it to a distant selenium receiver, lenses being employed to concentrate it. This receiver is placed in circuit with a battery and a Bell telephone. If the voice is thrown against the back of the mirror, the mirror will vibrate and put the beam of light in motion. The motion of this light on the surface of the selenium varies its resistance. The current of electricity passing through it is thus thrown into undulation. These undulations of the current correspond to those of the voice, and affect the ear at the telephone in the same way. Thus has this property of selenium been made useful in the transmission of sound by light.

111. Tellurium. — Tellurium is even more rare than selenium. When pure this element is bluish-white and lustrous. In physical properties it resembles the metals, but its chemical actions are much like those of selenium and sulphur.

VIII. — THE BIVALENT OR OXYGEN GROUP.

112. Chemical Resemblance of these Elements. — Oxygen, sulphur, selenium, and tellurium form a well-marked natural group. They enter into combination with the same substances, and their compounds have similar composition.

This analogy is seen in their hydrogen compounds. Thus: —

$$H_2O, \quad H_2S, \quad H_2Se, \quad H_2Te.$$

Their bivalent character is clearly shown by these formulas.

The last three elements of the group form analogous compounds with oxygen. Notice the formulas: —

$$\begin{array}{lll} SO_2, & SeO_2, & TeO_2, \\ SO_3, & SeO_3, & TeO_3, \end{array}$$

in which this resemblance clearly appears.

By combining with water, these oxides all form acids. Thus: —

$$\begin{array}{cccccc} SeO_2 & + & H_2O & = & H_2SeO_3, \\ \text{Selenium dioxide} & + & \text{Water} & = & \text{Selenious acid.} \end{array}$$

REVIEW.

I. — SUMMARY OF PRINCIPLES.

113. Sulphur is found in nature uncombined.

This element is a yellow solid, at ordinary temperatures very fusible, easily vaporized, but quite insoluble in water.

An allotropic form of sulphur is produced by heating it to near its boiling-point, and suddenly cooling it by cold water. In this condition it is dark-colored and elastic.

The compounds of sulphur are very numerous and important.

With hydrogen it forms the useful but unpleasant "sulphuretted hydrogen" or hydric sulphide, H_2S.

This gas is readily decomposed by many salts: their metals take the sulphur, and become sulphides.

With oxygen sulphur forms two oxides, or anhydrides, SO_2 and SO_3; and these unite with water to form the sulphurous and the sulphuric acids.

Both sulphurous oxide and sulphurous acid are good bleaching agents. They are used for bleaching silk, woolen, and straw goods.

Sulphurous oxide is a reducing agent.

Sulphuric acid is obtained by burning native sulphur, or iron sulphide, and oxidizing the sulphurous oxide thus formed by means of nitric acid and the nitrogen oxides.

This acid is the most useful among chemicals. It is a heavy, oily liquid, having a strong attraction for water. It combines directly with water, and also decomposes many organic substances in order to obtain their hydrogen and oxygen, which it takes from them in the proportions which form water.

This acid is dibasic. Its salts are sulphates, and these are either normal sulphates or acid sulphates.

The barium sulphate is the most insoluble in water or acids. It is formed whenever a barium compound is added to a solution of any sulphate. It appears as a white precipitate. This is a test for the presence of sulphuric acid, either when it is free or in combination.

Selenium and tellurium are rare elements, much resembling sulphur in their chemical actions.

Oxygen, sulphur, selenium, and tellurium are bivalent, and constitute a well-marked group, known as the oxygen or bivalent group of non-metals.

II. — EXERCISES.

Describe the preparation of sulphur. Describe the effect of heat on sulphur.

In what two ways may crystals be obtained? How may we obtain crystals of sulphur by fusion? How may we obtain crystals of sulphur by solution? Why is sulphur said to be a dimorphous substance? What class of compounds does sulphur form? How does sulphur occur in nature? What are some of the uses of this element?

What is sulphuretted hydrogen? How may it be prepared? Write the reaction. What are the properties of this substance? What is the chemical action of this gas on metals? Write the reaction with $Cu\,S\,O_4$.

Name the oxides of sulphur. Give their composition. Name the corresponding acids. Define anhydride. Show by the equation that the addition of the elements of water converts these oxides into acids.

What compounds of chlorine are anhydrides?

How may sulphurous oxide be obtained? Describe the gas. For what purpose is it useful?

What is a reducing agent? What is an oxidizing agent? Show by the equation how SO_2 reduces nitric acid.

How is sulphuric acid made?

Describe the process as conducted on a large scale. What is oil of vitriol?

What is the result of bringing sulphuric acid and water together? What action does H_2SO_4 exert on organic bodies? What are some of its uses?

What class of compounds does sulphuric acid produce? What is a normal salt? What is an acid salt? What is a dibasic acid? A monobasic acid? A tribasic acid? What is basic hydrogen?

What is the test for sulphuric acid? How would you detect this acid in vinegar?

Name the members of the bivalent group of non-metals. Give the formulas of their hydrogen compounds. Give the anhydrides of sulphur, selenium, and tellurium.

SECTION V.

THE TRIVALENT NON-METALS.

I. — Nitrogen.

114. Preparation. — There are large quantities of nitrogen in the atmosphere, but there are large quantities of oxygen with it. By burning phosphorus in a portion of air, the oxygen will be taken away, and the nitrogen left. For this purpose let a piece of phosphorus the size of a large pea be placed on a cork floating upon the water of a cistern. Touch it with a hot iron, and quickly invert over it a gallon jar. (See Fig. 49.) The phosphorus burns with a beautiful light, while milk-white vapors fill the jar. These vapors will be gradually absorbed by the water which will rise into the jar. The space above the water at last is filled with nitrogen.

115. Its Physical Properties. — The nitrogen is now seen to be a gas, perfectly colorless and transparent. It is without odor or taste, and a little lighter than air, its specific gravity (air = 1) being .972. It is slightly soluble in water, one hundred volumes of water absorbing about 1.5 volumes at ordinary temperature.

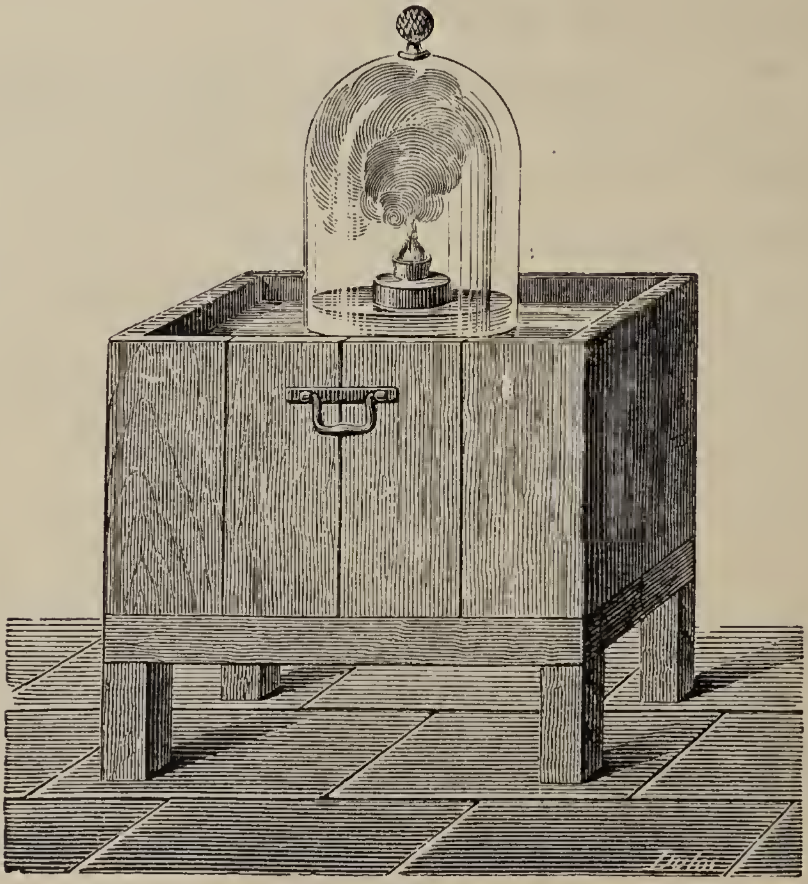


Fig. 49.

116. Its Chemical Properties. — If a lighted taper be lowered into a jar of nitrogen, it will be extinguished as quickly as if plunged into water. The gas will neither burn, nor allow other things to burn in it. So, too, if an animal were put into this gas, death would soon follow. It is not poisonous, but kills simply because it has no power to support life. A bandage over the face, which would shut all air away from the lungs, would kill in the same way. These facts illustrate the feeble disposition of nitrogen to combine with other elements. It can be made to combine with many of the elementary bodies; but the compounds are for the most part unstable, that is, easily decomposed. On this account the compounds of nitrogen are among the most energetic substances in chemical actions.

The atomic weight of nitrogen and its density, hydrogen being 1, are 14.

117. Occurrence in Nature. — Nitrogen is a very abundant element. It forms about four-fifths by weight of all the atmosphere, and besides this it is an important constituent in many animal and vegetable bodies.

II. — Nitrogen and Hydrogen.

118. Ammonia. — Nitrogen unites with hydrogen to form a single compound: it is the hydrogen nitride, or, as it is commonly called, ammonia.

Ammonia exists in small quantities in the atmosphere, not free but in combination as a carbonate and some other salts. Its compounds occur in small quantities in rain-water, and are to be found also in all fertile soils.

Nitrogen is a constituent in many of the substances which compose animal and some vegetable bodies, and by their putrefaction and decay ammonia is set free.

In this case the nitrogen and hydrogen are brought together just at the moment when they are liberated from combination. At this moment they unite in ammonia. Substances at the moment when set free from combination are said to be in the Nascent State. Their chemical activity seems to be greater at this time than at others: the hydrogen and nitrogen, for example, which readily unite in the nascent state, show no disposition at all to unite when simply mixed together.

119. Preparation. — Ammonia is usually obtained by heating together some ammonium chloride and "slaked lime."

The chemical change is as follows: —

$$\begin{array}{ccccccccc} 2NH_4Cl & + & Ca(OH)_2 & = & CaCl_2 & + & 2H_2O & + & 2NH_3, \\ \text{Ammonium chloride} & + & \text{Lime} & = & \text{Calcium chloride} & + & \text{Water} & + & \text{Ammonia.} \end{array}$$

That is: two molecules of ammonium chloride with one of calcium hydrate, yield one molecule of calcium chloride, two of water, and two of ammonia.

120. Physical Properties. — The ammonia obtained is a gas, having a well-known pungent odor. It is much lighter than air (.586). But its most remarkable physical property is its solubility in water; one volume of water at 0° C. absorbing no less than 1,148 volumes of the gas. As in the case of all gases, its solubility is less at higher temperatures, but at 15° C. one volume of water still dissolves 783 volumes of ammonia. Its solubility may be shown by experiment, using the same apparatus as for hydrochloric acid. The lower bottle should contain a little reddened litmus, and the ammonia may be obtained by heating the "liquor ammoniæ" of the shops.

This "liquor ammoniæ" is simply the solution of the ammonia-gas in water, and it parts with the gas at a gentle heat.

121. Chemical Properties. — Ammonia combines with the elements of water, and forms *ammonium hydrate.* Thus: —

$$\begin{array}{ccccc} NH_3 & + & H_2O & = & NH_4OH, \\ \text{Ammonia} & + & \text{Water} & = & \text{Ammonium hydrate.} \end{array}$$

The ammonium hydrate of the laboratory is made by passing ammonia-gas through water. A few drops of this solution turns reddened litmus blue, which shows it to be a hydrate. Indeed, it is one of the strongest. But on exposure to air the hydrate is decomposed, and the ammonia-gas escapes.

Ammonia has a strong attraction for hydrochloric acid. Let two little cups stand side by side: into one pour "liquor ammoniæ"; into the other, hydrochloric acid. The two gases arise, mix, combine, and a cloud of white fumes is produced. The new substance is ammonium chloride, or, as it was formerly called, sal-ammoniac. The presence of NH_3 may be tested in this way.

III. — NITROGEN AND OXYGEN.

122. Nitrogen Oxides and Acids. — No less than five compounds of nitrogen and oxygen are well known. Three of these, by uniting with the elements of water, yield acids. The most important of all these eight compounds is nitric acid, HNO_3.

123. Nitric Acid. — This acid is one of the most important in chemistry. It is largely used in the laboratory, and in great quantities, also, for many purposes in the arts. The nitrates are a numerous class of salts, many of which occur in large quantities in the earth. From these the acid is obtained.

124. Preparation. — Let potassium nitrate, often called saltpeter and sometimes niter, and sulphuric acid be heated together: the nitrate will be changed to a sulphate, and the nitric acid will be set free. Thus: —

$$\begin{array}{ccccccc} KNO_3 & + & H_2SO_4 & = & KHSO_4 & + & HNO_3, \\ \text{Potassium} & & \text{Sulphuric} & & \text{Potassium acid} & & \\ \text{nitrate} & + & \text{acid} & = & \text{sulphate} & + & \text{Nitric acid.} \end{array}$$

The nitrate is put into the retort (Fig. 50), whose neck is thrust into a receiver upon which a stream of cold water runs. The nitric acid goes over in vapor into the receiver, and is there condensed into liquid form. This liquid has a yellowish color due to the presence of nitric peroxide.

Sodium nitrate is much cheaper than niter, and is used instead for the manufacture of the acid for commerce.

125. Properties. — Nitric acid is a colorless liquid when pure, intensely acid, corrosive, and poisonous. Let it touch the skin, and it gives a yellow stain which time alone can remove; and, if long in contact, it occasions painful corrosion. Silk, wool, wood, and most organic bodies receive a bright yellow tint by its action.

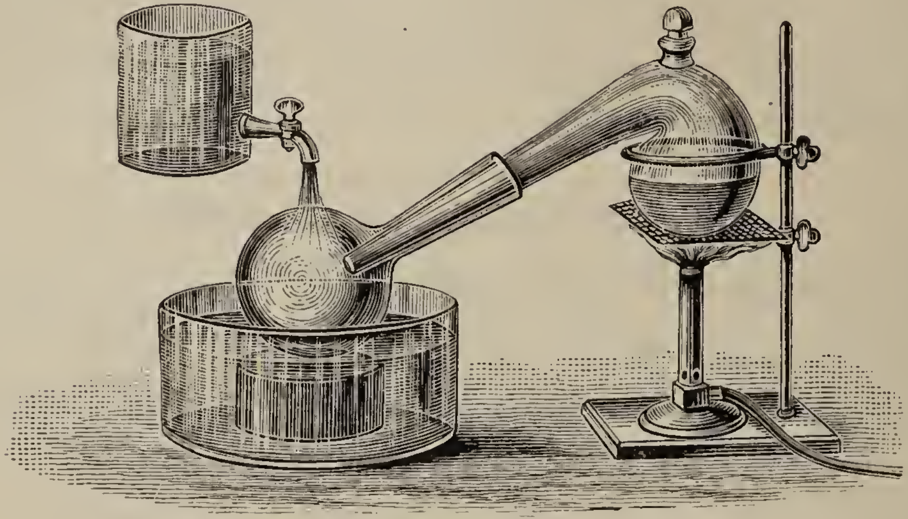


Fig. 50.

It is a powerful oxidizing agent, as a single experiment may illustrate. Place a little tin-foil in a cup, and pour dilute nitric acid upon it. Quickly volumes of red fumes arise, and the tin rapidly changes into a white compound of tin and oxygen. The tin is *oxidized* by the acid.

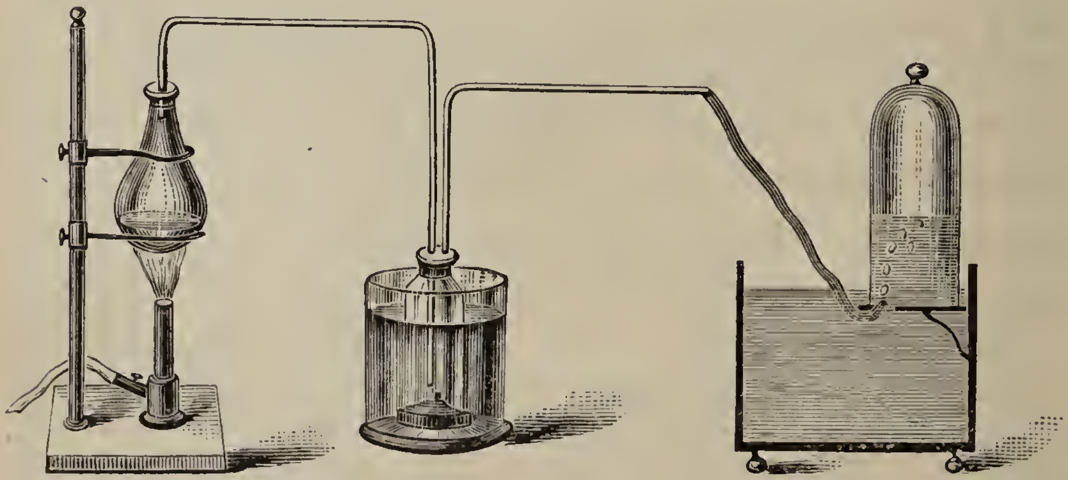


Fig. 51.

126. Nitrous Oxide. — First in importance among the nitrogen oxides is the lowest in the series, the nitrous oxide. It may be obtained by heating ammonium nitrate. The nitrate is put into a flask (Fig. 51), from which a bent tube reaches over into a small bottle standing in a vessel of cold water. Another tube passes from this bottle over to a jar on the shelf of the cistern. By heat the nitrate is melted, and afterward decomposed. Water and nitrous oxide are formed. The water is condensed in the cold bottle, while the oxide is collected in the jar.

$$\begin{array}{ccccc} NH_4NO_3 & = & 2H_2O & + & N_2O, \\ \text{Ammonium nitrate} & = & \text{Water} & + & \text{Nitrous oxide.} \end{array}$$

Properties. — Nitrous oxide is a colorless gas, a little heavier than air. The chemical force between its constituents is weak; a lighted taper decomposes it, and, taking its oxygen, burns with almost as great brilliancy as in oxygen itself. When breathed, its effects upon the system are peculiar. It often causes a lively intoxication, with a disposition to laughter: for this reason it has been called laughing-gas. It often produces entire insensibility, and is administered for this purpose, by surgeons, to patients upon whom they are to operate. If impure, or carelessly given, it may produce serious results.

Composition. — Nitrous oxide is composed of nitrogen and oxygen in the proportions by volume of two of nitrogen to one of oxygen. This composition is represented by the formula N_2O.


Fig. 52.

The Experiment. — The apparatus consists of a eudiometer, shown in Fig. 52. Four equal divisions are marked off from the closed end of the tube. Two of these divisions are filled with nitrous oxide; the remaining two are afterward filled with pure hydrogen. By an electric spark a violent explosion is made; steam is condensed on the side of the tube, and water from the cistern will rise, leaving the two upper divisions only filled with gas. This gas, when tested, is found to be nitrogen.

The Reasoning. — It is clear that the two volumes of hydrogen have taken oxygen enough — one volume — from the oxide, to form the water that was condensed on the tube, and have left two volumes of nitrogen. The nitrous oxide, then, was composed of two volumes of nitrogen and one of oxygen.

If we represent equal volumes of the two gases by equal squares, and their names by their symbols, the composition of the compound may be shown to the eye by the following diagram: —

$$\boxed{\mathrm{N} \;\big|\; \mathrm{N}} + \boxed{\mathrm{O}}$$

and represented also by the formula N_2O.

127. Nitric Oxide. — Nitric oxide may be obtained by the action of copper upon dilute nitric acid, in an apparatus similar to that used in the preparation of hydrogen. (Fig. 53.)

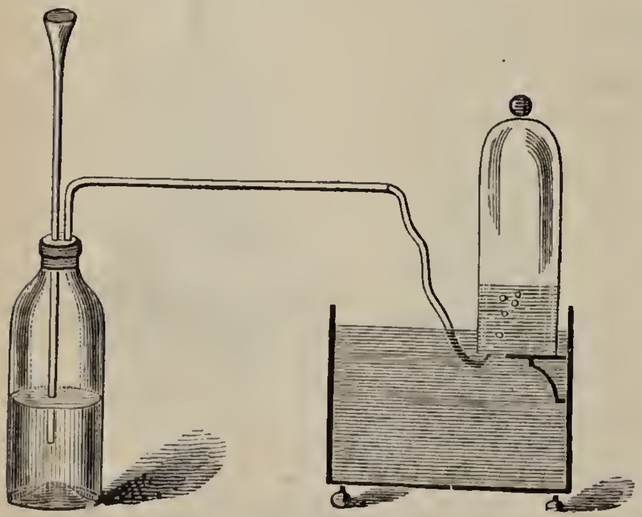


Fig. 53.

The copper decomposes the nitric acid; red fumes fill the bottle; but when the nitric oxide bubbles through the water into the jar it is seen to be colorless and transparent. A lighted taper will be instantly extinguished by this gas; but burning phosphorus will decompose it, take the oxygen from it, and burn with exceeding brilliancy.

Composition. — This gas is decomposed also by heated potassium, and when a measured quantity is used the composition of the gas may be found. It is composed of one volume of nitrogen and one volume of oxygen. This composition is represented to the eye by the following diagram : —

$$\boxed{\mathrm{N}} + \boxed{\mathrm{O}}$$

and its formula must therefore be written NO.

128. Nitrous Anhydride. — Nitrous anhydride is a third compound of nitrogen and oxygen, obtained with difficulty and imperfectly known. It has been found to consist of two volumes of nitrogen and three of oxygen. Thus : —

$$\boxed{\mathrm{N} \mid \mathrm{N}} + \boxed{\mathrm{O} \mid \mathrm{O} \mid \mathrm{O}}$$

The formula is therefore written N_2O_3.

129. Nitric Peroxide. — If nitric oxide is allowed to escape into the air, the dark cherry-red vapors which appear announce its combination with oxygen. This red substance is nitric peroxide. By measuring the volumes of nitric oxide and pure oxygen needed to produce this compound, its composition has been found to be, one volume of nitrogen to two volumes of oxygen. Hence the diagram : —

$$\boxed{\mathrm{N}} + \boxed{\mathrm{O} \mid \mathrm{O}}$$

and the formula which represents it is NO_2.

The production of this cherry-red gas is a useful test for the presence of free oxygen. Whenever the presence of oxygen is suspected, we may decide by passing into it a little nitric oxide, which instantly unites with oxygen to yield the red fumes of nitric peroxide.

130. Nitric Anhydride. — Finally we have a compound which, analyzed, is found to contain five volumes of oxygen and two of nitrogen. Thus: —

N	N	+	O	O	O	O	O

Its formula is N_2O_5. It is called nitric anhydride because it unites with water to form nitric acid. Thus: —

$$N_2O_5 + H_2O = 2HNO_3.$$

IV. — The Atmosphere.

131. Nitrogen is a Constituent of Air. — Let a jet of hydrogen be burned in a bottle of air. For this purpose an india-rubber tube from the bottle B (Fig. 54), ends in a jet-pipe, which passes tightly through one of two holes in a cork which fits the air-bottle A. Set fire to the jet of hydrogen, and insert it in the air-bottle, whose neck may then be quickly lowered into water. The hydrogen burns for a while, and then ceases, the water in the mean time rising a little distance into the bottle.

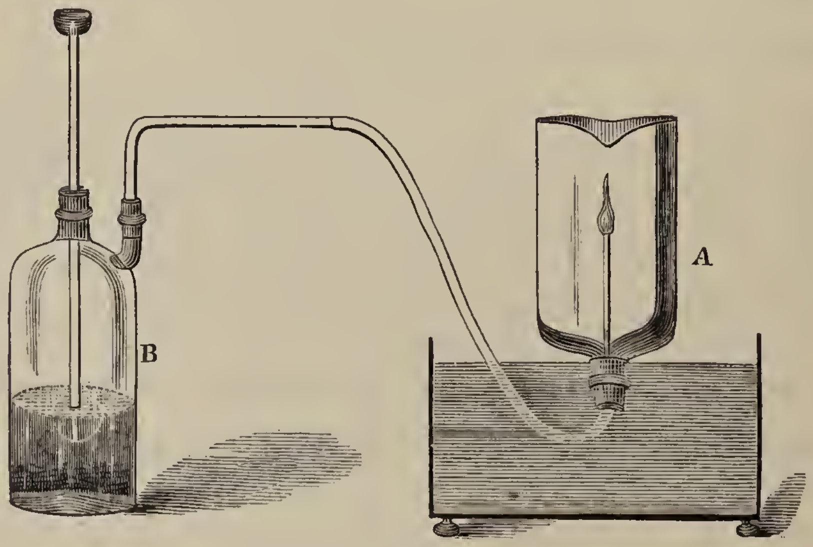


Fig. 54.

If the jet-pipe be now removed, and the remaining gas tested, it will be found to be nitrogen. Hence nitrogen is a constituent of the air.

132. Oxygen is a Constituent of Air. — We have seen that mercury, when heated for a long time in air, is changed to a red oxide of mercury, and that if this oxide is heated to a higher temperature the metallic mercury will be restored while oxygen will be given off. Now observe: the oxygen set free from the oxide in the last heating must have been taken from the air in the first. The experiment teaches that oxygen is a constituent of air.

Carbon Dioxide is a Constituent of Air. — Let a goblet of lime-water stand for a few hours exposed to the air: a crust will be formed on its surface. This crust consists of the well-known calcium carbonate. This substance is always formed when lime-water comes in contact with carbon dioxide, and its presence in this experiment shows that this contact has occurred, and hence that air contains carbon dioxide.

133. Water is a Constituent of Air. — That a vessel of ice-water on a warm day will in a little time be covered with drops of dew, is a fact sufficiently familiar. Now, the water, collected on the vessel, can have come only from the air. It must have been in the air in the form of invisible vapor, which, cooled by the cold sides of the vessel, is condensed. Hence water is a constituent of the air.

134. The Proportions of Nitrogen and Oxygen. — By far the larger part of the atmosphere consists of nitrogen and oxygen: their proportions may be found by experiment. A vessel, V (Fig. 55), has two openings closed with stop-cocks, *c* and *d*. To the upper stop-cock is attached a series of tubes, one, *t*, with a bulb in which is put some copper to be heated by a lamp below; others, *p p*, containing potash; and others still, *a a*, containing sulphuric acid.

The capacity of the vessel is accurately known, and at the beginning of the experiment it is quite full of water. When the stop-cocks are opened, air will flow through the series of tubes, entering at *o*, until the vessel is filled. In passing through *a a*, all its water will be taken out by the acid; in going through *p p*, all its carbon dioxide will be taken by the potash; in passing over the heated copper in the bulb it will give up all its oxygen; and the nitrogen will be left to enter the globe alone.

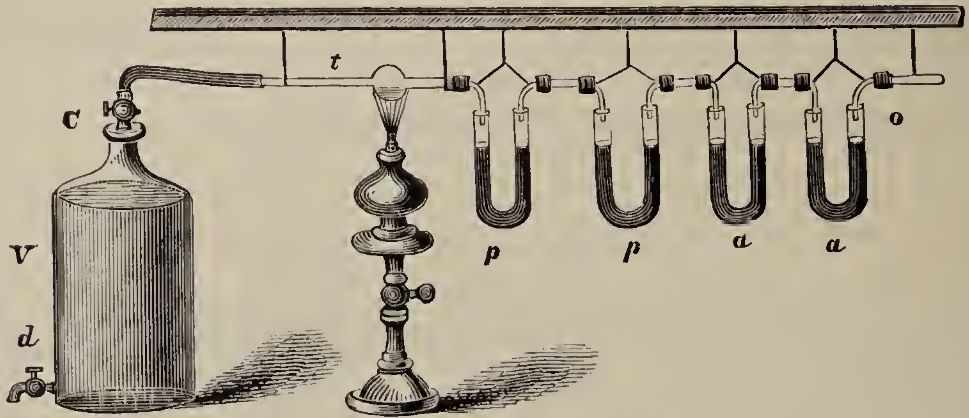


Fig. 55.

Now notice: if the tube *t* be accurately weighed before and after the experiment, what it has gained will be the weight of the oxygen taken from the air that passed through it. The vessel V is filled with the nitrogen from the same air; and, since the capacity of the vessel is known, the weight of the gas may be calculated. By this means it has been found that one hundred parts of air from which water and carbon dioxide have been taken, contain, by weight, 76.9 parts of nitrogen to 23.1 parts of oxygen.

135. The Proportions of Water and Carbon Dioxide. — Now, the quantities of water and carbon dioxide might also be found in this experiment by weighing the tubes in which they are left, were it not that in so small a quantity as the vessel full of air they are too small to be very appreciable. If, however, larger quantities of air are passed, the increase in weight will be quite enough. It has been found that the quantities of these substances vary at different times and places, being always very small.

The quantity of water is exceedingly variable. It depends upon locality and temperature. At 15° C. the largest quantity that air can hold is $\frac{1}{80}$ of its own weight. Commonly the amount is far less than this.

The proportion of carbon dioxide is usually given by volume. It varies from $\frac{1}{5000}$ to more than $\frac{1}{1000}$, averaging about $\frac{1}{2500}$.

136. The Air is a Mixture. — In the properties of the atmosphere we discover none that do not belong to one or another of its constituents. The oxygen causes bodies to burn, not so freely, of course, as if it were pure. The nitrogen, if pure, would quench all flames; in the air it hinders the burning which it can not quench. So, too, the water-vapor, by being cooled, is condensed into clouds and rain-water, just as this substance when pure first becomes visible as white vapor and then changes to water. And finally, the carbon dioxide of the air turns lime-water white just as the pure gas would do, only in a less degree. Hence the air is a mixture of its constituents, not a compound.

137. Diffusion. — The gases of the atmosphere have very different densities. The oxygen is heavier than the nitrogen, and the carbon dioxide is still heavier than oxygen. Yet the atmosphere is a *uniform mixture* of these gases; the heavy and the light remain mixed at all heights above the earth. Left to the force of gravity they would separate into layers, the heaviest at the bottom. Evidently they are under some other influence which overcomes this tendency; and this mixing of gases of different densities, when brought in contact, is called Diffusion. The same action is also observed in liquids.

138. The Diffusion of Liquids. — If in a tall jar some colored water is placed, alcohol may, with care, be poured upon it without mixing the fluids. Being lighter than water, the alcohol floats; and, being colorless, the line of division may be distinctly seen. After a few hours the liquids in the jar will be found colored uniformly throughout. The heavy water has risen; the lighter alcohol has sunk, and a uniform mixture is formed. Liquids, which when shaken together, will remain mixed, will diffuse when brought in contact; but those which separate again when standing a while will not diffuse.

Osmose of Liquids. — A thin membrane or porous substance will not prevent liquids from mixing. To illustrate: let a bladder be firmly tied over the end of a lamp-chimney, or other glass tube; and, being filled with alcohol, let it be pressed a little way into a vessel of water (Fig. 56). After a time the fluid will be seen gradually rising in the chimney. The water flows in to mix with the alcohol, and a smaller quantity of alcohol flows out at the same time. This mixing of liquids through porous solids is called OSMOSE.

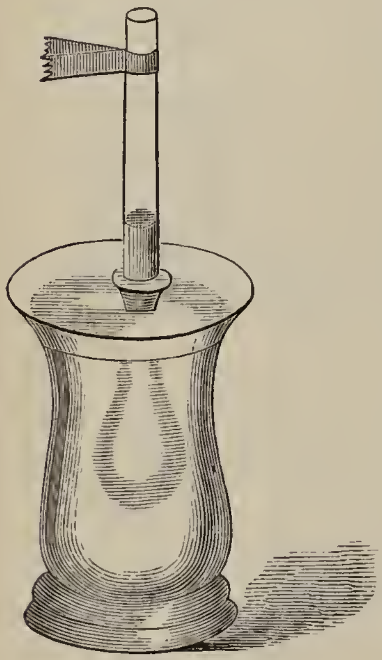


Fig. 56.

139. The Diffusion of Gases. — Diffusion among gases is more rapid than among liquids. The following experiment will clearly illustrate this important property. Two strong bottles with narrow necks, one filled with oxygen, the other with hydrogen, are placed with their open necks together, the oxygen being at the bottom. After considerable time the gases in both bottles will explode at the touch of a lighted taper. Neither oxygen nor hydrogen is explosive, but their mixture is. The oxygen must have risen into the upper jar, and the hydrogen must have fallen into the lower one, until a mixture was formed in both. Gases diffuse in spite of gravitation; nor will other mechanical forces prevent them: a gas will spread in direct opposition to a current of air.

On examination it is found that hydrogen will diffuse four times as fast as oxygen. Now, the densities of these gases are as 1 : 16 ; but notice, their diffusive rates are as 4 : 1. In other words, *their diffusive rates are inversely as the square roots of their densities.* This law applies to the diffusion of all gases.

The Osmose of Gases. — Porous substances can not stop the mixing of gases. Hydrogen can not for any length of time be kept pure in india-rubber bags: it will pass through the pores to mix with the air outside. The osmose of hydrogen and air may be shown very beautifully with the apparatus seen in Fig. 57. A porous cup (of Grove's battery) is fitted air-tight to a tube which reaches into colored liquid, contained in a bottle. If a jar of hydrogen is held over the cup, bubbles instantly rush out of the tube, showing that hydrogen has entered to mix with the air inside. Nor is this all: if the jar is removed, the liquid quickly rises in the tube, showing that the hydrogen leaves the cup, to mix with the air outside, and that it flows out so much faster than air can flow in, that a partial vacuum is formed.

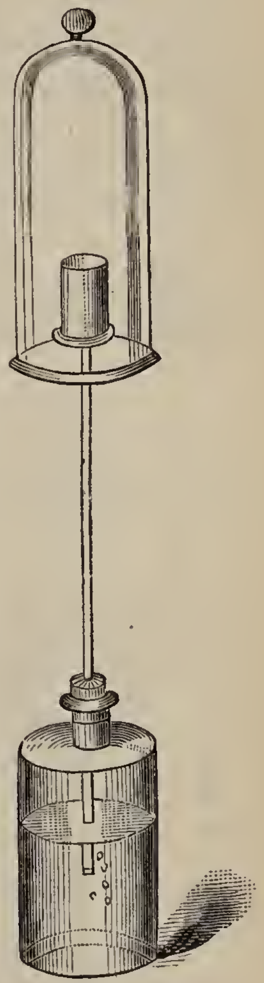


Fig. 57.

Application to the Air. — We can now see how gases of such different weights as those which form the atmosphere are kept uniformly distributed, instead of forming layers with the heaviest at the bottom. In obedience to this law of diffusion, the heavier gases are compelled to rise, and the lighter ones to fall, until the proportions of them all are the same throughout. Think, moreover, how vast the quantities of unwholesome, and even poisonous gases, which are given off by decaying substances, and from sewers, swamps and marshes. Were gases subject only to the laws of gravitation, the heavy and poisonous carbon dioxide, with miasms and effluvia, would render animal life impossible.

V. — Phosphorus.

140. Phosphorus in Nature. — Compounds containing phosphorus are quite common in soils and rocks. Plants receive them from the soil; while animals, living upon vegetable food, take these compounds from the plant. They are found in wheat and other grains, and in the brain and secretions of animals; so that, while not a very abundant element, phosphorus is widely distributed and very important. Its most common compounds are the phosphates, and of these the most common is the calcium phosphate. This compound of phosphorus occurs in bones, and from these the element is obtained.

141. Its Physical Properties. — Phosphorus is a solid element, having a pale yellow color, rather soft and wax-like at common temperature, but brittle at 0° C. In a dark room its clean surface shines with a feeble, pearl-white light. Its vapor also is beautifully phosphorescent. To show this, put a few small fragments of the solid into a flask of water, and boil it. The mixed vapors of water and phosphorus escape (Fig. 58), and, in a dark room, look like livid flames issuing from the mouth of the flask, while at the same time globules of melted phosphorus, on the surfaces of the water and the glass, appear like little balls of pearl.


Fig. 58.

This element is insoluble in water; but it dissolves freely in ether, and still more freely in carbon disulphide.

142. Its Chemical Properties. — Phosphorus combines most readily with oxygen. From a piece of the solid exposed to air, white fumes are continually falling, which consist of a compound of phosphorus and oxygen. A gentle heat — that of the fingers handling it is often enough — causes it to burst into violent combustion, forming the fumes in great abundance. On this account the element must be kept under water; and it should be held under water when cut, lest the friction of the knife set it on fire. The pieces to be used should be afterward dried by gentle pressure between layers of blotting-paper.

There is, however, an allotropic form of phosphorus, called red phosphorus, which will not produce the phenomena of combustion under a temperature of 260° C. Experiments with this variety are, of course, far less dangerous than with the common form. Sticks of the solid, exposed to light, gradually change to red phosphorus.

This element is a violent poison: even its vapors, inhaled, will cause wasting disease.

143. Uses. — The most important use of phosphorus is in the manufacture of matches, hundreds of tons being employed annually for this purpose.

Formerly the match was made by first dipping the end of the wood into melted sulphur and afterward into a paste of gum-water, phosphorus, and either niter or potassium chlorate. By friction the phosphorus is set on fire. The heat of the burning phosphorus is intense enough to ignite the sulphur, which in turn inflames the wood. But sulphur is seldom now employed, its place having been taken by paraffine or stearic acid. When potassium chlorate is used, the match burns with an explosive combustion.

The "safety match," now in quite common use, contains no phosphorus. This element is put upon the box, the surface on which the match is rubbed, instead of upon the match itself. The matter on the tip of this match contains potassium chlorate and sulphur or antimony sulphide, and sometimes red lead and potassium bichromate. The matter on the side of the box contains *red phosphorus.* Only by rubbing the match against this surface, can it be easily fired.

VI. — PHOSPHORUS AND HYDROGEN.

144. Hydrogen Phosphide. — Three compounds of phosphorus and hydrogen are known. One, PH_3, is a gas at ordinary temperatures; another, P_2H_4, is a liquid; and a third, P_2H, is a solid.

The gaseous compound may be obtained by the action of phosphorus on a solution of "caustic potash," KHO. The experiment is represented in Fig. 59.

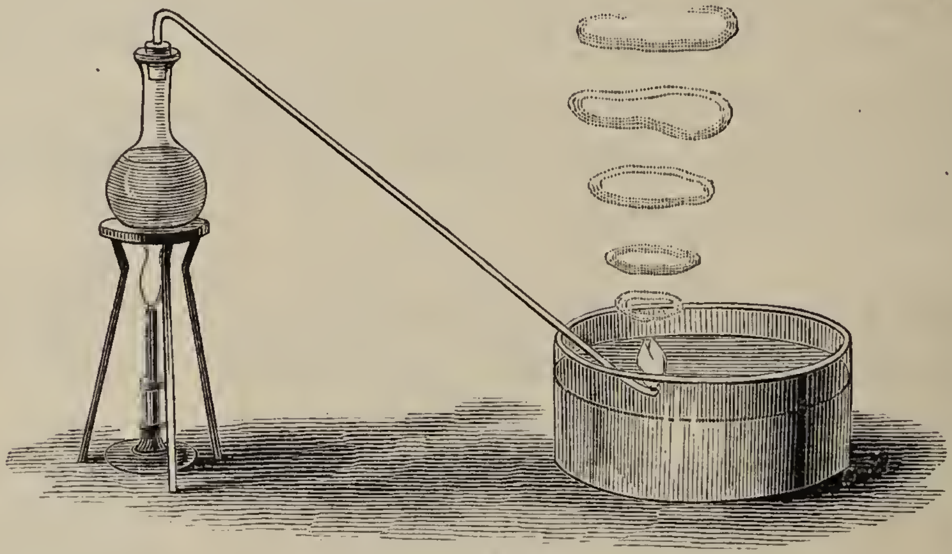


Fig. 59.

The small flask is two-thirds filled with a moderately strong solution of potassium hydrate, and a small piece of phosphorus is added. A few drops of ether are poured in, and the flask is then closed. The delivery tube dips into a cistern of water.

On applying a gentle heat the ether evaporates, and its vapor *expels the air* from the flask and tube. Very soon bubbles of gas arise from the phosphorus; and after a time larger bubbles emerge from the tube in the water, come in contact with the air, and instantly take fire with explosion. The bright white flame is followed by a beautiful vortex-ring of white vapor (P_2O_5) rising and expanding, and finally breaking and vanishing in the air.

And yet the *gaseous* phosphide is not spontaneously inflammable, as it would appear to be: it has been found that, made in this way, it contains some vapor of the *liquid* phosphide, and to this the spontaneous combustion is due.

VII. — PHOSPHORUS AND OXYGEN.

145. Phosphoric Oxide and Acid. — When phosphorus burns in *dry* oxygen, or in a full supply of *dry* air, a snow-white solid is produced: it is the phosphorus pentoxide, P_2O_5, also called phosphoric anhydride. This snow-like solid has so strong an attraction for water, that exposed to the air it takes moisture and melts away, or if brought in contact with water itself the two unite with such vigor as to produce a hissing sound, and much heat. The powder can be preserved only in hermetically sealed flasks.

If now some drops of the solution of this oxide are added to blue litmus, the color instantly changes to red, showing the liquid to be an acid. The P_2O_5 has combined with $3\ H_2O$, yielding $2\ H_3PO_4$, or two molecules of *phosphoric acid.*

146. Phosphorous Oxide and Acid. — If, instead of burning the phosphorus rapidly, it be made to burn slowly in a small supply of dry air, it produces a different compound. The phosphorous trioxide, P_2O_3, is formed.

This phosphorous oxide is also a white solid, and when brought into contact with water instantly unites with it. The P_2O_3 takes $3\ H_2O$, and yields $2\ H_3PO_3$, or two molecules of *phosphorous acid.*

147. Other Acids. — Besides the two acids of phosphorus just mentioned, there are three others, whose names and composition are here given for reference: —

Hypophosphorous acid	H_3PO_2,
Metaphosphoric acid	HPO_3,
Pyrophosphoric acid	$H_4P_2O_7$.

VIII. — ARSENIC.

148. Arsenic in Nature. — Arsenic sometimes occurs in the earth uncombined with other substances; but generally it is found as an oxide or a sulphide, mixed with similar compounds of the metals. "It was at one time supposed that arsenic entered into the composition of the flesh and bones of animals as a normal constituent; but it has been clearly proved that it is never found in the tissues, either of animals or vegetables, except when it has been introduced into them by accident or design." (BRANDE & TAYLOR.)

Its Physical Properties. — Arsenic is a very brittle solid, in appearance much like metals. It may be sublimed by heat; that is to say, it may be changed from the solid to the vapor state, directly, without melting.

Its Chemical Properties. — Heated in the open air, arsenic rapidly combines with oxygen, and even on simple exposure to air it forms an oxide. With some of the metals it forms arsenides. Owing to the criminal use of its poisonous compounds, this element has been most thoroughly and successfully studied.

IX. — ARSENIC AND HYDROGEN.

149. Hydrogen Arsenide. — Let fragments of zinc be put into a bottle (Fig. 60) with water enough to cover them. Let sulphuric acid be poured through the funnel-tube until a brisk evolution of hydrogen begins. Patiently wait until all air has been driven out of the bottle, and then touch a match-flame to the tip of the taller tube from which the gas is escaping. The almost colorless flame of hydrogen is thus obtained.

Next pour in through the funnel a few cubic centimeters of a solution of arsenious oxide As_2O_3: a moment afterward the flame will become enlarged, and shine with a livid whiteness.

The hydrogen has decomposed the arsenious oxide, and taken its arsenic to form hydric arsenide, AsH_3, commonly called arseniuretted hydrogen.

Properties. — This substance is a colorless gas of unpleasant odor, and very poisonous. It is very combustible, and the products of combustion are water and arsenious oxide. If the flame be cooled by holding the cold surface of a piece of porcelain across it, then only the hydrogen will burn away; the *arsenic will be deposited on the porcelain* in the form of a brown and lustrous mirror.

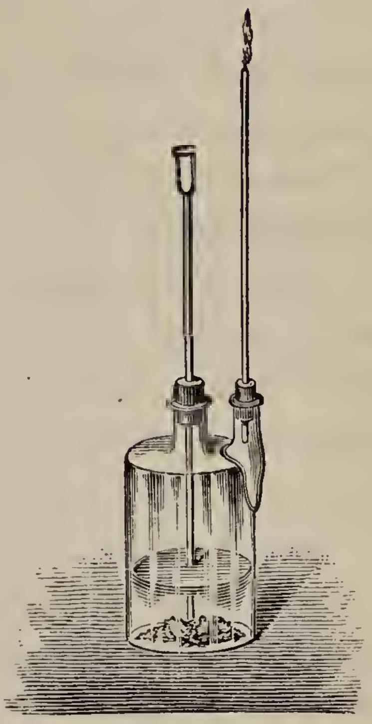


Fig. 60.

150. Marsh's Test. — This reproduction of arsenic as a mirror on porcelain is a most delicate and important test for the presence of the compounds of this poisonous substance. It is known as Marsh's test. It is applied by means of the apparatus shown in Fig. 61.

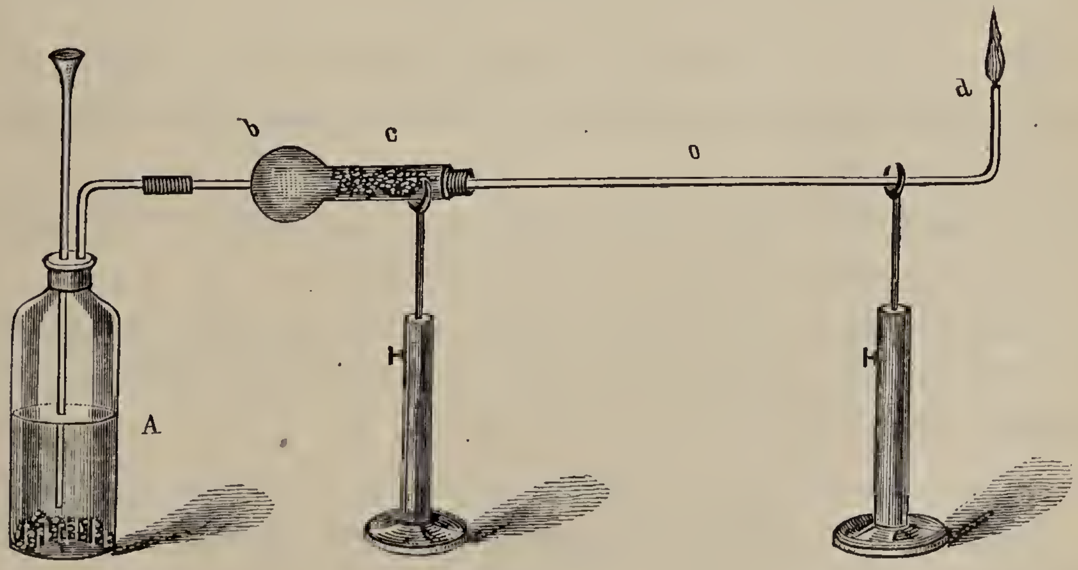


Fig. 61.

Into the bottle A are put *pure* zinc, water, and sulphuric acid for the evolution of hydrogen gas. This gas is allowed to flow until all air is expelled from the apparatus. The liquid supposed to contain the poison may be afterward poured through the funnel-tube. The gas formed in the bottle, passing through the bulb, *b*, loses a part of the water carried over with it, and going over calcic chloride in *c*, it is thoroughly dried. It finally escapes at the pointed end of the tube, *d*. After the air of the apparatus has been all driven out by the stream of gas, the hydrogen may be burned as it issues.

Now, if the liquid contain arsenic, there is at once formed arseniuretted hydrogen (H_3As). In presence of much arsenic the color of the flame turns to a livid hue, and sometimes gives off white fumes of arsenious oxide. But now comes the decisive test. A cold, clean, white porcelain surface is held in the small flame for a moment; *metallic arsenic* condenses on its surface, if the poison is present, and, when cold, appears a blackish-brown stain, with a bright *metallic luster*. The substance of this stain may be still further tested, until the last doubt of its character is removed.

The experiment is varied by applying heat to the middle part, *o*, of the small tube. The arseniuretted hydrogen, if present, will be decomposed, and the metallic arsenic will lodge upon the inside surface of the tube, forming a brilliant, mirror-like ring.

By Marsh's test, even so little as 0.01 of a milligram ($\frac{1}{7000}$ gr.) can be detected with certainty.

This is only one of many tests by which, taken together. the chemist can pronounce upon the presence of arsenic with absolute certainty.

151. Some other Compounds of Arsenic. — The arsenious oxide (As_2O_3) is the common "arsenious acid" or "white arsenic," — the well-known poison. It is a white solid, slightly soluble in water, with which it forms an acid, and this acid combines with metals to form arsenites.

These, also, are violent poisons; and yet some of them are used quite largely in the manufacture of aniline dyes, in calico-printing, and as pigments. Green wall-papers, many of them, have been colored with copper arsenite, well known as Scheele's green. Nor is green the only color of wall-papers containing arsenic: it has been found in yellow, pink, blue, and drab. These arsenical papers are guilty of poisonous action upon those who inhabit the rooms.

Another compound of arsenic and oxygen has the formula, As_2O_5: it is the arsenic oxide. This oxide is also an anhydride, since it unites with water to form arsenic acid. This acid is represented by the formula, H_3AsO_4.

$$As_2O_5 + 3\,H_2O = 2\,H_3AsO_4.$$

By passing sulphuretted hydrogen through a solution of arsenious oxide, made acid by a few drops of hydrochloric acid, a fine yellow precipitate is formed at once. This yellow solid is arsenious sulphide, As_2S_3. It has long been known under the name of orpiment, and was formerly much used as a pigment called king's yellow.

X. — BORON.

152. Boron in Nature. — The element boron is not found free in nature, but several of its compounds exist in considerable quantities. The most important of these are boric or boracic acid, and sodium biborate, commonly called borax.

The Element. — Boron is sometimes a dark-brown powder, and sometimes it occurs as lustrous crystals almost as hard as diamond. In neither of these allotropic forms does boron show much chemical activity. Boron is the only non-metal which forms no compound with hydrogen.

153. Boracic Acid. — The waters of the lagoons of Tuscany hold this acid in solution. These waters, evaporated by the heat of volcanic jets of steam issuing in the neighborhood, yield the acid in solid form. This is the source of it for the European market.

In the United States it is manufactured from the native borax found in sufficient quantity in the water of the borax lake in California.

Let a little borax be dissolved in two and a half times its weight of boiling water, and add a little more than half its weight of strong hydrochloric acid. On cooling, this solution will deposit glistening plate-like crystals. These crystals are boracic acid.

Place a little of this crystallized acid in a capsule, and add a spoonful of alcohol. After stirring it well, to dissolve the acid, set fire to the solution, stirring it the while, and the flame will exhibit a fine green color. This is a characteristic test for this acid.

154. The Borates. — With sodium this acid forms sodium biborate, or borax, $Na_2B_4O_7$. With other metals other borates are produced, the class being quite numerous.

Many of these are crystalline, and in the form of crystals almost always contain water. Borax, for example, in crystallizing, takes up ten parts of water, so that its formula becomes $Na_2B_4O_7 + 10\,H_2O$. This water contained in crystals is called WATER OF CRYSTALLIZATION. Oftentimes, when gently heated, the crystals will dissolve in this water which they contain. Heat crystals of borax in an iron spoon, and they "melt in their water of crystallization."

XI. — THE TRIVALENT GROUP.

155. The Group. — Nitrogen, phosphorus, arsenic and boron are trivalent elements. Many other chemical relations also mark them, with the exception of boron, as members of a natural group.

They are Trivalent. — The compounds of these elements with hydrogen have the following symbols: —

H_3N, H_3P, H_3As,

from which we learn that an atom of each can hold three atoms of hydrogen in combination. These three compounds are gases.

Other Chemical Relations. — The complete analogy in the composition of the compounds of these elements may be seen by a glance at the following formulas: —

N_2O_3,	N_2O_4,	N_2O_5,	NCl_3(?),		
P_2O_3,	. . .	P_2O_5,	PCl_3,	P_2S_3,	P_2S_5,
As_2O_3,	. . .	As_2O_5,	$AsCl_3$,	As_2S_3,	As_2S_5.

REVIEW.

I. — SUMMARY OF PRINCIPLES.

156. Nitrogen, phosphorus, arsenic, and boron are the members of the trivalent group of non-metals, one atom of each being equivalent to three atoms of hydrogen.

Nitrogen does not directly combine with oxygen. Phosphorus and arsenic do.

Nitrogen is a colorless gas; phosphorus, a translucent solid; arsenic also, a solid, steel-gray and metallic.

Nitrogen and hydrogen form a single compound, ammonia, a gas, very soluble in water with which it combines to form liquor ammoniæ, a strong base.

Ammonia combines directly with hydrochloric acid to form ammonium chloride.

Nitrogen unites with oxygen to form five compounds, illustrating well the "law of multiple proportions."

Two of these oxides are anhydrides; with water one forms nitrous acid, and the other nitric acid.

Nitric acid is a powerful oxidizing agent.

Nitrous oxide is an anæsthetic much used by dentists.

Nitric oxide is a test for the presence of oxygen, with which it instantly forms red vapors of nitric peroxide.

The atmosphere is a mixture of nitrogen, oxygen, carbon dioxide, and water-vapor, together with small and variable quantities of ammonia, and several other gases.

Compared with the nitrogen and the oxygen, the quantity of all other constituents is small.

By volume, the air consists of:—

Oxygen	20.77
Nitrogen	79.23
	100.00

By weight, the air consists of:—

Oxygen	22.92
Nitrogen	77.08
	100.00

For most purposes it is sufficient to say that four-fifths of the air is nitrogen, and one-fifth is oxygen.

Diffusion is the tendency of fluids or of gases to mix when brought in contact. The principle of diffusion is well illustrated by the uniform mixture of gases throughout the atmosphere.

Phosphorus burned in a full supply of air or oxygen yields phosphoric pentoxide, P_2O_5; but burned in a limited supply it yields phosphorous trioxide, P_2O_3.

These oxides are anhydrides: one of them with water forms phosphoric acid, the other phosphorous acid.

These acids form large classes of salts, one the phosphites, the other the phosphates.

The compounds of arsenic are remarkable for their poisonous character. Nevertheless some of them are used in many industrial arts.

There are two oxides of arsenic; viz., the arsenious oxide, As_2O_3, and the arsenic oxide, As_2O_5.

These oxides are anhydrides: with water, one yields arsenious acid, the other arsenic acid.

With metals these acids form salts: the first yields arsenites, and the second arsenates. "Scheele's green" is a copper arsenite, and Schweinfurt's green consists of copper arsenite and acetate. Both are used as pigments in coloring wall-paper.

The arsenic in paper can be easily detected. Moisten the paper with hydrochloric acid, and then add water. Into this solution thrust a bright copper wire. Arsenic, if present, will be deposited on the wire, giving it a gray coating.

The most sensitive and reliable test for arsenical poison is Marsh's test.

The antidote to arsenical poison is moist ferric hydrate, which can be made by adding ammonia-water to a solution of iron sulphate (copperas), after having first heated the solution with a little nitric acid. In absence of this antidote milk, or white of eggs, may be used. Or a powerful emetic may be administered.

Boron has two allotropic forms, one amorphous, the other crystalline.

It is the only non-metal which has no known compound with hydrogen.

Its most important compounds are boric acid and borax.

II. — EXERCISES.

Give the preparation of nitrogen. What are its physical properties? What is its chemical character?

What are its atomic weight and density?

How does it occur in nature?

What compound does nitrogen form with hydrogen? What is the source of the ammonia in the air? Define nascent state. What influence has the nascent state on combination?

How is ammonia usually obtained? Give the reaction. Consult the table of elements, and convert this reaction into a numerical equation. What does this numerical equation show? How much ammonium chloride is required to yield one hundred grams of ammonia? To yield one hundred liters of ammonia?

What are the physical properties of ammonia? What is ammonium hydrate? How is this hydrate made in the laboratory?

How many oxides of nitrogen are known? How many oxygen acids?

How is nitric acid prepared? Give the reaction.

What are the properties of nitric acid? Why is it called an oxidizing agent?

Describe the preparation of nitrous oxide. What are its properties? For what is it useful? How is its composition represented? Describe the experiment to determine its composition. Give the reasoning.

How may nitric oxide be obtained? What is the composition of nitric oxide? What formula represents its molecule? How does it serve to test the presence of oxygen?

Give the formula for nitrous anhydride. The formula for nitric peroxide. The formula for nitric anhydride.

What acid comes from nitric anhydride? Show by an equation the composition of the acid derived from nitrous anhydride.

Name the two most abundant constituents of the atmosphere. Name other constituents. Give the composition of the air by volume. By weight. How are these proportions determined?

What are the proportions of water-vapor and carbon dioxide? Why is the air called a mixture?

Is this mixture uniform at all heights? On what principle are the gases kept from settling into layers? Define diffusion. State the law of diffusion of gases.

How does phosphorus occur in nature? What are its physical properties? With what element has oxygen the strongest inclination to combine? Describe the allotropic form of phosphorus. Describe the composition of the match. What is the safety-match?

How does arsenic occur in nature? What are its physical properties? Its chemical properties?

How may arseniuretted hydrogen be prepared? What is its formula? Describe its combustion.

What is Marsh's test? Describe the apparatus employed. How is the experiment conducted?

What is the so-called "arsenious acid"? What is its true chemical name?

What is the most remarkable general character of the arsenic compounds?

To what uses are many arsenical compounds applied? How can the arsenic in wall-paper be detected?

Name the members of the trivalent group of non-metals. Show that they are trivalent by means of the formulas for their hydrogen compounds. In what else are they analogous?

SECTION VI.

THE QUADRIVALENT NON-METALS.

I. — SILICON.

157. The Element. — Silicon is a solid substance, sometimes obtained as a brown powder, sometimes as crystals with an iron-gray color and metallic luster. It is very hard, very insoluble, and very difficult to melt; but when heated in air it burns because of its attraction for oxygen.

158. Its Compounds. — Silicon is never found free in nature, but its compounds are very abundant. With oxygen it forms a silicon oxide, $Si\,O_2$, more commonly called silica.

This compound is one of the most abundant substances in the earth. The beautiful opal and amethyst and some other gems are almost pure silica: so is the more common rock crystal or quartz, while the sand of the sea-shore and every variety of sandstone rock are the impure forms of the same substance.

Silica is also an important constituent in organic bodies. It gives strength to the stalks of grains and grasses, while it constitutes the skeleton of whole tribes of some of the lower orders of animals.

Silica is used largely in the manufacture of glass.

Silicic acid, H_2SiO_3, is a limpid liquid of little known importance; but its salts, the silicates, occur in large quantities and are widely distributed in the earth.

159. Glass. — Glass is made from silica and the bases, potash, soda, lime, and some others, according to the variety required. These materials, being intensely heated, melt into a transparent pasty mass, portions of which may be taken from the furnace, and blown or molded into the different forms in which glass articles are made. These articles, being afterward annealed, are ready for the market. We may study the manufacture of glass a little more in detail.

The Materials. — In all true glasses silica is one constituent. This substance exists in almost pure form in flint, agate, and quartz; while all varieties of sand consist of the same in various degrees of purity. Flint was formerly used in the manufacture of glass: hence the name, *flint glass*, which one variety still retains. Sand is now the more general source of silica, great care being used to select a pure material.

Potash, soda, and lime are the most important bases used with the silica; but, besides these, lead oxide and oxides of tin and manganese are often employed.

The Varieties of Glass. — There are several varieties of glass, chief among which are four, viz.: green bottle-glass, Bohemian glass, window-glass, and flint-glass.

Green bottle-glass is made of cheaper and coarser material than any other variety. Its bases are more numerous: they are chiefly soda and lime; but the oxides of iron and manganese are among them, and to the first of these the glass owes its familiar color.

Bohemian glass is made from purer material than bottle-glass, and care is taken to have it free from color. Its bases are potash, lime, and manganese. Its lightness, the absence of color, and its power to stand high heat and sudden changes of temperature, make it very valuable for chemical purposes.

Window-glass consists chiefly of silica, soda, and lime.

Flint-glass consists chiefly of silica, potassa, and lead oxide (Pb_2O_3). This very beautiful variety of glass, sometimes called crystal glass, is chiefly made into such articles of domestic and ornamental use as tumblers, decanters, wine-glasses, and vases. It is very transparent, and refracts light powerfully: on this account it is valuable for lenses, and other optical instruments.

The Melting. — The raw material is heated in pots made of the purest and most infusible clay, and set in a conical furnace. When the heat is sufficient, the silica and the bases form silicates. In the case of window-glass, for example, the silica, soda, and lime form SODIUM and CALCIUM SILICATES. The compound of these silicates is a transparent half-fluid or pasty mass of *glass*, ready to be wrought into any desired form.

The Blowing. — By means of an iron tube four or five feet long, the workman takes out of the furnace a portion of the pasty and adhesive glass. He gives it regular shape by rolling it upon a smooth hard surface, and makes it hollow by blowing air through the tube. By keeping the tube in constant rotary motion while he blows, the bulb is enlarged into a globe; or if, at the same time, he swings his tube, pendulum like, a pear-shaped flask is made.

Some idea of the interesting process of making window-glass may be gained by studying the following cuts found in Muspratt's Applied Chemistry. In Fig. 62 some of the first stages of the process are shown. In front of each opening of the furnace is a stage, built over a pit about ten feet deep. Upon these stages the workmen stand. The ball of glass having been taken from the furnace, the workman blows through his pipe, while he at the same time skillfully rotates it in his hand, and swings it backward and forward, sometimes below his feet, sometimes over his head, until he has lengthened and enlarged it into a cylinder with uniform sides as long as the pane of glass is expected to be. The ends of this cylinder are then cut off, and a slit is made along one side. It is afterward placed upon a smooth and even stone plate, and heated in an oven. As the glass softens, the workman, with an iron rod, presses the sides of the cylinder open (see Fig. 63), and finally smooths it out upon the flat surface of the stone. It is then a *pane of glass*, needing only to be carefully cooled.

Fig. 62.

The Annealing. — Annealing is a process of slow cooling. All articles of glass must be annealed. For this purpose they are placed in hot ovens, whose temperature grows gradually less, until, at the end of four or five days, they are quite cold. The process is necessary, because if glass is cooled suddenly it becomes exceedingly brittle, and will sometimes break even without apparent cause; but when slowly cooled it is able to stand much pressure and sudden blows.

"Considered only with reference to its application in the study of natural phenomena, it is impossible to doubt the singular influence glass has exerted on the progress of science. It is chiefly by its aid that astronomy has attained a perfection so wonderful; by it, also, naturalists have been enabled to study under the microscope a host of phenomena which before escaped notice. But perhaps of still greater importance is the use made of it by chemists in their experiments. It requires no profound chemical knowledge to recognize the fact that to glass is chiefly owing the present advanced state of the sciences, so fruitful in marvelous application."

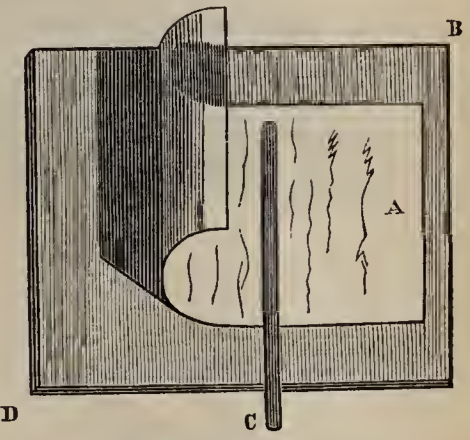


Fig. 63.

II. — Carbon.

160. Charcoal obtained from Wood. — When splinters of wood are heated in a test-tube closed with a cork through which passes a small tube, considerable quantities of vapor and gas escape, and the wood turns black. The black mass that remains when the gas is no longer given off is charcoal, one form of carbon. In the same way, on a large scale, wood is put into retorts, and heated intensely; its liquid and gaseous constituents are driven off, while the solid carbon remains.

Or, again: if we light one end of a splinter of wood, and slowly push the burning part into the mouth of a test-tube, the part in the tube will only glimmer in the small supply of air, or be extinguished altogether. The black residue is carbon.

On a Large Scale. — In a similar way, when charcoal is required in large quantities, piles of wood are burned under a covering of sod and moist earth, where but little air can reach them. Figs. 64 and 65 represent the way in which the method is carried out. In the former we see how the sticks are piled; in the latter we see the pile in its slow combustion. Air can enter only through holes in the earthy covering near the base. Smoldering sometimes for weeks, the gaseous parts of the wood are at last all driven off; while the carbon, keeping the form of the original sticks, with knots and

Fig. 64.

Fig. 65.

annular rings still perfect, all, however, much reduced in size and weight, remains.

Charcoal is without crystalline form, and on this account it is said to be Amorphous.

161. Other Forms of Amorphous Carbon. — Beside charcoal there are several other forms of amorphous carbon, among which we may mention lamp-black, coke, and animal charcoal.

Animal charcoal is carbon obtained from animal substances. *Bone-black*, for example, is animal charcoal obtained by charring bones in iron cylinders.

Coke is the carbon which remains when soft coal is heated to redness in the absence of air. Large quantities of soft, or bituminous coal, are used for the manufacture of illuminating gas, and large quantities of coke are left behind. It is a fuel of considerable value.

Lamp-Black. — Bodies which contain large proportions of carbon are likely to burn with a smoky flame. They are sure to do so if the supply of air in which they burn is limited. The blackness of the smoke is due to minute particles of carbon. Let a metallic plate be held over the flame of an oil-lamp, and we shall find an abundance of this finely

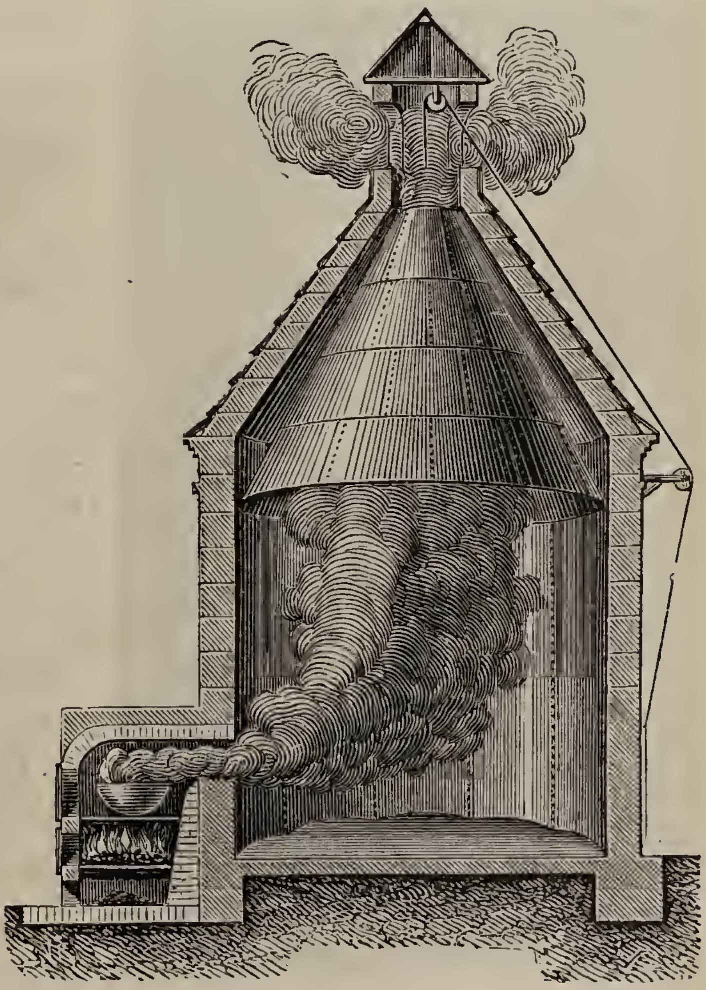


Fig. 66.

divided carbon deposited on it. In this form the carbon is called Lamp-black.

This form of carbon is much used in the arts. On a large scale it is made by burning tar, resin, or petroleum, in such a way that but little air will reach it. The dense black smoke is drawn over into large chambers (Fig. 66), on whose walls the soot, or lamp-black, is deposited.

Lamp-black is used as a black paint, and when mixed with oil it constitutes printer's ink. The finest kind is also used for preparing India ink and in calico-printing.

162. The Diamond. — The diamond is crystallized carbon. A roughly rounded pebble found in the sand and clay of India or Brazil, the diamond would hardly be picked up by one whose eye had not been trained to recognize it; but when polished it is at once the most beautiful and the most indestructible of gems. It can neither be melted nor dissolved: it may, however, be burned when heated intensely in oxygen-gas.

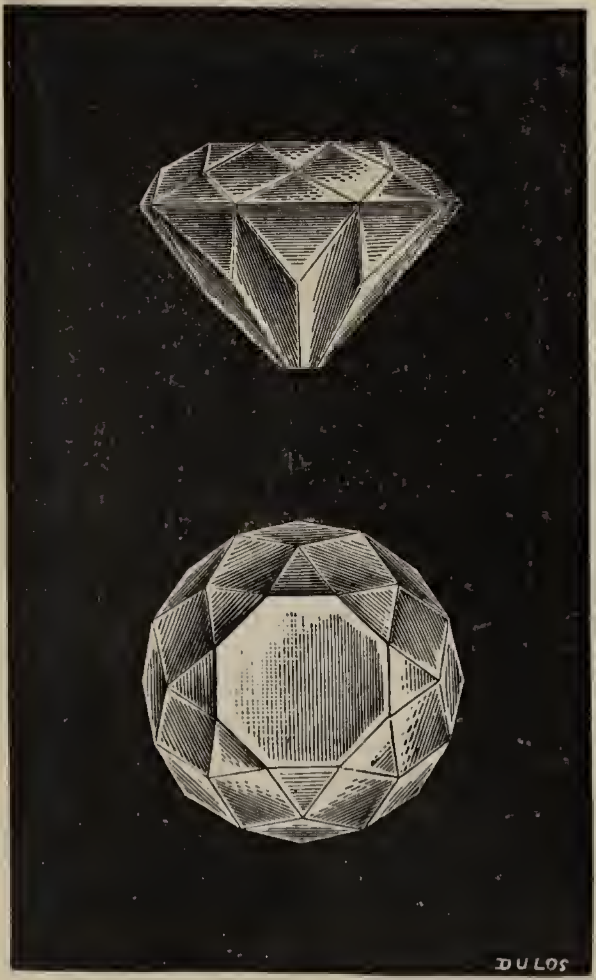


Fig. 67.

Its great power to refract light and its wonderful hardness are characteristic properties: these are the properties which render the gem so valuable in the arts. The light flashing from the different sides of the crystal makes it the most brilliant of ornaments: its extreme hardness renders it valuable in the construction of pivots in delicate instruments where friction is to be avoided.

There are two principal forms of this most elegant jewel. They are called the *brilliant* and the *rose*. The first of these is regarded as the finest. Its shape is well shown in Fig. 67. Look at the gem sidewise, and it appears as seen in the upper part of the picture; look at the top of it, and it appears as seen in the lower part. The form of the *rose* is seen in Fig. 68; a side view above and a top view below.

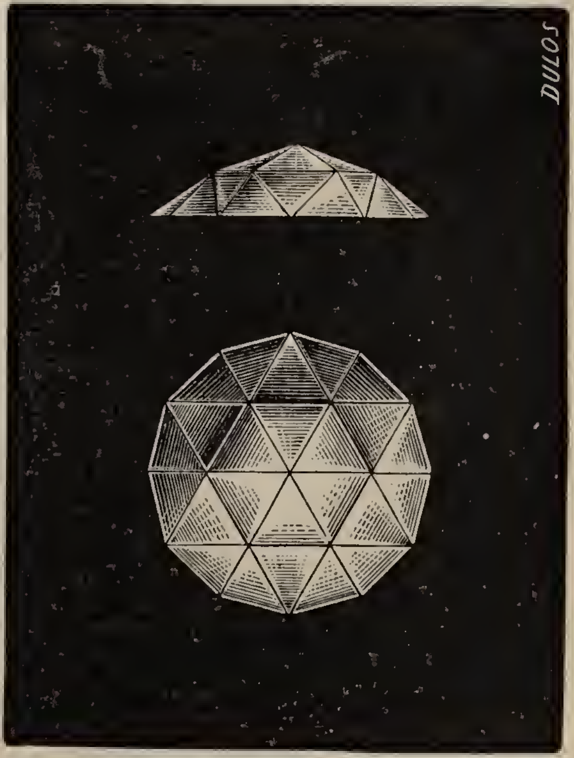


Fig. 68.

Value of the Diamond. — The diamond is the most highly prized of precious stones. Its value depends on its weight, and its freedom from color and from flaws. The weight of a diamond is described as so many *carats*, a carat being .2054 grm. (3.17 grains).

One of the finest diamonds in the world is known as the Pitt or Regent diamond, which weighs 136.25 carats, and is worth £125,000, equal to about $625,000. The largest diamond in Europe weighs 194¼ carats: it belongs to Russia, and is set in one end of the emperor's scepter. The famous Koh-i-noor is the property of the English crown. Originally it weighed 186 carats, but its weight has been reduced to 106 carats by the recutting to which it has been subjected.

163. Graphite. — Graphite is another crystalline form of carbon. It is also called plumbago, and is still more familiarly known as black-lead. It is taken from the earth in large quantities for the manufacture of lead-pencils. It has a shining steel-gray color, and it is a good conductor of electricity, being in these respects much like metals.

164. Allotropic Forms. — Charcoal, the diamond, and graphite, are but different forms of the element carbon. They differ in hardness, in color, in weight, and in many other physical properties. They are alike infusible, alike able to resist the action of substances which attack most other bodies, alike in being combustible, and alike in yielding the same substance (carbon dioxide) when burned. That they are also alike in being of vegetable origin, is believed; but of the mode in which the diamond and graphite have been made from vegetable matter, little, if any thing, is known.

It is Found in Nature. — All the different varieties of coal are but different impure forms of carbon, the remains of a vegetation which grew ages before the period of man's creation.

Enormous beds of limestone are found on every continent, but not a molecule of limestone occurs which does not contain an atom of carbon.

And more: not a single organized body, from the lowest form of plant to the highest form of animal, exists, in which carbon is not an important element.

III. — Carbon and Oxygen.

165. Oxides of Carbon. — With oxygen carbon unites in two proportions. We have

Carbon monoxide	CO,
Carbon dioxide	CO_2.

166. Carbon Monoxide. — Carbon monoxide may be prepared by heating gently, in a flask, four or five grams of potassium ferro-cyanide with about ten times as much strong sulphuric acid. It is a gas, and may be collected over water.

Properties. — Carbon monoxide is colorless, but has a slight odor of a peculiar kind. It is very inflammable, and burns with a blue flame, which may be seen playing over the surface of a brisk and new-made coal fire.

This gas is very poisonous. Even when breathed in small quantities it produces headache and giddiness: if the quantity is increased, insensibility and death follow. Accidents, not a few, have arisen from burning charcoal in open fires in rooms. The carbon takes oxygen from air to form this poisonous oxide, which escaping into the room destroys life.

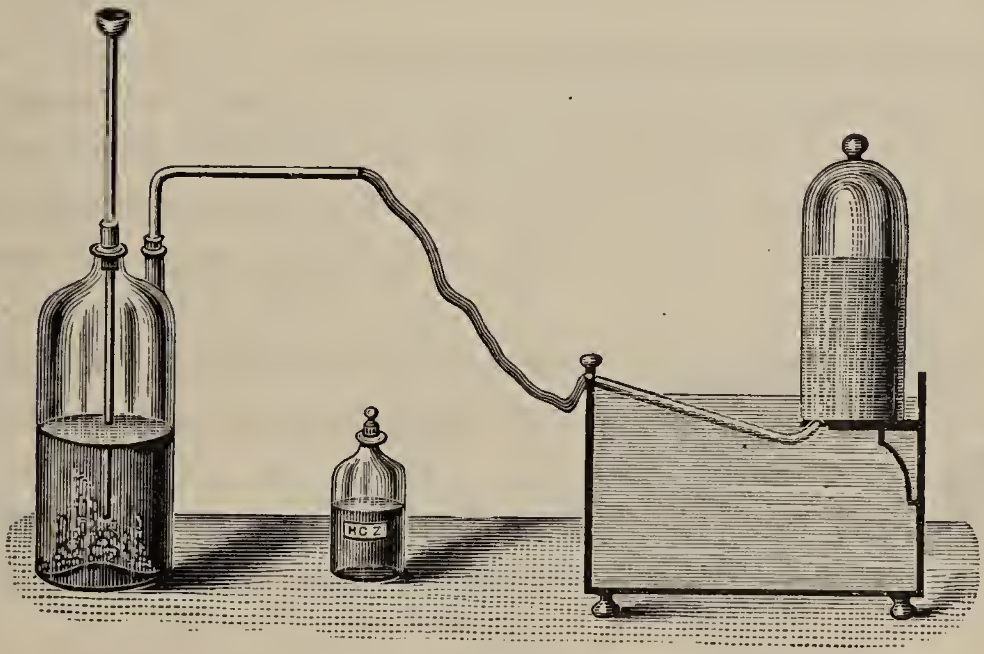


Fig. 69.

167. Carbon Dioxide. — Carbon dioxide may be obtained from marble by the following process. Small fragments of marble are put into a bottle whose cork is provided with two tubes, one reaching to the cistern, the other a funnel-tube reaching nearly to the bottom of the bottle. (Fig. 69.) Water is poured through the funnel until the lower end of its tube is covered, and hydrochloric acid is then added. A violent agitation quickly begins in the bottle: carbon dioxide is set free, and is collected in the jar over the water.

Marble is the calcium carbonate represented by $CaCO_3$. The reaction with hydrochloric acid is as follows: —

$$\begin{array}{ccccccccc} CaCO_3 & + & 2HCl & = & CaCl_2 & + & H_2O & + & CO_2, \\ \text{Calcium carbonate} & + & & = & \text{Calcium chloride} & + & & + & \text{Carbon dioxide.} \end{array}$$

Its Properties. — Carbon dioxide is a colorless gas. If a lighted taper is plunged into it, the flame is instantly extinguished; in this respect carbon dioxide resembles nitrogen. If lime-water is exposed to its action it will be turned milky: this effect can not be produced by other gases.

Carbon dioxide is much heavier than air; its density being 1.524. To illustrate: let the end of the bent tube from the bottle in which this gas is made be placed in a jar, standing with its open mouth upward In a little time a lighted taper, lowered into the jar, is quenched, showing that the jar is full of the gas. The gas remains in the open vessel as water would, and just for the same reason, — it is heavier than air. Indeed, it may be poured from one vessel to another like water: its flow can not be seen, but a lighted taper (Fig. 70) or lime-water will show its presence in the second vessel and its absence from the first.

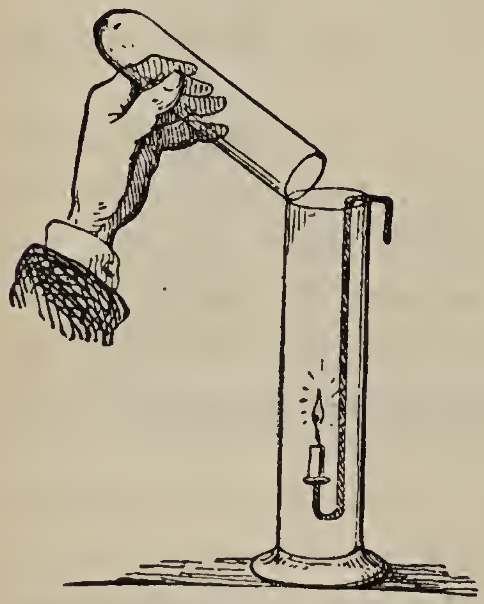


Fig. 70.

Carbon dioxide may be condensed to a liquid and even to a solid state. As steam when cooled becomes water, so this gas when made cold enough becomes a liquid: the temperature required is $-78.2°$ C. And as by being cooled still further water is frozen, so carbon dioxide may be changed from the liquid to the solid form. To reduce the temperature low enough, the liquid oxide is suddenly exposed to air at ordinary temperature and pressure: it evaporates so swiftly that the large amount of heat taken up by the change from the liquid to the gaseous form, leaves the rest of the liquid so cold that it freezes.

Gases may be liquefied, not only by cooling them: the same effect may be produced by pressure. Thus steam, whose temperature is kept up to 100° C., will be changed to water by pressure: so carbon dioxide, by a greater pressure, may be liquefied. At a temperature of 0° C. this requires a pressure of thirty-six atmospheres, equal to 540 pounds to the square inch.

168. Solubility. — Carbon dioxide is quite soluble in water. This liquid at 15° C. dissolves its own volume of the gas. At lower temperatures it will dissolve more, and at higher temperatures less.

The quantity which water will absorb, not only depends on its temperature, but also on the pressure to which it is subjected. If under twice the pressure of the atmosphere, the quantity of gas dissolved is doubled; and in all cases the quantity of carbon dioxide dissolved by water will be directly proportional to the pressure upon it.

Soda Water. — The refreshing summer drink known as "soda-water" is nothing but common water holding carbon dioxide in solution. The violent foaming of soda-water, when drawn from the "fountain," is due to the escape of carbon dioxide gas, which it has been made to dissolve by high pressure, but which it can not hold when under only the pressure of the atmosphere.

Water so highly charged with carbon dioxide is sometimes also called *seltzer* water. The true seltzer water comes from the celebrated spring of this name, and contains in solution many other things beside carbon dioxide. The artificial waters are greatly inferior to the natural waters, which are found, as at Saratoga, in such abundance. Yet they are wholesome and everywhere highly esteemed.

169. Carbon Dioxide and Respiration. — The presence of this gas in any considerable proportion is injurious to animals breathing it; and, in sufficient quantities, it produces death by suffocation. It is, however, generally admitted that the gas does not possess the poisonous character once attributed to it. It is given off in the breath; and the air in unventilated rooms quickly becomes charged with it, and unfit for respiration. Still the unpleasant odor of such air, and its poisonous effects when breathed, are believed to be due to other gases furnished by the breath, rather than to carbon dioxide.

Test for its Presence. — This gas is so heavy that it is likely to accumulate in wells, in cellars, and in coal-pits. Indeed, it is often found in coal-mines: the miners call it *choke-damp*. Before entering a place where its presence may be suspected, let a lighted candle be introduced. If the candle continues to burn, the air is pure enough to be breathed; but if the flame is extinguished, life would be also.

170. Carbonic Acid. — Carbon dioxide is an anhydride. With water it combines, and forms carbonic acid, H_2CO_3.

$$CO_2 + H_2O = H_2CO_3.$$

This acid can not be obtained pure and free; but its presence in water, which has dissolved CO_2, is demonstrated by adding a little blue litmus solution, which becomes red.

IV. — The Group.

171. The Quadrivalent Group. — Carbon and silicon are the two quadrivalent non-metals. One atom of carbon can hold four of hydrogen in combination; one of silicon can do the same. The symbols of these compounds are CH_4 and SiH_4. Carbon forms a host of other compounds with hydrogen, some of which will be noticed in due time.

Other Properties. — These two elements have each three allotropic forms. The common form of carbon is charcoal: its crystalline forms are graphite and the diamond. So the common form of silicon is a dark-colored solid, which may be changed to two crystalline forms, like graphite and diamond.

Both these elements are very hard, very infusible, and able to resist the action of chemicals in greater degree than most others.

The compounds of carbon are especially numerous among organic substances, while those of silicon are to be found almost wholly among minerals.

REVIEW.

I. — SUMMARY OF PRINCIPLES.

172. Silicon and carbon are the only quadrivalent non-metals.

Each exists in three allotropic forms.

Silica is the compound of silicon with oxygen, represented by the formula $Si\,O_2$. It is very abundant in the earth.

Silicic acid is represented by $H_2\,Si\,O_3$. The salts of this acid are silicates, and are very numerous.

Glass is a silicate containing two or more metals. It is made by melting together silica and such bases as soda, potash, and lime. The character of the glass depends on the bases used.

Carbon is found in the form of coal, of diamond, and of graphite.

With oxygen, carbon forms two oxides, carbon oxide or monoxide, and carbon dioxide. They are represented by $C\,O$ and $C\,O_2$.

Carbon oxide is a colorless, combustible, and very poisonous gas.

Carbon dioxide is also colorless, but not combustible. It does not support respiration.

It may be obtained by the action of any strong acid upon a carbonate. Hydrochloric acid and marble are conveniently chosen.

This gas is soluble in water, the quantity depending on the temperature and the pressure.

The presence of carbon dioxide may be determined in wells and mines by means of a candle-flame, which it will extinguish.

The test used by chemists is lime-water, which the gas renders milky.

Carbonic acid is represented by $H_2 C O_3$. The salts of this acid are the carbonates, and are very abundant in nature. They are among the useful substances in the laboratory also, and in many arts.

II. — EXERCISES.

Describe the element silicon.

What is silica? Where is it found? What is silicic acid? What are silicates?

From what is glass made? Give a brief description of the process.

Describe the materials used in making glass.

Name four varieties of glass. Of what different materials are they made?

Describe the process of melting.

Of blowing. Tell how a pane of glass is made.

Describe the process of annealing.

How many ways of making charcoal are given? Describe the first. The second.

What is said of the abundance of carbon?

What properties make the diamond so valuable? What is said of its occurrence? In what two forms is it cut? What is a carat? Mention a few of the largest diamonds.

What is said of graphite?

Name the three allotropic forms of carbon. How do they differ? In what respects are they alike?

Name the compounds of carbon and oxygen. Give the formula for each.

Describe carbon oxide briefly.

Of what is carbon dioxide composed?

How may this gas be obtained?

In what respects does it resemble other gases? In what action does it differ from all others? How may we show it to be heavier than air? How can this substance be obtained in a liquid state? In a solid state? What is the effect when this gas is breathed?

What is carbonic acid?

What is the quantivalence of silicon and of carbon? How are these elements related by other properties?

SECTION VII.

COMBUSTION.

1. — Nature and Products of Combustion.

173. Combustion. — Combustion is a mutual chemical action, generally between oxygen and some other substance, and which, when rapid enough, evolves heat and light.

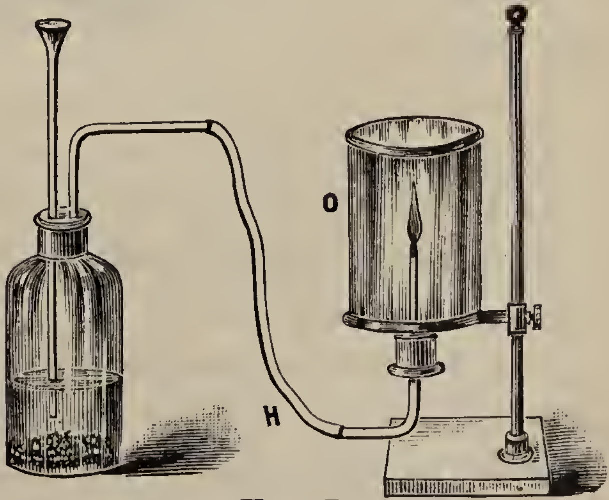


Fig. 71.

When a substance will burn in air, it is said to be *combustible:* the air at the same time is said to be the *supporter of combustion.* But really there is no difference in the part played by the two things in the action: the chemical change is mutual.

Illustrated by Experiments. — Experiments may be made to illustrate this fact. Fig. 71 represents a jar of oxygen, with a jet of hydrogen burning in it. It is the hydrogen which appears to burn. In Fig. 72 oxygen is flowing from a jet, and burning in a vessel kept full of hydrogen: in this case the oxygen appears to burn. The truth is, that in both cases alike the two gases are combining to form the liquid water.

Take an alcohol lamp, and, by spreading its wick, make a large flame. The center of this flame is dark, and is filled with the vapor of alcohol. Let the oxygen in a gas-bag be forced out by a very gentle pressure, and while it is thus flowing in a gentle stream let the end of the jet-pipe be slowly pushed into the edge of the dark center of the alcohol flame. The jet of oxygen will take fire while entering the flame, and then continue to burn with a very distinct and pretty light. In this little flame the oxygen seems to be the combustible, and the alcohol vapor the supporter; while in the large flame around it the alcohol seems to be the combustible, and the oxygen of the air the supporter. The fact is, both are equally combustible.

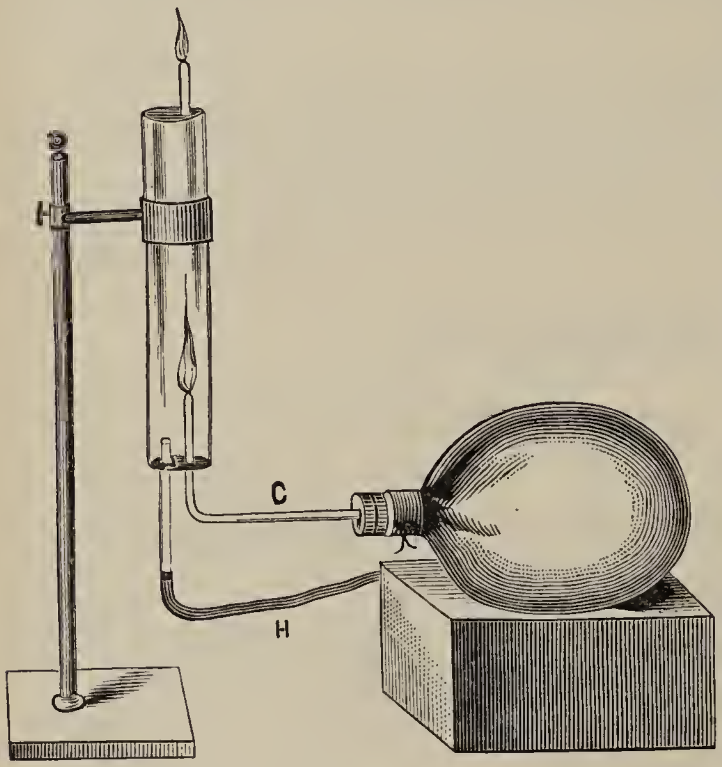


Fig. 72.

174. Combustion is an Oxidation. — In all ordinary combustion oxygen is one of the active substances. When wood burns in air, it is because the carbon and hydrogen of the wood combine with the oxygen of the air. Few, indeed, are the exceptions to this: the burning of a wax taper in a jar of chlorine is one.

The products of combustion must, therefore, be oxides.

Evolving Light and Heat. — The evolution of light and heat accompanies all familiar cases of burning; but the same chemical action, going on more slowly, gives off heat without light. An iron wire, when burned in moist oxygen, produces a splendid light, but when allowed to slowly rust in air, no light is ever seen. The chemical action is the same in both cases, — oxygen combines with the iron, and they form an oxide. Moreover, to change the iron to an oxide takes the same amount of oxygen in the two cases; and, still further, the same aggregate quantity of heat is given off. Both are cases of combustion; the one being rapid, the other slow.

The *amount* of heat depends upon the amount of material taking part in the action: the *intensity* depends upon the rapidity with which the action goes on.

175. The Kindling Temperature. — Thin shavings of the dryest pine wood, as is well known, will rest in air for any length of time unburned; but when heated by a burning match how quickly do they disappear! The match, by its heat, simply raises the temperature of the wood until, having reached a certain point, the oxygen of the air begins rapidly to combine with its elements. The temperature at which a substance begins to burn is tolerably constant, and it may be called its KINDLING POINT. The kindling point varies greatly in different substances: phosphorus, for example, inflames sometimes by the gentle heat of the hand; while sulphur must be heated to about 560° F. before it will take fire, and ordinary fuels to about 1,000° F.

Temperature kept up by Chemical Action. — But let the burning once begin, and the kindling point is kept up by the combustion itself: a match may kindle the fire which destroys a city, but if by any means the burning body can be cooled below the kindling point the fire is quenched. Over the flame of a gas-jet press down a sheet of wire-gauze

(Fig. 73): the gas goes through the gauze, the flame does not. The cold metal reduces the heat, cooling the gas below its kindling point. The metal finally becomes red-hot, after which the flame freely passes it.

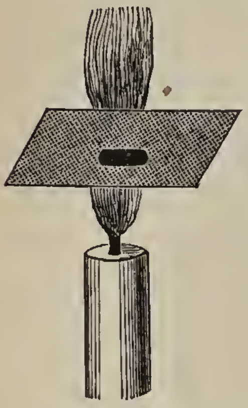


Fig. 73.

Upon this principle the "safety-lamp" is made (Fig. 74). It consists of a wire-gauze cylinder completely enveloping the burner of the lamp. The combustible "fire-damp" of a mine may burn some time inside the cylinder, *warning* the miners of its dangerous presence, before the cold meshes of the gauze will allow the flame to pass out and explode the mine.

176. The Composition of Fuel. — All fuel is of vegetable origin. That wood and charcoal are so, is evident; not less so are all forms of coal found in the earth. They are the remains of an ancient vegetable or forest growth, which, under the influence of heat and pressure, has at last parted with its oxygen and hydrogen, while the carbon still remains. The resins, oils like petroleum, and coal-gas, more rarely used for fuel, are substances which have also come from the decomposition of plants. But all these substances are compounds of carbon and hydrogen in different proportions. In all ordinary fuels, carbon and hydrogen are the chief constituents.

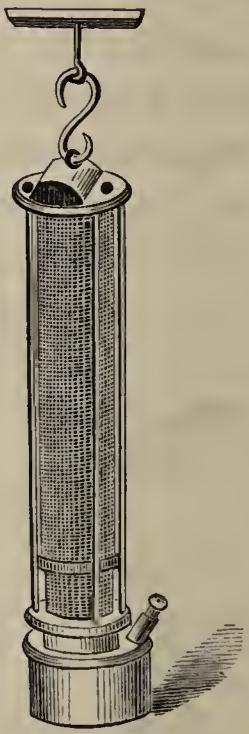


Fig. 74.

The Products of Combustion. — Upon the bottom of a plate or pan stand a burning wax taper, and cover it with a glass jar (Fig. 75). The flame soon grows dim, and is finally extinguished. Let the taper be removed, and a small quantity of lime-water poured into the jar. The milkiness of the fluid shows the presence of *carbon dioxide*. The carbon of the wax has united with the oxygen of the air, and carbon dioxide is a product of the combustion.

Again, over the flame of a burning gas-jet hold a cold glass jar. The sides of the jar are quickly dimmed with dew, which must have been formed by the hydrogen of the gas uniting with the oxygen of the air.

Water and carbon dioxide are the important products of all ordinary combustions.

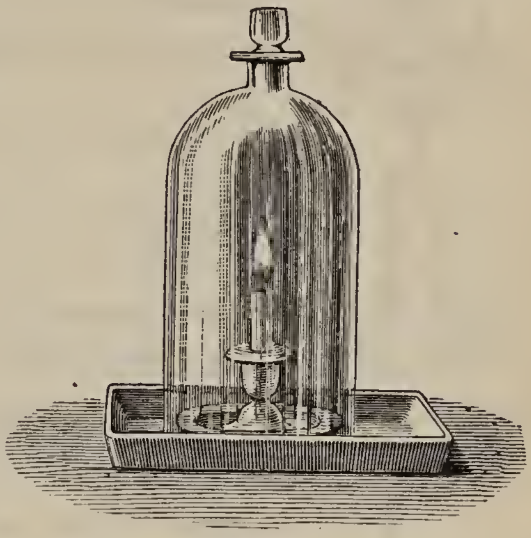


Fig. 75.

177. Gases only burn with Flame.—The flames which accompany the combustion of most substances are due to burning gases. The flame of a candle is as truly a gas flame as is the flame of a jet of illuminating gas. The solid wax or tallow in the wick must be first melted and then vaporized by the heat of the match before it will take fire. The gas once lighted causes a flame, whose heat melts the wax below; and the liquid, lifted by capillary force through the wick, is changed to vapor as fast as it is needed in the flame. Take the wax in a test-tube with a jet-pipe through its tight cork, and heat it. The wax melts, then boils, and the vapor at length issues from the jet, when, if touched with a lighted match, it burns (Fig. 76) with exactly the same kind of flame, only less steady.

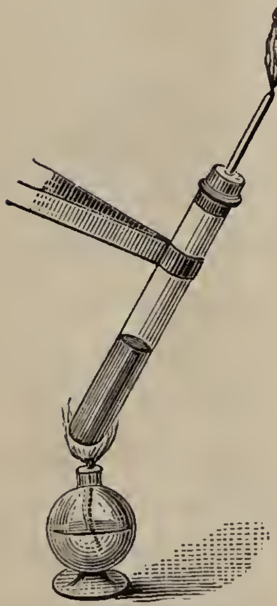


Fig. 76.

Alcohol and oil burn with flame only after heat has changed them to vapors. Bituminous coal and wood burn with flame because they contain substances which the heat applied can vaporize; while hard coal burns with no flame, because it gives off no gases when heated.

The fine blue flame which is often seen over an anthracite fire has been already explained as due, not to the combustion of the coal, but to the combustion of carbon oxide. The carbon at first takes oxygen, and becomes carbon oxide; and then this gas, coming in contact with oxygen at the surface of the fire, combines with additional oxygen, and becomes carbon dioxide.

$$CO + O = CO_2.$$

This combustion of the gas, CO, yields the blue flame.

178. Three Parts to a Common Flame. — The dark center and the luminous cone of common flames have been noticed by us all. This dark center or nucleus consists of the gases formed from the fuel; but they are not in a burning condition. The luminous envelope consists of these gases combining with the oxygen of the air. In this part of the flame only does combustion occur. Outside of this, is an envelope made up of water-vapor and carbon dioxide produced by the burning.

Fig. 77.

The three parts may be easily seen in a common candle flame (Fig. 77).

The *nucleus is not burning:* for the end of a match, plunged to the center of the flame, does not burn while there; it takes fire while coming out. And even gunpowder may rest in this center of a flame unharmed! This can be easily proved. For this purpose invert a dinner-plate or common bowl upon the table: its bottom is a very shallow dish which will hold a small quantity of alcohol. Put some gunpowder on the end of a small cork, and place it at the middle of the plate, and then touch the alcohol with a match. A large flame is formed, in the center of which the powder remains unburned (Fig. 78), until some air-current wafts the flame, or until the alcohol is nearly consumed.

The *burning takes place in the luminous envelope.* We are taught this by the following experiment: over the flame of an alcohol-lamp suddenly lower a sheet of writing-paper, holding it for a moment across the middle of the flame (Fig. 79). Remove it, and a scorched ring will be seen; the paper is burned just where it was in contact with the luminous envelope.

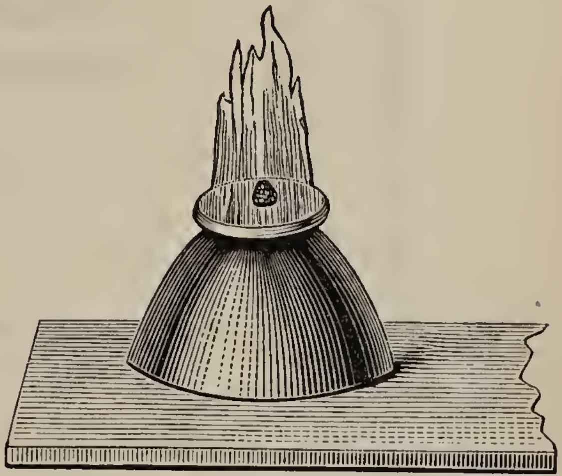


Fig. 78.

Hold a cold glass plate in place of the sheet of paper for a moment on the flame: a ring of dew is condensed upon it where it comes in contact with the non-luminous envelope.

II. — Light and Heat.

179. The Practical Value of Combustion. — Most chemical processes are valuable on account of their products; but the products of combustion, water and carbon dioxide, are of no practical value as products, and combustion is never resorted to for their sake. This process is useful simply for the light and heat which accompany it.

Fig. 79.

180. Conditions for Production of Light. — Witness the following experiments. Let a glass jar be inverted over a burning taper (Fig. 80). The flame continues for a time, and then dies. It goes out because its supply of oxygen is exhausted. Nor does it help the matter much to raise the jar a little space above the plate on which the taper stands, for the jar remains full of impure air which keeps the pure air out.

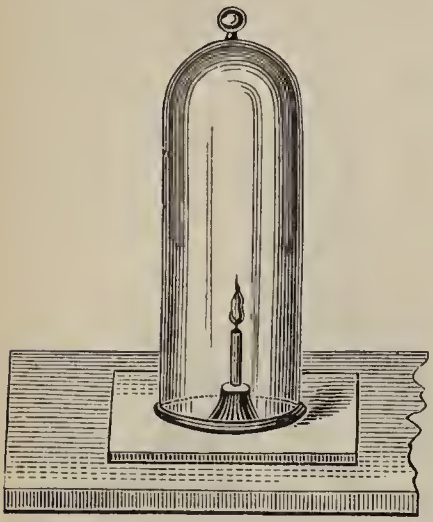


Fig. 80.

Now again, put the taper at the bottom of a jar whose open mouth is upward, and, if necessary, partly cover it (Fig. 81). The flame dies almost as quickly as before, and for the same reason. We learn from these experiments that a supply of fresh air is absolutely needed for combustion.

A Full Supply. — Go further: take a lamp-chimney, and so place it that a *very small* opening admits air at the bottom: the flame does not die, but burns dimly. And now, again, almost cover the top of the chimney, leaving the bottom open: the taper burns, but with a dim or smoky flame. Finally, open both top and bottom. A current of air passes freely up through the chimney, and we see the taper burning brightly. These experiments teach us that a *full supply* of air is necessary to produce a luminous flame.

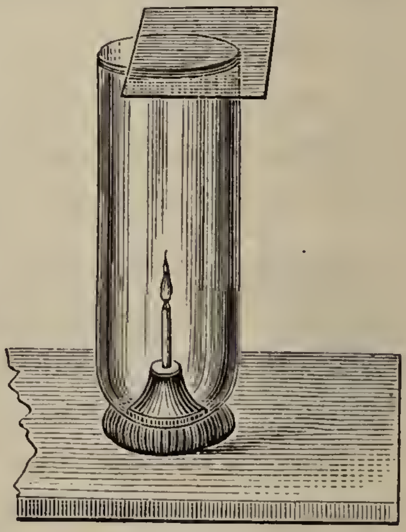


Fig. 81.

To the Surface of the Flame. — But it will not do to mix the air with the combustible gas: such a mixture burns with an almost lightless flame.

This effect is shown in the Bunsen's burner (Fig. 82). The gas is brought from the chandelier by means of the rubber tube, and issues from a jet-pipe inside the tube A. Just below the end of this jet are two large holes on opposite sides of the tube, one of which is shown at *b*. A movable collar *c*, with corresponding holes, may be turned around the tube so as to open and close these holes at pleasure. Now a constant supply of *gas* from the jet, and of *air* entering at the holes, keeps the tube (A) full of a *mixture*, which is lighted as it issues from the upper end. This burning mixture gives an intense heat, — a glass tube may be quickly softened or melted by it; but the light is very feeble.

The instrument is much used in the laboratory for heating glass vessels, for which it is well adapted, because its flame is without smoke and deposits no soot.

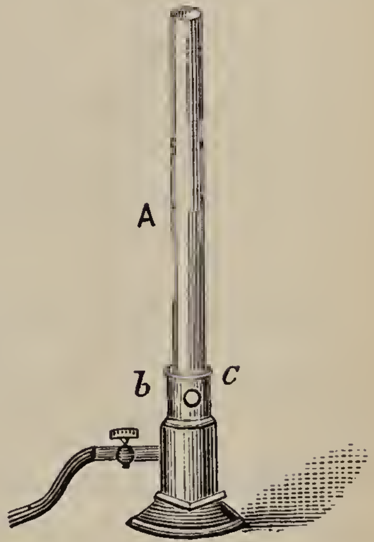


Fig. 82.

A full supply of air must be brought in contact with the *surface* of the gas-jet if the greatest light is to be obtained.

This is done in the Common Lamp. — The oil or kerosene lamp is too familiar to need description. Notice that the *flat* wick usually employed, exposes a large surface to the air, and that the chimney, open at the top and bottom, allows a constant current of fresh air to pass up through it. By this means an abundance of air is brought in contact with the flame, at its surface only, and the bright light is the result.

The Argand Burner. — In the argand burner an artifice is resorted to, by which the surface of the flame exposed to air is increased. The wick is a hollow cylinder, and the air passes up through the inside of it as well as around its outside. The gases from the wick expose a double surface to the air, and give off a greater light accordingly.

The Gas Burner. — The burner of a gas-chandelier is so made that the gas escapes in a fan-shaped jet. This is done in many ways. In the end of some burners we may notice two small holes; and by putting pins into these we find them to be the ends of two little tubes slanting toward each other, so that if continued outward they would meet. Now, the two jets of gas from these tubes strike against each other with force enough to flatten both out into a single fan-shaped jet. In this way a large surface of gas is exposed to air without making a mixture of the two substances, and the luminous flame is produced.

181. Source of the Light. — Sift some finely-powdered charcoal into the almost non-luminous flame of a Bunsen's burner, and we find that the flame becomes more or less luminous. What happens is this: The small particles in the hot flame are heated to whiteness, and shine with a bright light. The light quickly vanishes, however; because when these hot particles of carbon come in contact with air, they instantly combine with oxygen, and become carbon dioxide. The light is due, while it lasts, to the *intensely heated solid particles*.

Doubtless on this principle we may, in part, account for the light of common flames. We know that the burning gas is being decomposed, and that one of its constituents is carbon. We know, too, that carbon is a solid, even at the highest heat, but that heated in air it combines with oxygen. At the instant, then, when set free from the burning gas, and before its union with oxygen, each little molecule of carbon is heated white-hot, and, shining brightly, adds its mite to the light of the flame.

The Oxyhydrogen Light. — The luminous power of solid bodies in a heat-flame is illustrated by the so-called oxyhydrogen or lime light. It is made by simply placing a piece of lime in the flame of the oxyhydrogen blow-pipe. (Fig. 84.) So very intense is this light, that the eye is blinded by its direct rays; and, used in light-houses, it has been seen miles away at sea.

Observe now: lime remains solid, even in the most intense heat of the blow-pipe; but the burning mixture gives heat enough to raise the lime to a white heat, and in this condition it shines with a blinding light.

Light Due also to Dense Vapors. — But that all the light of flames is not to be explained in this way, is seen clearly from the fact that in some of the most dazzling there is no substance present which can remain solid at the temperature of the flame. Remember the blinding light of phosphorus burning in oxygen-gas. "Now, phosphoric anhydride, the product of this combustion, is volatile at a red-heat; and it is, therefore, manifestly impossible that this substance should exist in the solid form at the temperature of the phosphorus flame, which far transcends the melting point of platinum." (*Dr. Frankland.*) Many other examples of a similar kind might be given. The light of such flames, and doubtless much of the light of all flames, is due to intensely heated dense vapors.

182. Conditions for the Production of Heat. — If heat is the object sought for in the process of combustion, two conditions should be fulfilled. There should be a *proper supply of air* provided, and this supply should be *thoroughly mixed with the fuel.*

Object of Mixing the Air and Fuel. — By mixing the air with the fuel we enable the combustion to take place in every part at once. It is therefore more rapid and complete, and the heat is more intense.

In the common stove and furnace, for example, the air entering at the draught *mixes* very thoroughly with the fuel; and the heat is much more intense than it would be if the two came in contact only at the surface.

The Proper Supply of Air. — By the proper supply of air we mean just so much as is needed to furnish oxygen enough to combine with all the fuel.

Too little air will leave some of the fuel unburned, and heat will be lost. Too much air will cool the fire by carrying away some heat into the draught.

To Calculate the Quantity. — If we know the composition of the fuel, we can then calculate the quantity of air needed for its combustion. Suppose, for example, we have ten liters of hydrogen, and wish to get the most intense heat by burning it: how much air must we give it? We first find how much oxygen it requires to combine with it. We know that two volumes of hydrogen take one volume of oxygen to form water. Hence our ten liters of hydrogen must have five of oxygen. We next find how much air is required to furnish the necessary oxygen. We know that oxygen constitutes very nearly one-fifth of the air; hence it will need *twenty-five liters* of air to give the five liters of oxygen to burn the ten of hydrogen.

Whatever the composition of the fuel may be, knowing it we can calculate just how much oxygen will exactly oxidize all its constituents, and then the quantity of air required to furnish this amount.

183. The Intensity of the Heat. — The intensity of heat produced depends on the rapidity of the combustion. A brisk fire is a hot fire. A smoldering fire may burn longer, and in the end give out as much heat, but at no time is it able to scorch, consume, or fuse with vigor.

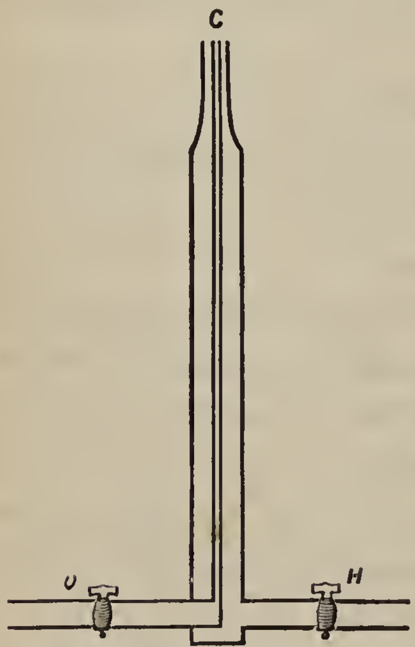


Fig. 83.

The Oxyhydrogen Blow-pipe. — A mixture of hydrogen and oxygen, in the right proportions to form water, burns with heat of surprising intensity. By means of the oxyhydrogen, or compound blow-pipe, this flame may be obtained with little danger of explosion. Notice Fig. 83, which shows a section of one form of the jet, and Fig. 84, which shows the instrument in use.

The gases are stored in separate bags, or other reservoirs.

Hydrogen passes from one of these into the larger or outside tube of the jet, and flows out at *c;* while oxygen from the other passes through the inside tube, and out at the same point. It will be seen that the two gases *mix* just at the point of the jet, and that here the combustion takes place. This is called the *concentric jet.* In another form the jet contains no inner tube. The two gases are brought into the base of the jet, where they mix, and the mixture issues from the tip. This is called the *mixed jet.*

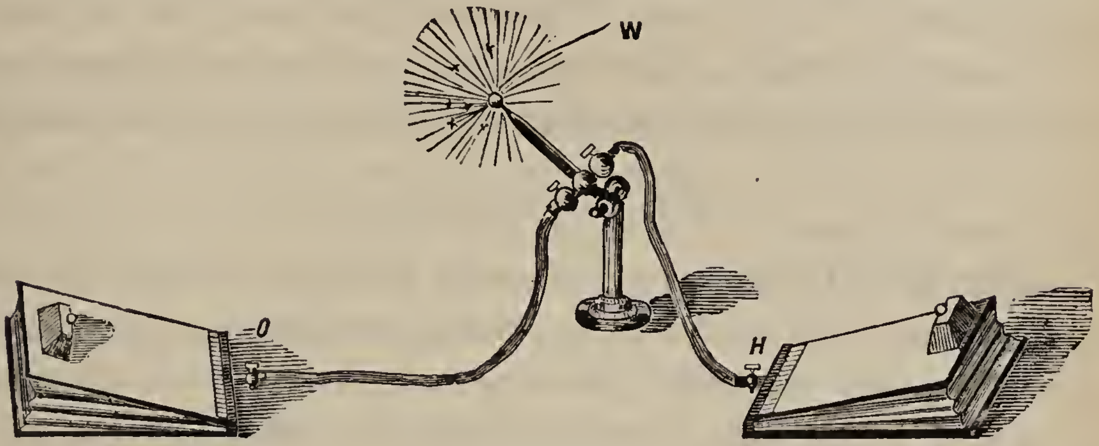


Fig. 84.

The heating effects of this flame are astonishing. Zinc and antimony are vaporized by it. Iron and steel (Fig. 84) burn in it like thread in a lamp-flame. Platinum, and even quartz and other rocky matter, may be melted or softened. Only the heat of the electric arc exceeds this in intensity.

184. The Quantity of Heat. — It has been proved by numerous experiments that the *quantity* of heat evolved by combustion does not depend on the rapidity of the action, but only on the *weight of the material burned.* The burning of a pound of hydrogen will yield precisely the same amount of heat, whether it burn swiftly or slowly. A pound of carbon will yield always exactly the same quantity of heat when it is changed into carbon dioxide by combustion. If, however, the carbon burns to carbon monoxide, the pound will yield less, because the quantity of oxygen burned is less. The general principle may be stated as follows: —

The same weight of the same substance will invariably yield the same amount of heat when the products of combustion are the same.

A slow fire will in the end give just as much heat as a brisk one, if the same quantity of fuel and air are consumed.

III. — Respiration.

185. Respiration. — The walls of the air-cells of the lungs are covered with a net-work of minute capillary blood-vessels. Into the air-cells successive portions of fresh air enter, to be at once thrown out again, while, at the same time, the impure blood of the system is constantly coursing through these vessels. We must consider what changes are being made, first in the air, and second in the blood.

The Chemical Changes. — If one breathes through a tube into a vessel of lime-water, its milky color soon shows the presence of *carbon dioxide.* If he breathe on a cold surface, the moisture condensed there shows the presence of *water-vapor.* To determine what other substance is exhaled, we may make an experiment, which is represented in Fig. 85. A bell-jar is placed with its open mouth in a vessel of water. The jar is provided with a cork,

Fig. 85.

through which passes a tube which joins it to the bottom of a tall jar containing caustic potash, KHO. The top of this potash jar is also provided with a tube terminating with a mouth-piece. Applying the lips to this mouth-piece, the air in the jar may be drawn into the lungs, and then returned to the jar. Let this be done two or three times. In going back and forth between the jar and the lips, the air must pass over the caustic potash, which will effectually remove from it, both the water-vapor and the carbon dioxide. Let the flame of a taper be afterward plunged into the jar, and it will be extinguished, showing the absence of *oxygen* and the presence of *nitrogen*.

From all these experiments we conclude that the air while in the lungs gives up its oxygen, and receives carbon dioxide and water-vapor.

Could we examine the blood, we should find that while in the lungs its color changes from purple to bright red, due to the loss of carbon dioxide and water-vapor, and the receipt of oxygen.

A Process of Slow Combustion. — Now, how can these changes be explained? It has been found that particles of the body itself are constantly being worn out. Not an act can be done, a word spoken, nor can a thought occur, without the disintegration of some portion of the organs. These waste particles are in great part thrown into the blood: they are its impurities. These particles are composed chiefly of carbon and hydrogen. The pure blood is charged with oxygen; and, as it flows, this gas combines with the carbon and hydrogen of the waste, at all points in its course; and the carbon dioxide and water-vapor, thus formed, pass into the lung-cells, to be finally exhaled from the system. The action is a process of slow combustion. The waste particles are the fuel, oxygen is supplied, and carbon dioxide and water-vapor are the products.

Moreover, it is by the heat evolved in this combustion that the body is kept warm.

Large Quantities of Air Spoiled. — It will be at once seen that the air of our rooms is being constantly made unfit to keep up this action by which the blood is purified. In the first place, its oxygen is being taken out; and, in the second place, impurities are being thrown into it. Not only does it receive the worn-out matter of the system from the breath, but we may here add that other portions of waste, even more offensive, are incessantly being given into it from the body by perspiration. The total amount of impurity thus added to the air of a close room is frightful; since, upon the average, about twenty thousand cubic inches of air pass through the lungs of a single person every hour!

Hence the Need of Ventilation. — As the flame of a taper dies in a closed jar, so a human being would die if confined long enough in an air-tight room. As the flame flickers and burns dim when a small supply of air is furnished, or the products of combustion are not removed, so the life of human beings flickers and grows feeble in rooms where fresh air is not supplied.

IV. — DECAY.

186. Decay. — The decay of wood or other vegetable matter is a slow process of combustion. By the gradual loss of carbon dioxide and water, the decaying wood is finally changed into a brown or black mold, to which the name *humus* is given.

The Decay of Wood. — If fine sawdust is thoroughly moistened, and put into a stoppered bottle containing air, and afterward exposed to a temperature of about 16° C. (60° F), it will, after a long time, be found partially decayed. By the usual tests it will be found that the air in the bottle has given up a part of its oxygen, and received carbon dioxide in return.

Wood, of whatever kind, when exposed to the continued action of moist air and warmth, will, like the sawdust in the experiment, be gradually decomposed. Moisture and warmth are essential to decay; since it is found that wood, exposed to the cold of the arctic regions, or to the dry air of Egypt, will remain, even for centuries, in good condition.

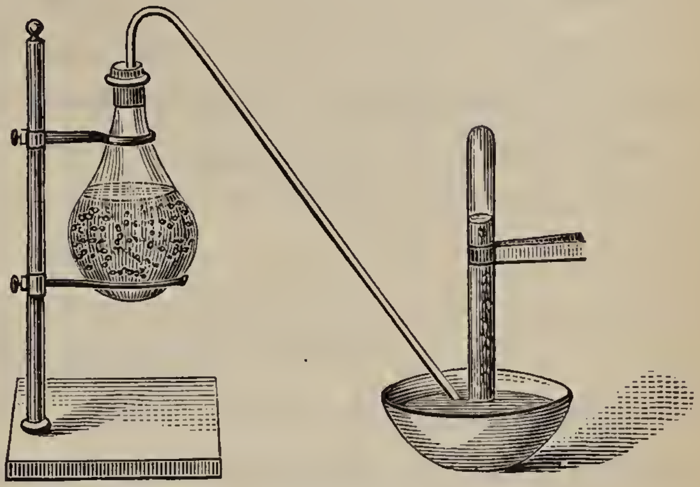


Fig. 86.

Other Vegetable Matter. — If a flask (Fig. 86) partly filled with peas and water is furnished with a bent tube reaching over to an inverted larger tube filled with water, there may, after a time, be seen bubbles of gas rising into the tube. This gas, if tested, will be found to be carbon dioxide; and when the peas are examined they will be found to be partially decayed.

A Slow Combustion. — Now, the experiments with the sawdust and the peas are simple illustrations of what occurs whenever any kind of vegetable matter is long exposed to the action of air under the influence of moisture and warmth. They are slowly decomposed: a part of their carbon and of their hydrogen unites with oxygen to form carbon dioxide and water; while the rest of these elements, in combination with oxygen, remains in the form of a loose, solid, black mold.

The chemical action in the process of decay is, clearly, very similar to that of combustion. Indeed, it differs from combustion. chiefly in being less rapid. Heat is evolved in decay as in combustion, and from equal quantities of material the same amount. In some rare cases a feeble light is also given off by decaying vegetable matter. Decay is to be considered as but a process of slow combustion.

The combustion is, however, very incomplete. A large part of the carbon, and much of the hydrogen, of the decaying body, is left unoxidized.

Humus. — The brown or black mold that is left after the decay of plants is called Humus. This term, however, is not the name of a single substance, but rather of a mixture of several compounds of carbon, hydrogen, and oxygen in various proportions. Humus gives to fertile soils their rich brown or black appearance.

REVIEW.

I. — SUMMARY OF PRINCIPLES.

187. Combustion is a mutual chemical action, generally between oxygen and some other substance, and which, when rapid enough, evolves heat and light.

A certain temperature must be reached before a substance will kindle; but once started, the chemical action will produce heat enough to keep the action going.

Every combustible has a kindling point peculiar to itself. Phosphoretted hydrogen kindles on contact with air at ordinary temperatures. When this is the case, the burning is said to be Spontaneous Combustion. On the other hand, ordinary fuels have a kindling point of about 1000° F.

Flame is extinguished by being cooled below the kindling point of the substance. For this reason flames will not pass through small openings in cold bodies, as, for example, the interstices in wire-gauze.

All fuel is of vegetable origin, and consists chiefly of carbon and hydrogen.

Combustion is the oxidation of the elements of the fuel.

Carbon dioxide and water are therefore the chief products of combustion.

Flame is produced only by the burning of gaseous matter. A solid, burning as such, yields a glow of light but no flame.

Combustion is resorted to merely for the heat and the light it produces; not as in other chemical processes, for the material which it forms.

To give the greatest light, there should be a full supply of air to the surface of the gas-jet, as in the common lamp, the argand burner, and ordinary gas flames.

To give the greatest heat, the air should be thoroughly mixed with the burning gas.

The proper supply of air is that which will furnish just oxygen enough to unite with the elements of the fuel.

The quantity of heat depends on the weight of the material burned.

The intensity of the heat depends on the rapidity of the combustion.

When the oxidation takes place so slowly that no light is evolved, it is described as *slow combustion.*

Respiration is a process of slow combustion. The waste particles of the body are the fuel: they are oxidized by the oxygen of the air which is absorbed by the blood in its passage through the lungs: carbon dioxide and water-vapor are produced and exhaled in breathing.

The heat of this slow combustion is the natural heat of the body.

The decay of organic bodies is also an analogous process of slow combustion.

II. — EXERCISES.

What is combustion? By what experiment is it shown to be a mutual action?

Between what substances does it occur?

Do light and heat always accompany combustion? Illustrate by the combustion of iron. Upon what does the amount of heat depend? Upon what, its intensity?

Why do not substances take fire when simply exposed to air? What is meant by the term, kindling temperature? Give examples.

How is the temperature kept up? Illustrate. Describe the experiment with the wire-gauze. What does this experiment teach? What application has been made of this principle?

Of what is fuel composed? What are the chief products of combustion?

What is the origin of all fuel? How can this be true of coal? Of resins and petroleum?

Describe the experiment with the wax taper. What does it teach? The experiment with the gas-jet. What does it teach?

Do solids, liquids, or gases burn with flame? How can this be true of the candle flame? Of alcohol? Of wood? Of hard coal?

Name the three parts of a common flame. Of what is each formed? Prove that the nucleus is not burning. How may we learn that the burning takes place in the luminous envelope? How can we prove that there is water in the non-luminous envelope?

For what is combustion resorted to?

In what way is the greatest heat produced?

Describe the Bunsen's burner. Explain its action.

Describe the oxyhydrogen blow-pipe. What effects may be produced by it?

To get the greatest light from combustion, what is necessary?

Describe the first experiment with the burning taper. The second. What do these experiments teach? Describe the experiments with the lamp-chimney. What do they teach? Should the air and gas be mixed?

How is the full supply of air in contact with the surface of the flame secured in the common lamp?

In the argand-burner?

In the gas-burner?

To what is the light of a flame due?

Illustrate by the lime-light.

What reason have we to suppose that all the light is not due to solid particles? To what is it, then, due?

What kind of an action is respiration?

By what means are the air and blood in the lungs brought in contact? How may the presence of carbon dioxide in the breath be proved? Of water-vapor? How may we prove the absence of oxygen and the presence of nitrogen in the breath? What change, then, occurs in the *air* while in the lungs? What change occurs in the blood there at the same time? Explain these changes.

How, then, is the air of inhabited rooms being spoiled? How much impurity is thus being added to the air?

Why is ventilation absolutely necessary?

What kind of a process is decay? What change takes place in decaying bodies?

Describe the experiment with sawdust. How do we know that moisture and warmth are necessary to cause decay?

Describe the experiment with the peas.

What do these experiments illustrate? In what respects are decay and combustion alike?

What is left after the decay of plants?

CHAPTER III.

THE COMPOUNDS OF CARBON.

SECTION I.

GENERAL STATEMENTS.

188. Organic Chemistry. — The compounds of carbon are more numerous than those of all other elements taken together. Many of them are very complex in composition. They are, for the most part, constituents of, or products from, organic bodies. For these reasons they have usually been studied by themselves, in what is called ORGANIC CHEMISTRY.

Organic chemistry is now defined to be the chemistry of the carbon compounds.

Carbon Compounds. — By the carbon compounds we here refer to compounds of carbon containing *hydrogen*. These two elements alone unite to form a very large number of bodies, which are called HYDROCARBONS.

Another very large class contains carbon and hydrogen with oxygen; and a smaller class contains a few other elements also, among which are nitrogen and sulphur.

189. Organized Bodies. — Plants and animals are *organized bodies.* Their bodies have been nourished and enlarged by means of food taken into and distributed throughout their interior. The plant, by means of certain organs, receives the sap into its roots, and sends it to every part. Every leaf and stem is built of substance which is taken from this sap and from the air. The animal takes its food, converts it into blood, distributes it to the most minute fibers throughout, and every part of its body is built from material thus furnished.

Every distinct part of these bodies is also organized. The leaf, the twig, and the tendril are organized bodies; and so are the hairs, the claws, and the tissues of animals.

Organic Substances. — These organized bodies consist almost entirely of the compounds of carbon and hydrogen, oxygen and nitrogen. Sugar ($C_{12}H_{22}O_{11}$), starch ($C_6H_{10}O_5$), and alcohol (C_2H_6O) are three among the thousands of substances which the chemist is able to get by decomposing organic bodies. Such compounds may be called ORGANIC SUBSTANCES, in distinction from ORGANIZED BODIES from which they are obtained.

Many of these organic substances exist ready formed in the organized body, but many more do not. These last are *produced* only by the decomposition of the organized tissues.

190. Chemistry deals only with Organic Substances. — Now, it would seem that Nature has fixed a barrier, which the chemist may not pass, between organic substances and the organized bodies which they form. The chemist can, by synthesis, make a few of the simplest organic substances, and he has good reason to believe that even the most complex are produced by chemical force, governed by the ordinary laws of combination. But, on the other hand, the simplest organized body, although it consists of the same elements, is entirely beyond his reach. He can not make a single cell. By what chemistry the leaf and the flower are made, he does not know. His laboratory teems with elegant crystals whose growth he has himself guided; but his garden is filled with still more delicate forms, the secrets of whose chemistry are known only to the Divine Chemist who made the earth and the sun, and all that they contain.

SECTION II.

MARSH-GAS AND THE MARSH-GAS SERIES.

191. Marsh-Gas. — The simplest hydrocarbon known is marsh-gas (CH_4), in which one atom of carbon is united with four atoms of hydrogen.

This gas is found in nature. It is the principal part of the gas which escapes in bubbles when the muddy bottom of a stagnant pool is disturbed.

It may be collected by receiving these bubbles in an inverted bottle (Fig. 87).

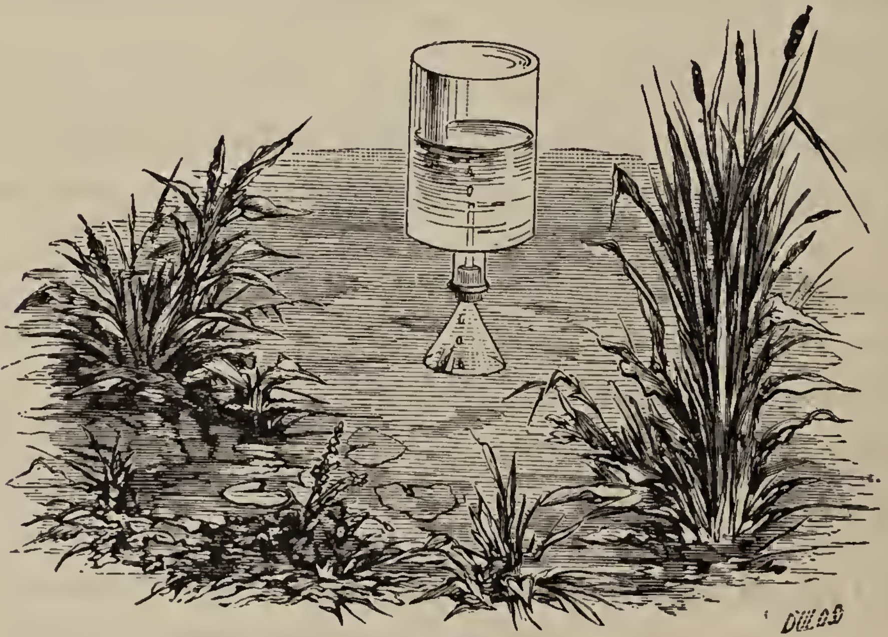


Fig. 87.

In this case marsh-gas has been produced by the decay of dead leaves, or other organic matter; and its name comes from the fact of its occurrence in such places. This same gas collects sometimes in mines; and, by explosions of which we have all heard, now and then extinguishes the lamps and the lives of the miners. On this account it has also been called FIRE-DAMP.

Properties. — Marsh-gas is colorless, and only eight times heavier than hydrogen, or a little more than half as heavy as air (.558). It burns vigorously, with a bluish flame which is slightly luminous. Its mixture with twice its volume of oxygen explodes with great violence.

Composition. — The composition of marsh-gas is represented by its formula, $C\,H_4$. That the molecule contains four atoms of hydrogen, may be proved in this way: We remember that a molecular volume is always two: hence two volumes represent a molecule of the gas. Now pass electric sparks through two volumes of marsh-gas: it will be decomposed, and will yield just four volumes of hydrogen. But the atomic volume of hydrogen is one, and hence four volumes must represent four atoms. Therefore, the molecule of marsh-gas yields four atoms of hydrogen.

Names. — Marsh-gas is the old and common name of this substance; but the chemist calls it Methane, and at other times he calls it Methyl Hydride.

This last name, methyl hydride, is intended to show that the gas is a *binary compound* of hydrogen and methyl. If we write the formula $C\,H_3\,H$ instead of $C\,H_4$ we show what that methyl is: it is $C\,H_3$, which with hydrogen forms the hydride. Chemists have much reason to suppose that the molecule, $C\,H_4$, is made up of *two parts*, viz., $C\,H_3$ and H; and the name, methyl hydride, expresses this view, the name of the $C\,H_3$ being methyl.

192. The Marsh-Gas Series. — Methane is the simplest of a large number of compounds, which contain carbon and hydrogen holding a very curious relation. The relation is this: *each one in the series may be obtained by adding* $C\,H_2$ *to the one before it.* Let us start with $C\,H_4$ in the following table, and illustrate this relation: —

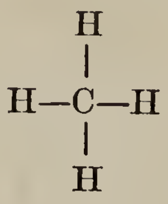
$C\,H_4$	Methane,
$C\,H_4 + C\,H_2 = C_2\,H_6$	Ethane,
$C_2\,H_6 + C\,H_2 = C_3\,H_8$	Propane,
$C_3\,H_8 + C\,H_2 = C_4\,H_{10}$	Butane,
$C_4\,H_{10} + C\,H_2 = C_5\,H_{12}$	Pentane,
etc. + etc. = etc.	etc.

All these compounds actually exist, and many more whose formulas the reader can make for himself by continuing to add CH_2.

This series of hydrocarbons, beginning with methane, and increasing in complexity by the constant addition of CH_2, is called the MARSH-GAS SERIES. They are also sometimes called the PARAFFINES.

Explanation. — Carbon alone among the elements is able to form so many hydrogen compounds. Chemists try as follows to explain this fact: *In the first place*, the carbon atom is quadrivalent; and, *in the second place*, one atom can combine with another atom of carbon, and saturate one or more of its units of quantivalence, leaving the other units to be saturated by atoms of hydrogen. This explanation is considered very satisfactory.

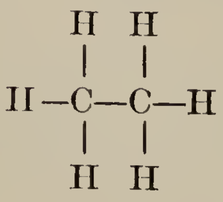
Illustration. — If we write the graphic formula for the first one in the series, it stands thus: —



$$\begin{array}{ccc} & H & \\ & | & \\ H- & C & -H \\ & | & \\ & H & \end{array}$$

. . . . . . CH_4, Methane.

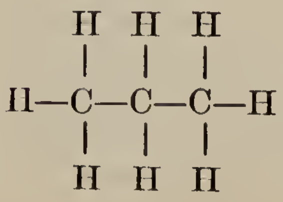
If, now, we put two atoms of carbon together, we must *add* two atoms of hydrogen in order to saturate them. Thus: —



$$\begin{array}{cccc} & H & H & \\ & | & | & \\ H- & C- & C & -H \\ & | & | & \\ & H & H & \end{array}$$

. . . . . . C_2H_6, Ethane.

If, again, we add a third atom of carbon, we must add two more atoms of hydrogen to give the saturated molecule. Thus: —



$$\begin{array}{ccccc} & H & H & H & \\ & | & | & | & \\ H- & C- & C- & C & -H \\ & | & | & | & \\ & H & H & H & \end{array}$$

. . . . . C_3H_8, Propane.

Thus we see, that, carbon remaining quadrivalent, and its atoms combining together by one unit, the constant addition of CH_2 must produce the successive members of the series.

193. Properties of the Marsh-Gas Series. — The first five members of this series are gases at ordinary temperatures. Several which follow are liquids, while from $C_{20}H_{42}$ onward in the series the substances are solids.

Each one of the liquids has its own boiling-point. Pentane, for example, boils at 30° C., and hexane at about 60° C. *The boiling-point rises about 30° C. for each addition of* CH_2.

The series is remarkably indifferent to nearly all chemical agents. Its members are all combustible, and in other chemical actions they exchange atoms of hydrogen for atoms of other elements.

194. Petroleum. — Petroleum is an oily liquid of a greenish color, found in the rocks of various parts of the world. It is especially abundant in Pennsylvania, where it is obtained in large quantities for commercial purposes.

Composition. — Petroleum is a mixture of hydrocarbons; and these are chiefly, but not wholly, of the marsh-gas series. All the members of the series from C_4H_{10} to C_9H_{20} are present.

Fractional Distillation. — Since each hydrocarbon has its own boiling-point, it is possible to separate several, when mixed, by distillation. Let the mixture be heated up to the *lowest* boiling-point, and kept steadily at that temperature; only the vapors of the one substance having that boiling-point will be driven off, and these may be condensed in a receiver. When these vapors cease to flow, let the temperature be raised up to the next boiling-point, and maintained: only the vapors of the substance having this next higher boiling-point will be driven off, and these may be condensed in a separate receiver. Similarly each one in the mixture may be obtained in a separate vessel. This distillation of a liquid at the successive boiling-points of its constituents is called FRACTIONAL DISTILLATION.

For Commercial Purposes. — Kerosene oil, naphtha, and gasoline are among the products obtained from petroleum for various uses in the arts. None of these are single compounds: each is a mixture of several hydrocarbons. They are obtained from petroleum by fractional distillation; although the temperatures used are not the successive boiling-points of its *chemical* constituents, but are arbitrarily chosen instead. On this account each product is a mixture.

The following table, from *Wagner's Chemical Technology* (Crookes), shows temperatures used, and names the commercial product obtained: —

Below 37.7° C.	Rhigoline,
At 76.6° "	Gasoline,
At 137.0° "	Naphtha,
At 148.0° "	Benzine,
At 183.0° – 219.0° C.	Kerosene.

SECTION III.

THE ALCOHOLS.

195. Common Alcohol. — When the juices of fruits containing sugar are kept warm for several hours, a peculiar change occurs. Their sweetness becomes less and less; bubbles of carbon dioxide escape; and, at the same time, alcohol is formed in the liquid. The chemical action, called fermentation, is to be described more fully in the future.

The alcohol may be separated by distilling the liquid in which it is formed; its boiling-point being about 85.5° C. (186° F.).

But mere distillation from the fermented liquor, while it may furnish a concentrated spirit, can not give one entirely free from water. The attraction of alcohol for water is so strong that a small portion will be retained by it after the most careful distillation. It can be removed by the stronger attraction of quicklime, and when this is done the product is called *absolute* alcohol; but on exposure to air it soon absorbs water again, so that absolute alcohol is of rare occurrence. The specific gravity of absolute alcohol is .794; of the strongest commercial alcohol, which contains about ten per cent of water, it is .825.

Properties. — Alcohol is a colorless, limpid liquid, lighter than water. It burns readily with a hot, but not a bright, flame. It is able to dissolve resins and oils, and many other substances which water can not; and this solvent power renders it useful in the arts.

Composition. — The composition of alcohol is shown by its formula C_2H_6O. This is its empirical formula. It is often written also C_2H_5HO, which represents the belief that its molecule consists of two groups of atoms, C_2H_5 and HO.

It is interesting to know why the chemist should feel inclined to arrange the atoms in the molecule in these two groups, and a little attention will show the reason.

Constitution of the Molecule. — He finds, that if he mixes ethane, C_2H_6, and chlorine, he may bring about the following reaction: —

$$\begin{array}{ccccccccc} C_2H_6 & + & 2\,Cl & = & C_2H_5Cl & + & HCl, \\ \text{Ethane} & + & \text{Chlorine} & = & \text{Ethyl chloride} & + & \text{Hydrochloric acid;} \end{array}$$

and, if he then treats the ethyl chloride with silver hydrate, he produces alcohol. Thus, —

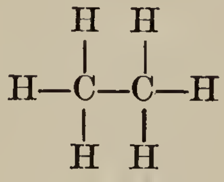
$$\begin{array}{ccccccccc} C_2H_5Cl & + & AgHO & = & C_2H_5HO & + & AgCl, \\ \text{Ethyl chloride} & + & \text{Silver hydrate} & = & \text{Alcohol} & + & \text{Silver chloride.} \end{array}$$

Now, the C_2H_5 seems to have come along down from the ethane unbroken, while the HO, from the silver hydrate, has been added to it; and the union of these two groups, C_2H_5, HO, seems to give the molecule of alcohol.

This is only one of the chemical actions which can be best explained by this view of the arrangement of atoms in the molecule of alcohol. It is very generally accepted.

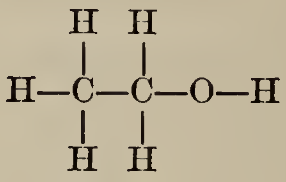
Derived from Ethane. — Compare C_2H_6 with C_2H_5HO, and we see that the latter would be obtained by substituting HO for H in the former.

That is, *alcohol may be derived from ethane by substituting a molecule of hydroxyl for one atom of hydrogen.* This may be clearly shown by the constitutional formula: —



$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & — & C & — & C & — & H \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

Ethane



$$\begin{array}{ccccccccc} & & H & & H & & & & \\ & & | & & | & & & & \\ H & — & C & — & C & — & O & — & H \\ & & | & & | & & & & \\ & & H & & H & & & & \end{array}$$

Alcohol

where it is seen that if we only put the two atoms, OH, in place of the right-hand atom of hydrogen in ethane, we change the molecule of ethane into a molecule of alcohol.

196. The Series of Alcohols. — Now, hydroxyl may be substituted for an atom of hydrogen in other members of the marsh-gas series as well as in ethane.

Methane,	CH_4,	becomes	CH_3OH,
Ethane,	C_2H_6,	“	C_2H_5OH,
Propane,	C_3H_8,	“	C_3H_7OH,
Butane,	C_4H_{10},	“	C_4H_9OH,
Pentane,	C_5H_{12},	“	$C_5H_{11}OH$.

The new compounds are all of them ALCOHOLS. Alcohol is, therefore, the name of a large class of bodies, instead of a single compound. *An alcohol is a substance derived from a member of the marsh-gas series by substituting hydroxyl for hydrogen.*

Their Names. — The first in the series is called *methyl* alcohol, from the hydrocarbon group, CH_3, the name of which is methyl. The second is called *ethyl* alcohol, from the group C_2H_5, or ethyl. The third is *propyl* alcohol, from propyl, C_3H_7. In every case *the name of the alcohol is the name of the hydrocarbon in it;* and the name of the hydrocarbon in it is the name of the corresponding member of the marsh-gas series with the ending *ane* changed to *yl.*

197. Radicals. — These hydrocarbons in the alcohols are called Alcohol Radicals. This term radical is applied to any group of atoms which is not broken up during chemical changes, but which seems to go bodily from one molecule into another.

Let us take the case of methane, CH_4, and represent it as methyl hydride, CH_3H. Now, if we mix chlorine with this, the molecule CH_3H will be changed into CH_3Cl; and then, if we treat this with silver hydrate, $AgOH$, the molecule CH_3Cl will be changed into CH_3OH. We here see that the group CH_3 remains unbroken while the substance changes. This is what the term radical implies.

These alcohol radicals are all unsaturated molecules with one free unit of quantivalence: they are univalent.

SECTION IV.

THE ETHERS.

198. Ether. — Ether is a transparent liquid, with a peculiar odor, and a sweetish taste. When breathed, it causes exhilaration at first, but perfect insensibility at last. On this account ether has been used to render patients insensible to the pain of surgical operations.

The evaporation of ether produces intense cold: a pretty experiment can illustrate this. Let a drop or two of water be covered with a few drops of ether, and by a bellows or a blow-pipe let a current of air be blown against the ether. By its rapid evaporation the ether takes away the heat, until the water is frozen. A mixture of ether with solid carbon dioxide will produce a temperature of −110° C. (−166° F.). Pure ether has never been frozen.

Liquid ether is much lighter than water, but its vapor is much heavier than air.

Ether is very combustible, burning with a bright flame; and a mixture of its vapor with air is explosive.

Ether is used in medicine, and extensively by the chemist as a solvent. Oils and resins, caoutchouc, and many other organic substances, are soluble in this liquid. It is a more powerful solvent than alcohol.

Preparation. — When a mixture of concentrated sulphuric acid and alcohol is heated to about 140° C., a very volatile substance is formed, whose vapors may be condensed in a separate vessel. It is *sulphuric ether;* or, since it is the only ether of commercial importance, it is called simply ETHER. The chemical change is complicated; but the result of it all is, that the alcohol is changed into ether and water.

$$\begin{array}{cccc} 2\,C_2H_5OH & = & (C_2H_5)_2O & + \; H_2O, \\ \text{Alcohol} & = & \text{Ether} & + \; \text{Water.} \end{array}$$

Two molecules of alcohol yield one molecule of ether and one of water.

199. The Series. — Just as common or ethyl alcohol yields common ether when treated with sulphuric acid, so each alcohol will yield a corresponding ether. Thus: —

Methyl alcohol	yields	Methyl ether,
Ethyl alcohol	"	Ethyl ether,
Propyl alcohol	"	Propyl ether.

Hence the term ether, like alcohol, is the name of a large class of compounds. Indeed, this series of simple ethers is only one of several series. Nitric acid, also, converts the alcohols into ethers; and any other strong acid will yield a series of compounds belonging to this same large family. *An ether is a compound produced by the action of a strong acid on an alcohol.*

We may notice that common ether is an oxide of the radical ethyl. This is shown by its formula $(C_2H_5)_2O$. So all this series of ethers obtained by sulphuric acid are oxides in which one atom of oxygen is combined with two molecules of a radical.

SECTION V.

OLEFIANT GAS AND THE OLEFINES.

200. Olefiant Gas. — Olefiant gas is another compound of carbon and hydrogen, whose composition is shown by the formula C_2H_4. It is a colorless and very combustible gas, and burns with a bright flame. It is found in small quantities in illuminating gas, and was formerly thought to be the most important constituent in it; but this idea has been abandoned. Mixed with air it explodes violently.

Olefiant gas is sometimes called Ethylene, and sometimes Ethene.

201. The Olefines. — We have seen that marsh-gas is the simplest or first member of a long series of hydrocarbons, in which there is a constant difference of CH_2. Olefiant gas is like marsh-gas in this respect; for it is also the simplest or first member of a long series, in which there is a constant difference of CH_2. Thus we have

C_2H_4 . . . . Ethene,
$C_2H_4 + CH_2 = C_3H_6$ . . . . Propene,
$C_3H_6 + CH_2 = C_4H_8$ . . . . Butene,
$C_4H_8 + CH_2 = C_5H_{10}$ . . . . Pentene.

The members of this series are produced by distilling organic bodies, and they are found also in petroleum.

202. Homologous Series. — Any series of compounds, the members of which differ by the common difference, CH_2, is called a Homologous Series. The paraffines and the olefines are only two homologous series of hydrocarbons:

many others are known, and all together include a multitude of compounds. The members of each series agree in certain relations and properties, so that a description of one is to some extent a description of them all. Were it not that they may be arranged in these natural groups, any thing like an exhaustive study of the hydrocarbons would be an overwhelming task.

SECTION VI.

DESTRUCTIVE DISTILLATION.

203. Destructive Distillation. — When wood or other vegetable substance is heated in close vessels, it is decomposed. The process is called Destructive Distillation. Some of the products are solid, some are liquid, and some are gaseous: these three classes of substances are always produced by destructive distillation.

204. Destructive Distillation of Wood. — Let the following experiment be tried. Into a small flask (Fig. 88) put some fine splinters of some hard wood: beech-wood answers the purpose well. Let the flask be tightly corked, and provided with a bent tube reaching over into a bottle which is kept cold by the water which surrounds it. On heating the flask the wood soon turns black: volatile matter is driven over, some of it being condensed in the bottle, while another part escapes in the form of gas. This gas may be collected in a second bottle over water, by passing it

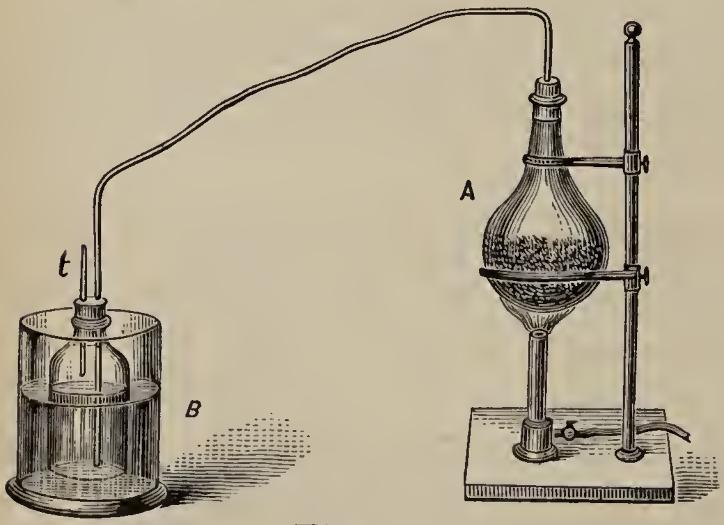


Fig. 88.

through a rubber tube reaching from the short tube *t*. The black solid left in the flask is charcoal: the gas collected is a mixture of several, — marsh-gas and olefiant gas being among them; while the fluid in the first bottle consists of pyroligneous acid and wood-tar.

Charcoal and the gases have already been described.

205. Pyroligneous Acid. — Pyroligneous acid is often called *wood-vinegar*. Dry beech-wood yields it in greatest abundance. It is a dark-brown liquid, with a sour, smoky taste. It contains acetic acid; and on this account has been largely used in making such acetates as are employed in the arts, especially in calico-printing and dyeing. Sodium acetate and lead acetate are examples.

206. Wood-Tar. — Wood-tar is a very dark-colored resinous fluid. There are several varieties. One, largely used in ship-building and other arts, is obtained by a rude distillation of resinous pine-wood: another is obtained from hard wood. It is sometimes used as a covering for wood to preserve it from decay: the more volatile constituents of the tar passing away, leave the harder part (pitch) in the pores of the wood. Water is thus kept out of the pores of the wood; and this, together with the action of creosote, to be soon noticed, prevents decay.

207. Methyl Alcohol. — When pyroligneous acid is distilled, a very volatile liquid may be obtained, which, being afterward rectified by the use of quicklime, constitutes the METHYL ALCOHOL of commerce. It is a limpid liquid, very inflammable. It may be used instead of alcohol for many purposes, especially for dissolving resins in making varnish. It is the simplest member of the alcohol series.

208. Creosote. — The smoky taste of pyroligneous acid, and the power of both this acid and wood-tar to prevent decay, is due to the presence of a curious compound of carbon, hydrogen, and oxygen ($C_8H_{10}O$), called CREOSOTE.

One pound of the acid contains about a quarter of an ounce of it in solution. Its most curious and valuable property is its power to prevent decay: indeed, it is among the most powerful antiseptics known. Flesh remaining a few hours in a fluid made by dissolving one part of creosote in one hundred parts of water will not afterward decay. That meats are often *cured* by exposure to smoke, is familiar to all: now, the preservation and the peculiar flavor of smoked meat is due to the action of creosote in the smoke. This substance is often used in medicine; but when taken internally, except in very small quantities, it is a corrosive poison.

209. Paraffine. — Paraffine may be obtained by distilling wood-tar. It is a crystalline solid, with neither taste, color, nor smell. Its most remarkable property is its indifference to the chemical action of other substances. It can resist the action of the strongest alkalies and of the most corrosive acids. It is, however, combustible, and its flame is white and smokeless. Paraffine candles rival the most costly wax candles in luster and in the strength and beauty of their light.

210. Other Substances. — Numerous substances, besides the few just described, may be obtained by the destructive distillation of wood. Each different kind of wood and of other vegetable matter yields some different products; but for the most part they are all compounds of carbon with either hydrogen or oxygen, often with both.

It should be noticed that none of these products are supposed to exist ready formed in the plant. By heat the substance of the plant is broken up; its elements are rearranged, and these hydrocarbons formed thereby.

Application of this Process. — Destructive distillation is used on a large scale in the manufacture of illuminating-gas. In this manufacture the solid and liquid products are of secondary importance, while the gaseous products are secured as pure as possible.

Any complete apparatus for illustrating the process of destructive distillation illustrates also the essential parts of the apparatus of the gas-house. For example, in Fig. 89, there is a *flask*, in which the organic matter is heated ; a *bottle* surrounded by cold water, in which impurities are condensed ; and a *receiver* over water, in which the gas is collected. These three parts of the apparatus are represented in every "gas-works."

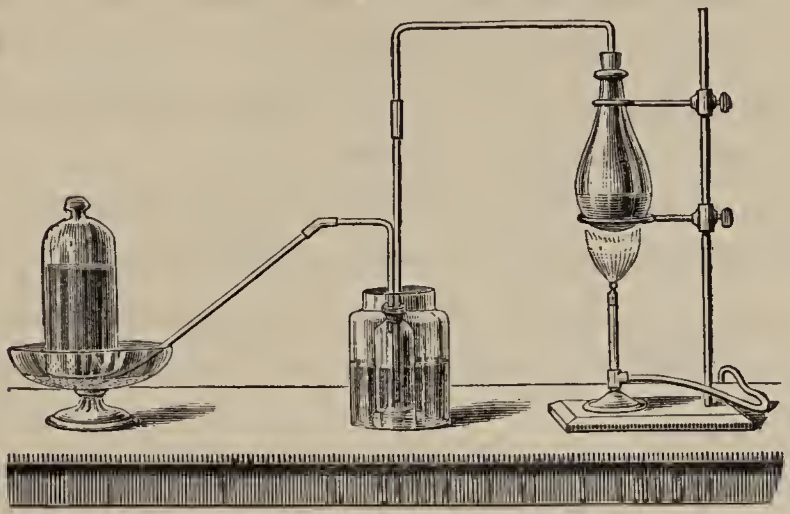


Fig. 89.

211. Manufacture of Illuminating Gas. — By the destructive distillation of bituminous coal, and sometimes of other substances, illuminating-gas is made. The material being heated in iron retorts, its volatile constituents are driven off. The gas thus formed is purified by passing first through cold pipes, and then over lime. It is afterward collected in gas-holders, from which it is pressed out into the pipes of the city, and through these into the chandeliers of the houses.

The following is an outline of this important manufacture.

212. Bituminous Coal. — In the United States alone, there are about 130,000 square miles of workable coal-fields. The coal found so abundantly in nature is called MINERAL COAL : it consists of carbon mixed with other matter, especially with volatile compounds of carbon and hydrogen. The two varieties of mineral coal, *anthracite* and *bituminous*, differ chiefly in the amount of their volatile constituents. The first contains very little of the hydrocarbons, is very hard, and burns with a very feeble bluish flame ; the second contains a large amount of volatile compounds, is much softer, and burns with a bright flame. From bituminous coal illuminating-gas is generally made.

Heated in Iron Retorts. — The vessels in which the coal is heated are called retorts. They are usually of iron, about seven feet long, and scarcely more than a foot in diameter. Fig. 90 shows the ends of five of these retorts placed in a single furnace. Each one, after receiving a charge of from one hundred pounds to one hundred and fifty pounds, is closed air-tight as shown at G, and made red-hot by the fire in the furnace, whose doors are shown at A. In the course of a few hours the volatile matter of the coal is driven off : the residue, called coke, is then raked out, cooled, and used for fuel.

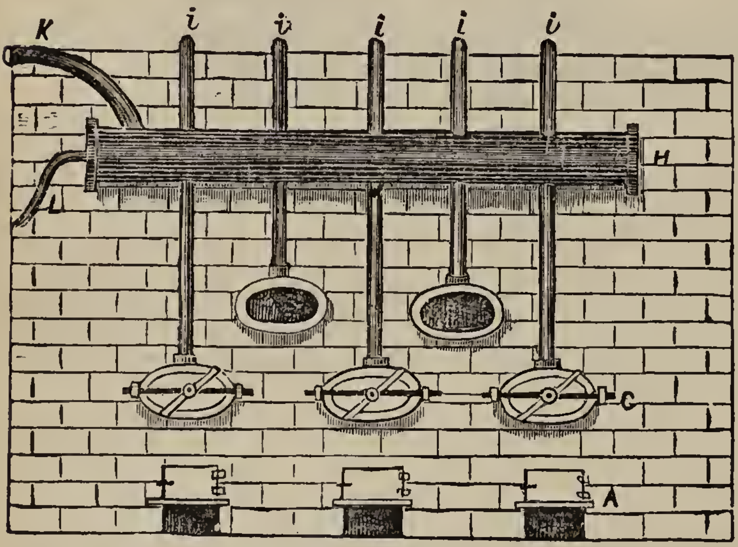


Fig. 90.

Its Volatile Constituents. — The gaseous mixture driven off by heat contains marsh-gas and many other hydrocarbons, carbon dioxide, hydrogen, ammonia, hydrosulphuric acid, and coal-tar, besides other substances. This mixture is totally unfit for use, and must be purified.

Pass through the Hydraulic Main. — From each retort a vertical pipe (*i i i*, Fig. 90) ; is an outlet for this mixture of gases. This pipe, bending over at the top, reaches down into a larger and horizontal tube or trunk, H, called the HYDRAULIC MAIN. In the beginning of this process this main is filled half full of water, and the pipes *i i* dip into this fluid. The gas coming over from the retorts bubbles up through the water, which prevents its return. Now the vapors of coal-tar will be condensed, in part, by the lower temperature of the main; but, as the fluid increases, it runs off through the tube L to a tar-cistern. Much of the coal-tar is left in the hydraulic main, while the gas passes out of it through the pipe K.

Through the Condensers. — This pipe, K, leads the gases over to a series of upright pipes, C C (Fig. 91), called the condensers. Passing up one, and down another, until they have traversed the whole series, the gases are exposed to a large extent of cold surface, and the condensable gases are changed to a liquid state. The tar and ammoniacal liquor thus condensed run into a cistern below, from which they may be drawn off at pleasure.

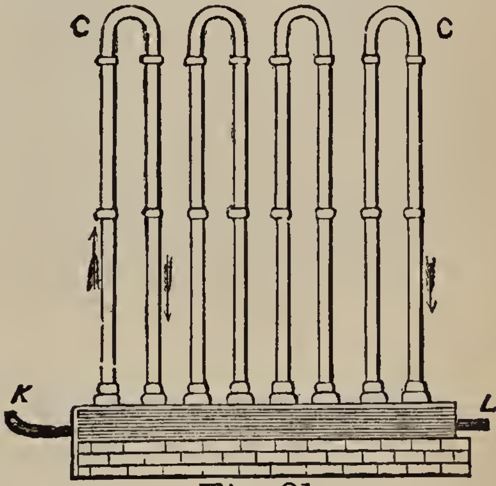


Fig. 91.

Through the Lime Purifier. — The gas, still containing sulphur compounds and carbonic acid, passes from the condensers through a pipe (L) into a chamber (Fig. 92) in which are several sieve-like shelves covered with *slaked lime*. In passing through the lime the gas loses its carbon dioxide and hydrosulphuric acid.

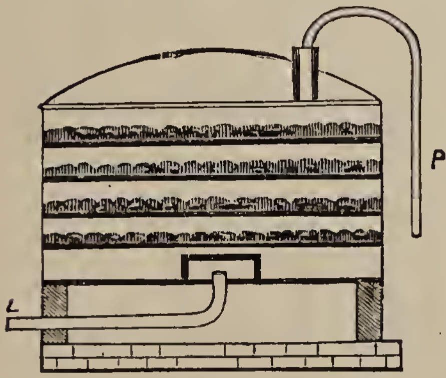


Fig. 92.

Into the Gasometer. — The purified gas, leaving the lime chamber through a pipe (P), passes into the *gas-holder* (Fig. 93), an immense sheet-iron cylinder, closed at the top and opened at the bottom, hung by chains which run over pulleys at the top, and which carry weights to balance it. Below it is a well of water large enough and deep enough to let this cylinder down until it is filled with water. As the gas enters it the gas-holder rises; and, when it is filled, the gas is ready to be pushed out through the pipe S into the streets, and finally into the houses of the city, furnishing to all a convenient and beautiful light.

"In the iron arteries under towns, in the constellations of burners that rule the nights of favored days, rising over the chaotic oil-lamps of old, what a creation!"

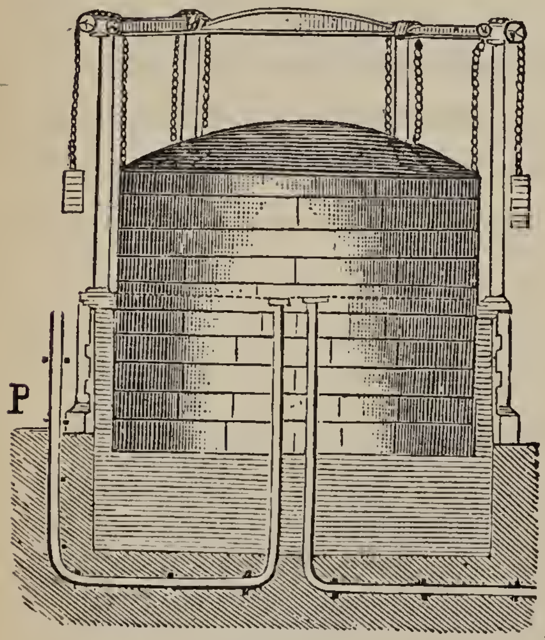


Fig. 93.

213. Gas from Petroleum. — A very elegant and successful process of making gas from petroleum has in recent years been successfully carried out. In this process petroleum vapor and water-gas are brought together in a white-hot retort, from which they issue as a mixture of permanent gases of good illuminating power. The following outline will show the principles of the process.

Water-Gas. — Steam from a boiler is carried into a furnace, and heated to a high temperature. This "superheated steam" is driven into a retort containing some red-hot anthracite coal. Here a chemical reaction occurs.

$$\begin{array}{ccccccccc} C & + & H_2O & = & CO & + & H \\ \text{Carbon} & + & \text{Steam} & = & \text{Carbon monoxide} & + & \text{Hydrogen.} \end{array}$$

This mixture of carbon monoxide and hydrogen is called Water-Gas.

Decomposition of the Petroleum. — The water-gas passes over into another retort kept intensely hot, into which petroleum is introduced. This liquid is at once vaporized and decomposed. Its heavy hydrocarbons C_9H_{20}, C_8H_{18}, etc., are broken into light hydrocarbons, such as marsh-gas, olefiant-gas, and other permanent gases.

The Illuminating Gas. — These hydrocarbons, mixed with carbon monoxide and hydrogen, constitute the illuminating gas. Several of the hydrocarbons are rich in carbon, and would burn with a smoky flame if they were alone; but they are diluted with the marsh-gas, carbon oxide, and hydrogen, until they are able to burn with a clean and brilliant light. Because the marsh-gas, carbon oxide, and hydrogen are useful to dilute the heavy hydrocarbons, these constituents are called the DILUENTS, while the light-giving gases are called the LUMINANTS.

214. Coal-Tar. — The coal-tar of the gas-works is a very complex substance. When distilled, vapors containing ammonia first pass over, and then a light oil, known as *coal naphtha*, followed by a heavier one called *dead oil*, containing a small portion of paraffine. A black *pitch* or *asphalt* is left.

But these are only proximate constituents of the tar: each is itself made up of many simpler ones. The naphtha, for example, is made up of several distinct kinds of oil which may be separated by careful distillation, each having its own peculiar boiling-point.

215. Carbolic Acid. — One of the most important substances obtained from coal-tar is carbolic acid or phenol (C_6H_6O). It is not a direct product of distillation; but it is obtained from the naphtha which comes over between 300° and 400° F., by the action of sodium hydrate (caustic soda). When pure, it is a white solid, soluble in alkalies, with a smell like creosote. Its most important property is its power to interrupt decay: it is a good disinfectant, and is in much demand for this purpose. It is used to some extent in medicine, and quite largely in the manufacture of colors for dyeing silk and woolen goods.

216. Benzol. — Benzol (C_6H_6) is another important constituent of coal-tar. It is a colorless liquid, very volatile and very combustible. It is a powerful solvent of oils and other fatty substances, and may be used to remove grease spots from silk or woolen fabrics.

217. Nitro-Benzol. — By mixing benzol with strong nitric acid a reaction is brought about, shown in the equation,

$$C_6H_6 + HNO_3 = C_6H_5NO_2 + H_2O.$$

It will be noticed that one atom of hydrogen in the benzol is replaced by one molecule of the radical NO_2, forming $C_6H_5NO_2$. This new substance is called NITRO-BENZÓL.

Nitro-benzol is a fluid having the odor of bitter almonds. It is used to some extent in making perfumes, but its most important use is in the production of aniline.

218. Aniline. — Nitro-benzol may be changed to aniline in different ways. Hofmann's method consists in acting upon it by hydrogen set free from sulphuric acid by zinc. The following reaction takes place: —

$$C_6H_5NO_2 + 6H = C_6H_7N + 2H_2O.$$

Losing two combining weights of oxygen, and gaining two of hydrogen, the nitro-benzol is changed to C_6H_7N. This new substance is aniline.

Manufacture. — On a large scale aniline is made by a more economical method (Bechamps'), in which the change is effected by iron and acetic acid. One hundred parts of the crude nitro-benzol is mixed with nearly its own weight of strong acetic acid; and to this is added, little by little, about one hundred and fifty parts of iron-turnings. A complicated reaction takes place; and when the mixture is afterward heated, impure aniline is obtained. It is purified by treating it with lime or soda, and re-distilling it. By this means the crude aniline of commerce is obtained.

Properties. — Aniline, when pure, is a colorless liquid, heavier than water, soluble in alcohol and ether, and very slightly in water. Its most remarkable property is that of acquiring various and rich colors by the action of different oxidizing agents.

Aniline Red. — If, for example, aniline is heated with arsenic acid to about 150° C., a part of its hydrogen will be extracted as water, and there will remain what is called ROSANILINE, in combination with the arsenic. Sodium hydrate may then be added: it will remove the arsenic, and precipitate the rosaniline.

Let this rosaniline be treated with hydrochloric acid, and a rosaniline chloride will be formed. This salt is the so-called ANILINE RED. It occurs in crystals, having a green luster, but its solution in water or in alcohol is a beautiful red.

Aniline Dyes. — By the action of other chemicals on aniline or its salts, many coloring matters are obtained. Various rich shades of red, yellow, green, blue, and black, indeed, almost every variety of tint, are made, and largely used in the arts.

219. Slow Destructive Distillation. — The decay of vegetable matter, when buried in the moist earth, or covered by the water and mud of bogs and marshes, is somewhat different from its decay when exposed to the air. Instead of giving off carbon dioxide and water, and crumbling to a black mold, it gives off marsh-gas and other hydrocarbons, and yields a residue of coal.

The chief constituents of wood are carbon, hydrogen, and oxygen; the first two being far the most abundant. On exposure to air, oxygen from the atmosphere combines with hydrogen and carbon of the wood to form water and carbon dioxide, leaving oxygen in combination with what remains of both these elements, forming humus; but when the air is excluded, the process of decay must consist chiefly in the re-arranging of the elements of the wood itself. Its oxygen takes carbon enough to form carbon dioxide; its hydrogen takes carbon also, and forms gaseous or liquid compounds;

while the excess of carbon, not thus used, is left in the form of coal.

Varieties of Coal. — Vast quantities of vegetable matter, accumulating in low wet lands of warm countries, gradually become covered with water; and sometimes, by the sinking of the land, they are buried under mud and sand brought over them by streams or floods. Thus shut off from the air, a slow decay goes on, by which they are at last changed to coal. The different varieties of coal mark the different stages of the process. In peat the change is only well begun: in anthracite coal the process is at an end. The warmth of the earth assists the change; and the great pressure of the material, accumulating for ages upon it, must have had much to do with the final compactness of the remaining coal. In bituminous coal the liquid hydrocarbons remain, but may be driven away by heat: from anthracite they have already escaped.

Origin of Petroleum. — Numerous and extensive beds of coal have thus been produced by the slow distillation of vegetable matter during past ages of the earth's history.

But what has become of the liquid hydrocarbons which must have been formed, but which are no longer held in the hard coal? Moreover, during the deposition of other rocks in which no coal is found, there is abundant evidence that plants were growing, and they, too, must have been decomposed in a similar way: what has become of the products of their decay?

The gaseous products would, of course, for the most part, escape into the air; and it would be natural to suppose that the liquid products would gradually collect in cavities and fissures in the rocks. Now, inflammable, oily substances, issuing often in large quantities from the fissures of rocks, have been long known. To them the general name of *petroleum* has been given. They resemble the liquid products obtained by destructive distillation of wood; and it is believed that they are the products of the slow decomposition of organic matter, chiefly vegetable. Petroleum, or rock-oil, is, then, the liquid hydrocarbon substances given off in the slow process of the decay of vegetable matter long buried in the earth.

SECTION VII.

THE SUGARS.

220. Sugar. — The term "sugar," when used in chemistry, is not the name of any single substance, but the name of a class. The sugars are, all of them, compounds of carbon, hydrogen, and oxygen; and in their composition there is this peculiarity, viz., *The hydrogen and oxygen are in the proportions to form water.* There are many varieties of sugars; but they may be grouped in three classes, — the *sucroses*, the *glucoses*, and the *amyloses*.

The Sucroses. — Cane-sugar, so common and well known, and milk-sugar, or lactose, obtained by evaporating the whey of fresh milk, are members of the first class. Cane-sugar occurs in the juices of many plants. It is obtained by evaporating the sap of the sugar-maple, or the juice of the beet, and in far larger quantities from the juice of the sugar-cane. Its general character is well known. Its composition is represented by the formula $C_{12}H_{22}O_{11}$. When strongly heated, it yields water and a dark-colored residue called caramel. One of its most curious properties is shown in its action upon polarized light: it turns the plane of polarization to the right hand.

The Glucoses. — Grape-sugar and fruit-sugar are glucoses. They are found together in many kinds of fruit, especially in the grape. They have the same composition, $C_6H_{12}O_6$, and yet differ in several properties. Grape-sugar easily crystallizes: fruit-sugar never does. The latter is more soluble, and rotates the plane of polarization *to the left*, the former to the right.

Sucrose Changed to Glucose. — When cane-sugar is acted on by dilute sulphuric acid, a reaction takes place by which the cane-sugar is changed into grape-sugar and fruit-sugar, by taking the elements of a molecule of water: —

$$\begin{array}{ccccccccc} C_{12}H_{22}O_{11} & + & H_2O & = & C_6H_{12}O_6 & + & C_6H_{12}O_6, \\ \text{Cane-sugar} & + & \text{Water} & = & \text{Grape-sugar} & + & \text{Fruit-sugar.} \end{array}$$

The Amyloses. — Starch is the most familar example of the amyloses. It consists of a white powder, composed of granules, which have different size and shape in different varieties: those of potato-starch being about .007 inches in diameter; of beet-root, about .0002 inches. These granules are not soluble in cold water, but when heated in water they swell and split open; and, if the paste thus formed is boiled in a larger quantity, the starch is at last dissolved.

When free iodine is brought in contact with starch, a compound, having a rich blue color, is made. This action is the most delicate test for the presence of starch. To show its presence in a potato, for example, let the freshly cut surface of the vegetable be washed with a solution of iodine.

The composition of starch is shown by its formula, $C_6H_{10}O_5$. By the action of dilute sulphuric acid, starch is changed into dextrine and grape-sugar.

Dextrine. — Dextrine, or, as it is more commonly called, British gum, is another amylose. It is very soluble in water, and it is used sometimes instead of gum-arabic in calico-printing and other arts. It is made from starch, not only by means of dilute sulphuric acid, but by simply heating the starch to about 150° C. (302° F.), or by the action of *diastase* (a substance contained in malt). By the continued action of the diastase the starch is first changed into dextrine and grape-sugar, and the dextrine is finally changed also into grape-sugar.

$$\begin{array}{ccccccc} 3(C_6H_{10}O_5) & + & H_2O & = & 2(C_6H_{10}O_5) & + & C_6H_{12}O_6, \\ \text{Starch} & + & \text{Water} & = & \text{Dextrine} & + & \text{Grape-sugar.} \end{array}$$

Dextrine and starch have the same composition, $C_6H_{10}O_5$, but their properties are different. Dextrine is soluble in water, and is reddened, instead of being turned blue, by iodine. Bodies having the same composition, but different properties, are said to be *isomeric*. Dextrine and starch are isomeric substances.

221. Ferment. — By the term *ferment* we mean an organic compound containing nitrogen, and which readily decomposes on exposure to air. Any substance containing nitrogen, and partially decomposed, will act as a ferment. Yeast is the most familar example. When the sweet juices of vegetables are exposed to the air, a ferment is soon formed in them; and, the smallest quantity of ferment being present, an action is started by it which goes on until the entire body of liquid is decomposed.

Fermentation. — The decomposition caused by ferments is called FERMENTATION. It may be easily illustrated by experiment. Dissolve about one hundred grains of honey, or it may be molasses, in a pint of water; fill a small flask with the solution, and add a few drops of brewer's yeast. Close the neck of the flask with the hand, and invert it in a dish holding some of the same sirup, and leave it in a warm place for twenty-four hours. Fermentation soon begins; a colorless gas collects in the flask, which, by lime-water, may be shown to be carbon dioxide, while alcohol remains in the fluid. Thus: —

$$\begin{array}{ccccc} C_6H_{12}O_6 & = & 2\,C_2H_6O & + & 2\,CO_2, \\ \text{Sugar} & = & \text{Alcohol} & + & \text{Carbon dioxide.} \end{array}$$

All fermentation which produces chiefly alcohol and carbon dioxide is called the ALCOHOLIC or VINOUS fermentation. The process goes on best at a temperature of 25° or 30° C.

Spirituous Liquors. — The spirituous liquors of commerce, such as brandy, gin, and whiskey, are produced by distilling fermented liquids. The fermented liquid obtained from malted grain is called BEER; that from the juice of the grape is called WINE. By distilling these, and adding various substances to color and flavor the result, different kinds of liquor are made. Brandy is made by distilling wine; gin is made from different kinds of corn spirits, its flavor being given by juniper-berries, sweet-flag, licorice-powder, and several other substances. Whisky is also obtained by distilling the fermented liquor from corn.

The intoxicating principle in all these liquors is alcohol, which has been produced by fermentation.

222. The Acetous Fermentation. — An alcoholic liquid, which contains a small quantity of a ferment, and is in the presence of air, yields acetic acid. Acetic acid is the acid which gives sourness to vinegar.

Fermentation in which acetic acid is the chief product is called the ACETOUS FERMENTATION.

Production of Vinegar. — When an alcoholic liquid is exposed to the air in a warm place, a little yeast or other nitrogenous matter in it will start an action by which the alcohol is changed into vinegar. "A good extemporaneous vinegar may be prepared by dissolving one part of sugar in six of water, with one part of brandy, and a little yeast. The mixture is put into a cask, with the bung-hole open, and kept at a temperature between 70° and 80° F. In from four to six weeks the clear vinegar may be drawn off." (*Brande & Taylor.*) A still more simple process consists in soaking crushed apple-skins in soft water for a few days, straining the juice, and letting it stand exposed to the air in a warm place for several days: an excellent vinegar is the result.

Acetic Acid. — Common vinegar is a very dilute acetic acid. Its quality depends upon the proportion of acid it contains, and the absence of other impurities. The composition of acetic acid is shown by the symbol $C_2 H_4 O_2$. It is a colorless liquid, with a powerful and peculiar odor, which, once experienced, is afterward easily recognized.

Fermentation of Alcohol. — The chemical action by which alcohol is changed to acetic acid is called the *acetous fermentation.* And yet it is not in all respects a true fermentation. It is not a *decomposition*, but rather an *oxidation*, as may be seen by comparing the symbols of alcohol and acetic acid. In this respect the action is a case of combustion rather than of fermentation. It can take place only in the presence of air, so that the action is not entirely due to a ferment. Yet, on the other hand, it will not occur, except in presence of a nitrogenous substance, to which the term *ferment* has been given ; and hence the reaction is very naturally called a fermentation.

REVIEW.

I. — SUMMARY OF PRINCIPLES.

223. Organic chemistry is the chemistry of the compounds of carbon.

The compounds of carbon are the products of the decomposition of organic bodies : with some exceptions, we have no evidence that they exist ready formed in the living plant or animal.

A large number of these substances contain only carbon and hydrogen ; others contain oxygen in addition ; and a few contain other elements, especially nitrogen.

Marsh-gas is the simplest hydrocarbon. Its molecule is represented by $C\,H_4$.

It is found in nature ready formed. In many places it issues in large quantities from the earth : it collects in mines, and it also issues from the mud of stagnant pools.

A homologous series is a series of compounds containing the same elements, in which there is a common difference in the molecules of successive members.

Marsh-gas is the first member of a homologous series of hydrocarbons, in which the common difference is $C\,H_2$.

The densities of these hydrocarbons increase regularly as the molecules become more complex.

Carbon is quadrivalent, and its atoms are able to combine with one another: these two assumptions furnish an explanation of the paraffine series.

Petroleum contains several members of this series from C_4H_{10} to C_9H_{20}, inclusive, beside several hydrocarbons belonging to other series.

The boiling-points of these hydrocarbons differ: hence they may be separated by fractional distillation. Several commercial products are obtained by the fractional distillation of petroleum.

Common alcohol is obtained by distilling a fermented liquid. It is a colorless liquid, lighter than water, very volatile and combustible.

The molecule of alcohol is C_2H_6O. Its rational formula is C_2H_5HO. According to this formula its chemical name should be ethyl hydrate.

It may be regarded as derived from ethane, C_2H_6, by the substitution of HO for one atom of H.

Common alcohol is one of a large number of analogous substances which constitute a series of alcohols.

An alcohol is a compound derived from a hydrocarbon by substituting a molecule of hydroxyl for an atom of hydrogen.

A radical is a group of atoms which passes bodily from one combination to another in chemical actions. The hydrocarbon combined with hydroxyl in an alcohol is called an alcohol radical.

Ether is a limpid liquid, very volatile, very combustible, a powerful solvent, and an anæsthetic.

It is obtained by treating ethyl or common alcohol, with sulphuric acid. Its composition is represented by $C_4H_{10}O$, or by the rational formula $(C_2H_5)_2O$. According to this formula its chemical name should be di-ethyl oxide.

Common ether is one member of a large class of substances called the ethers. An ether is a substance produced by the action of a strong acid on an alcohol.

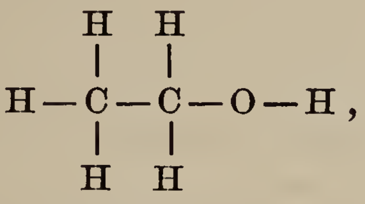
The composition of methyl ether is represented by C_2H_6O, or by its rational formula $(CH_3)_2O$.

But the composition of common alcohol is also represented by the same formula, C_2H_6O. Hence these two substances furnish an example of *isomerism*, — having the same chemical composition, but different properties.

Isomerism is explained by supposing the atoms of the substances to be differently arranged in the molecules.

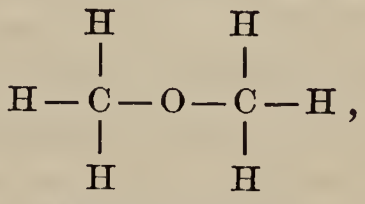
The rational formula may show the different groupings of the atoms which we suppose to exist. Thus we write for alcohol C_2H_5OH, and for methyl ether $(CH_3)_2O$.

The graphic form of the rational formulas shows this difference still more clearly. Thus: —



$$\begin{array}{ccccccccc} & & H & & H & & & & \\ & & | & & | & & & & \\ H & - & C & - & C & - & O & - & H, \\ & & | & & | & & & & \\ & & H & & H & & & & \end{array}$$

Alcohol.



$$\begin{array}{ccccccccc} & & H & & & & H & & \\ & & | & & & & | & & \\ H & - & C & - & O & - & C & - & H, \\ & & | & & & & | & & \\ & & H & & & & H & & \end{array}$$

Methyl ether.

In the first we discover two groups, C_2H_5 and OH. In the second we discover two groups, CH_3 and CH_3, bound together by oxygen.

Isomeric substances are very numerous among the compounds of carbon.

Olefiant gas is a colorless gas, very combustible and explosive. It is represented by C_2H_4.

This gas is the first member of a homologous series, — the olefines.

Destructive distillation is the decomposition of an organic substance by heat in the absence of air. The products are solid, liquid, and gaseous.

Illuminating gas is made by the destructive distillation of soft coal, sometimes of other materials.

The solid residue is coke. The liquid products are collectively called coal-tar and ammoniacal water. The gaseous product contains the luminants, the diluents, and the impurities.

Carbolic acid and benzol are among the valuable products obtained directly from coal-tar. From benzol comes nitro-benzol, and finally aniline. From aniline many rich and valuable dyes are made.

Organic matter buried in the earth undergoes a slow process of destructive distillation. The varieties of mineral coal, and petroleum or rock-oil, are doubtless the products of such a process.

Sugar is a compound of carbon, hydrogen, and oxygen, in which the last two are in the proportions to form water.

There are three classes, — the sucroses, the glucoses, and the amyloses.

The sucroses may be changed to glucose by dilute sulphuric acid. The same is true of the amyloses.

Fermentation is a decomposition caused by ferments. The vinous fermentation consists in the change of glucose into carbon dioxide and alcohol.

II. — EXERCISES.

Define organic chemistry. What is an organized body? What are organic substances? With which of these does chemistry deal?

Define hydrocarbon. Name the simplest hydrocarbon, and give its formula. Where is marsh-gas found in nature? What are its properties? How may we prove that its molecule contains four atoms of hydrogen?

By what other names is marsh-gas known? What does the name methyl hydride signify? How would its formula be written?

Give the formulas and names of the first four members of the marsh-gas series. How do chemists explain the fact that hydrogen and carbon can form so many compounds? Write the graphic formulas for the first four members of the series.

What are the properties of this series?

What is petroleum? What is its composition? Define fractional distillation. What is kerosene? Benzine? Naphtha?

How may alcohol be obtained? What is absolute alcohol? How may it be obtained? What are the properties of alcohol?

What is the composition of alcohol? If we call it ethyl hydrate, how should we write its formula? What evidence is given that the molecule does contain these two groups?

By what substitution may alcohol be derived from ethane?

Define the term alcohol. How are the alcohols named?

Define the term radical. Show that $C H_3$ remains unbroken while methane is passing through the changes to become methyl alcohol.

What is ether? What is its effect when breathed? What is the effect of its evaporation? What are some of its uses?

How is ether prepared? Give the chemical change.

To what class of substances is the name ether given? Which of these is the common ether? What is its formula?

What is olefiant gas? What are some of its properties? What are its other names?

What are the olefines? From what sources may these bodies be derived?

Define homologous series.

What is destructive distillation?

Describe the experiment showing the destructive distillation of wood.

What is said of pyroligneous acid?

What is said of wood-tar?

What is said of methyl alcohol?

What is said of creosote? What is its most curious and valuable property?

How may paraffine be obtained? What are some of its properties?

What is said of other products of destructive distillation? Do these substances exist in the plant?

Give a brief description of the manufacture of illuminating gas?

What is said of mineral coal? What are two principal varieties? Which is used in making the gas?

Describe the heating in iron retorts.

What substances are driven off?

What becomes of this mixture? What occurs in the hydraulic main?

Describe the next step in the process.

What impurities still remain? How are they removed?

Into what is the purified gas finally conducted? Describe the gasometer, and explain its action.

From what other substance is illuminating gas made? What is water-gas? How is it made? How are water-gas and petroleum converted into illuminating gas?

Of what does the illuminating gas consist?

What is the formula of carbolic acid? How is this acid obtained? What are some of its properties? Its uses?

What is the formula of benzol? What are some of its properties?

Give the reaction by which nitro-benzōl is formed? What are its uses?

Show how nitro-benzol may be changed to aniline. Describe the method adopted in the arts. What are the properties of aniline?

How is aniline changed into aniline red? How are other aniline dyes made?

How have mineral coal and petroleum been formed?

Describe the decay of vegetable matter when kept from the air. Show wherein this decay differs from decay in air.

Explain the formation of the varieties of coal.

In the formation of mineral coal, would these gases and volatile liquids be driven off? What would naturally become of them? Is any thing of the kind now to be found in the rocks? What name is given to them? Describe them.

Define the term sugar. Name the three classes.

Name varieties of sucrose. How is cane-sugar obtained? What is its composition? What are its properties?

Name varieties of glucose. What is their composition? Wherein do they differ? Give the reaction by which cane-sugar is changed into grape and fruit sugar.

What are the properties of starch? The test for its presence? Its composition?

What is dextrine? How is it made from starch? Give the symbols, and explain the reaction.

What is a ferment? Name a familiar example.

What is fermentation? Describe the experiment, and explain the reaction. What is the alcoholic fermentation?

How are spirituous liquors obtained? Name some common kinds, and tell how they are made.

What is the intoxicating principle in all kinds?

How may alcohol be changed to acetic acid?

Give the formula for acetic acid.

Describe the acid.

What is common vinegar composed of? What is the acetous fermentation? Is it in all respects a true fermentation?

CHAPTER IV.

THE METALS.

SECTION I.

GENERAL DESCRIPTION.

224. Characteristic Properties. — The peculiarities of metallic elements are more or less familiar. Their *luster*, as of silver; their *malleability*, as of gold and zinc; their *ductility*, as of iron and copper; together with their power to conduct heat and electricity, — are their most characteristic properties. One or more of these properties are possessed by some non-metals; and, on the other hand, some metals have them only in a slight degree. Indeed, nature seems to have drawn no precise line of division between the two classes. Certain elements, arsenic and antimony for example, have been classed with metals, but are considered now to be non-metals; while even hydrogen, because of its chemical relations, is thought by some to be a metal.

Melting-Points. — The metals, with the exception of mercury, are solids at ordinary temperature. Some are easily melted, as potassium, at 62° 5 C. (144° 5 F.); others melt with difficulty, as iron at 1,600° C. (2,912° F.); while others, like platinum, melt only in the intense heat of the oxyhydrogen blow-pipe.

Density. — Osmium is the heaviest of metals, 22.47 times heavier than water; others, as potassium and sodium, are so light that they will float on the surface of water. Lithium is the lightest of all: its specific gravity is only .594.

Condition in Nature. — A few metals, copper and gold for example, are found in nature in the metallic state: this condition is commonly called Native. But in the native state they are seldom pure: two or more are combined. Combinations of metals are called Alloys. The metals are, however, usually found combined with non-metals; and such compounds are called Ores.

225. Classification of the Metals. — In the following table the metals are classed so as to bring together those which most closely resemble each other. Many of the metals are rare, and not important to the general student: others are of the greatest use and interest. Of these last the names are printed in capitals, and to the description of them we are to pay the more particular attention.

1. — *Metals of the Alkalies.*

POTASSIUM . . .	K.$^{\mathrm{I}}$
SODIUM	Na.$^{\mathrm{I}}$
Lithium	Li.$^{\mathrm{I}}$
Cæsium	Cs.$^{\mathrm{I}}$
Rubidium . . .	Rb.$^{\mathrm{I}}$

2. — *Metals of the Alkaline Earths.*

CALCIUM	Ca.$^{\mathrm{II}}$
Strontium . . .	Sr.$^{\mathrm{II}}$
Barium	Ba.$^{\mathrm{II}}$

3. — *The Aluminium Class.*

ALUMINIUM . . .	Al.$^{\mathrm{II,\,IV}}$
Indium	In.
Gallium	Ga.

4. — *Cerium Class.*

Cerium	Ce.
Ytrium	Y.
Erbium	Er.
Lanthanum . . .	La.
Didymium . . .	Di.

5. — *Zinc Class.*

ZINC	Zn.$^{\mathrm{II}}$
Magnesium . . .	Mg.$^{\mathrm{II}}$
Cadmium . . .	Cd.$^{\mathrm{II}}$
Beryllium . . .	Be.

6. — *Iron Class.*

IRON	Fe.$^{\mathrm{II,\,IV}}$
Manganese . . .	Mn.$^{\mathrm{II,\,IV}}$
Nickel	Ni.$^{\mathrm{II,\,IV}}$
Cobalt	Co.$^{\mathrm{II,\,IV}}$

7. — *Tin Class.*

TIN	Sn.$^{\mathrm{II,\,IV}}$
Titanum	Ti.$^{\mathrm{II,\,IV}}$
Zirconium . . .	Zr.
Thorium	Th.

8. — *Chromium Class.*

Chromium . . .	Cr.$^{\mathrm{II,\,IV}}$
Molybdenum . .	Mo.
Tungsten . . .	W.
Uranium	U.

9. — *Antimony Class.*

ANTIMONY	$Sb.^{III}$
ARSENIC	$As.^{III}$
BISMUTH	$Bi.^{III}$
Vanadium	V.
Niobium	Nb.
Tantalum	Ta.

10. — *Lead Class.*

Lead	$Pb.^{II}$
Thallium	Th.

11. — *Silver Class.*

Copper	$Cu.^{II}$
Mercury	$Hg.^{II}$
Silver	$Ag.^{I}$

12. — *Gold Class.*

GOLD	$Au.^{III}$
PLATINUM	Pt.
Palladium	Pl.
Rhodium	Rh.
Ruthenium	Ru.
Iridium	Ir.
Osmium	Os.

SECTION II.

METALS OF THE ALKALIES.

226. General Description. — The metals of the first class are very soft and light, having so violent an attraction for oxygen that they can decompose water at any temperature. They are univalent.

Of this class only potassium and sodium are of interest to the student; and even these metals are of little use, and are important only because their compounds are of great value in the arts.

Illustration of these Class Properties. — A piece of potassium or of sodium may be molded in the fingers like wax, and if dropped upon water it floats.

Upon a piece of ice, in a small cavity made for the purpose, place a small fragment of potassium: a purple flame springs up, as if the ice had been set on fire. A smart explosion usually ends the experiment; and if we then examine the water that is left in the cavity of the ice, we find it to contain potassium hydrate (potash). So strong is the attraction of this metal for oxygen, that it decomposes water, even at the temperature of ice. This is true of the other members of the group.

The Reaction. — If we study the reaction taking place in the experiment, we shall see that the metal is univalent. Thus: —

$$H_2O + K = KHO + H.$$

One combining weight of potassium has simply replaced one of the two combining weights of hydrogen in the molecule of water. A similar reaction would occur with the other members of the group. The potassium hydrate formed is an ALKALI; the other hydrates of the group are also alkalies. The alkalies differ from most other metallic hydrates in their power to withstand heat: heat alone will not decompose them.

227. Manufacture of Potassium Carbonate. — The ashes of wood, mixed with about five per cent of lime, are placed in tubs, and drenched with successive portions of fresh water. As the water soaks through the ashes, it dissolves out the soluble constituents, among which is the potassium carbonate. This process is called leaching. The solution known as LYE is put into broad and shallow pans, and evaporated. The solid residue is called CRUDE POTASH. By strongly heating this substance, much of its impurities may be driven off: the purer carbonate remaining is called PEARLASH.

A pure salt may be obtained by dissolving the pearlash, and then letting it crystallize. The salt crystallizes while the impurities are still in solution: at this point the fluid is drawn off, and the crystals left. The symbol of the pure substance is K_2CO_3.

228. Preparation of Potassium Hydrate. — Potassium hydrate (KHO) is obtained by boiling a solution of the carbonate with slaked lime ($Ca(HO)_2$). A reaction occurs, in which the potassium of the carbonate is substituted for the calcium of the calcium hydrate, or slaked lime, and by this action potassium hydrate is formed, which remains in solution, while the calcium carbonate produced falls to the bottom as a heavy powder.

The clear solution is afterward evaporated to dryness, and the solid hydrate is then fused and run into molds.

Properties. — Potassium hydrate is very soluble in water, and has a strong affinity for carbonic acid. It greedily takes both these substances from the air, until at length the entire mass of hydrate is changed into a sirup of the carbonate. To indicate this property of dissolving in water absorbed from the air, the term DELIQUESCENCE is used.

229. Sodium Chloride. — Sodium chloride, so well known as common salt ($NaCl$), is everywhere abundant. In many parts of the world it occurs in thick beds, from which it may be mined. Large quantities are obtained by evaporating the water of salt springs, while immense quantities in solution give its characteristic saltness to the sea.

The uses of this substance are important: not among those least important is its use in the manufacture of the other sodium compounds, especially of sodium carbonate.

230. Manufacture of Sodium Carbonate. — Enormous quantities of sodium carbonate are used in bleaching, soap-making, glass-making, and other arts. The several processes in its manufacture are as follows: —

Sodium Chloride changed to Sodium Sulphate. — When salt is heated with sulphuric acid in a reverberatory furnace, mutual decomposition takes place. Sodium sulphate and hydrochloric acid are the products of the reaction. This may be understood by the equation: —

$$2\,NaCl + H_2SO_4 = Na_2SO_4 + 2\,HCl.$$

The sodium sulphate thus formed is valuable, aside from its use in making the carbonate. It is well known under the name of GLAUBER'S SALTS. In the manufacture now being described, it is called SALT CAKE.

The Sodium Sulphate changed to Sodium Sulphide.

— When sodium sulphate is decomposed by carbon, the following reaction occurs: —

$$Na_2SO_4 + 4C = Na_2S + 4CO.$$

Sodium sulphide and carbonic oxide are produced. In the manufacture of sodium carbonate, the sulphate, with small coal and chalk, or limestone (calcium carbonate), are thoroughly mixed and melted together in a furnace. The above reaction takes place.

The Sodium Sulphide changed to Sodium Carbonate. — The sulphide formed by this reaction is at once changed by the calcium carbonate according to the following equation: —

$$Na_2S + CaCO_3 = Na_2CO_3 + CaS.$$

Sodium carbonate and calcium sulphide are formed. The mixture has a blackish-gray color, and is called BLACK ASH.

The black ash is afterward thoroughly leached; during this process the water dissolves out the carbonate, but leaves the sulphide; and then, finally, the solution is evaporated to dryness: the residue is the crude sodium carbonate of commerce, generally known as SODA ASH.

231. Acid Sodium Carbonate. — Hydrosodium carbonate (bicarbonate of soda) is obtained by exposing sodium carbonate to the action of carbonic acid. Its formula is $HNaCO_3$. This is the substance familiarly known as SODA, and used so commonly instead of yeast in bread-making. It is used in medicine: it is also much used in making effervescing drinks.

SECTION III.

METALS OF THE ALKALINE EARTHS.

232. General Description. — The metals of the second class are bivalent. They form carbonates which are not soluble in water, unless it contains carbonic acid: in this respect they differ from the metals of the first class. The metals themselves are of little use, but some of the compounds of calcium and barium are of considerable importance.

Illustrations of these Class Properties. — The metals of this class, like those of the first, decompose water at all temperatures, but the reaction is somewhat different. If calcium is used for the purpose of illustration, it will be:

$$H_2O + Ca = CaO + H_2.$$

Observe that one combining weight of calcium replaces the two combining weights of hydrogen in water. This illustrates the bivalent character of calcium: the other members of the group are also bivalent.

The calcium oxide (lime), shown in the reaction just given, combines with water, evolving much heat, actually boiling the water (Fig. 94) to form the calcium hydrate (slaked lime), which is very slightly soluble in water, forming LIME-WATER.

Fig. 94.

233. Calcium Oxide. — The oxide is made on a large scale, to be used in making *mortar* and cements, so valuable in building. For this purpose fragments of limestone are mixed with coal and burned in kilns. The carbonic acid of the limestone is driven off by the heat; and the other constituent, lime, or, as it is often called, QUICKLIME, is left, still in the form of hard and compact stone. In contact with water the stone swells, grows intensely hot, and crumbles to powder. The slaked lime thus made is mixed with sand to form MORTAR.

234. Calcium Carbonate. — The members of this group form *carbonates.* Limestone and marble of every kind are composed chiefly of calcium carbonate. This compound is very slightly soluble in water unless it contains carbonic acid, but in water charged with this gas it dissolves readily. The formation of STALACTITES is a beautiful illustration of this action. Water, charged with carbonic acid, flowing through soil and over rocks where limestone is abundant, dissolves this substance. Finding its way to caverns, it falls, drop by drop, from the roof. Exposed to the air, carbonic acid evaporates; the water can no longer hold the carbonate in solution, but deposits it wherever it rests. Drop by drop, for a moment clinging to the roof, leaves its mite of carbonate behind, until pendent masses, like icicles, sometimes of curious shape and beauty, are formed. The carbonate forms, at the same time, on the bottom of the cave, upright masses called STALAGMITES.

SECTION IV.

METALS OF THE EARTHS.

235. General Description. — Aluminium is the most important metal of the third class. The others are metals lately discovered by means of the spectroscope, — indium in

1863, and gallium in 1875. These metals decompose water at high temperatures, and form sesquioxides with its oxygen.

236. Aluminium. — This element has a combination of properties which renders it one of the most interesting in the whole series of metals. It has the hardness and luster of silver; and, since it does not tarnish when exposed to air and vegetable acids, it would seem to be fitted for the practical uses to which silver is put. It melts only at a high temperature, and may then be cast into any desired form. This, together with its malleability, ductility, and tenacity, would enable it to replace iron for many purposes, while its lightness (density 2.56) and beauty give it advantage over that metal.

Occurrence in Nature. — In connection with these valuable properties we find that aluminium is one of the most abundant elements in nature. It is a constituent of clay and marl, of slate, and indeed of most rocks and soils. It must constitute about one-twelfth of the solid parts of the earth.

But no cheap method of extracting the metal is yet known, and the expense stands in the way of its application in the arts. It is now manufactured on a commercial scale in England and France, and it has been used for ornamental work and in making physical apparatus where strength and lightness are required.

237. Compounds. — The most important compounds of this metal are aluminium sulphate and alum. The formula of the first is $Al_2 (S O_4)_3$. It is largely used in calico-printing.

Alum is a double salt: it contains the metals potassium and aluminium, both as sulphates. Its formula is $K_2 Al_2 (S O_4)_4 + 24 H_2 O$, and its chemical name is potassio-aluminium sulphate. The water, $24 H_2 O$, is held in its crystals, and is called its "water of crystallization." Heat a few crystals of alum, and this water will be expelled: the alum will remain bulky, white, and opaque; it is then called "burnt alum."

Other metals beside potassium may combine with the aluminium, as sulphates, and other kinds of alum are thus produced. Thus common alum is a potassium alum. So also we have

$Na_2Al_2(SO_4)_4$ . . . Sodium alum,
$(NH_4)_2Al_2(SO_4)_4$ . . . Ammonium alum.

SECTION V.

METALS OF THE ZINC CLASS.

238. General Description. — The metals of this class are alike fusible at quite low temperatures, volatile at temperatures at or below a bright red heat, and combustible when heated in the air. They are bivalent, and form, each, but one oxide, chloride, and sulphide. They decompose water at high temperatures, and dilute acids at low temperatures, liberating hydrogen gas.

Illustrations of the Class Properties. — The melting-points of these metals are comparatively low; of zinc at 423° C., of magnesium a little higher, and of cadmium a trifle lower.

Heated to a bright red heat, magnesium is changed to vapor; at a low red heat cadmium vaporizes, and zinc at a temperature between these extremes.

At a red heat in the air these metals burn. Cadmium gives the vapors of its oxide; zinc, with a blue flame, forming clouds of vapor; and magnesium, with a flame of most dazzling brightness, sometimes used as a source of light in photography and in optical experiments.

At high temperatures they decompose water, and form oxides. In this reaction one atom of metal replaces two of hydrogen, and forms an oxide with the one atom of oxygen; and hence they are bivalent.

On contact with sulphuric or hydrochloric acid they displace the hydrogen, and form salts. Indeed, by this action of zinc we have seen that hydrogen is generally prepared.

239. Zinc. — Zinc *blende* (a sulphide), *calamine* (a carbonate), and the *red oxide*, are the ores of zinc found most abundantly; and from these the metal is extracted. When either of the first two is used, it is first *roasted*, that is, heated in presence of air. By this means it is changed to the form of *oxide*. The oxide is then mixed with coal, and heated in a close vessel having an iron tube reaching over into a cold receiver. The oxide is decomposed; and its zinc, in the form of vapor, goes over to be condensed in the receiver.

Properties of Zinc. — Zinc is a bluish-white metal, about seven times (6.9) as heavy as water.

At low temperature zinc is brittle; heated to about 200° C. it is also very brittle; but between these extremes (130° C.) it is malleable, and is rolled or hammered into thin sheets for various uses.

Zinc is not easily acted upon by air, and on this account it is sometimes used as a coating to protect iron from rust. Iron covered with a thin coating of zinc is said to be *galvanized.*

SECTION VI.

METALS OF THE IRON CLASS.

240. General Description. — In this group we find iron, manganese, nickel, and cobalt. The last three named more or less resemble iron, both in physical and chemical properties. It needs the intense heat of the furnace to melt iron: the same is true of the others. Even at so high a temperature iron does not volatilize: neither do the others. Iron is generally bivalent, but sometimes quadrivalent: so are the rest; and each forms several oxides, sulphides, and chlorides. Like the preceding class, these metals decompose water at high temperatures, and acids without the application of heat.

241. Iron. — Pure metallic iron is of very rare occurrence in nature. The metal is found, however, in great abundance in combination with non-metals; its oxides and its sulphides being the most common ores. The *magnetic oxide* occurs in large quantities in many parts of the United States, and is used extensively in the manufacture of the metal. It is the ore from which the best "Swedish iron" is also made. In England an impure carbonate, called *argillaceous iron-ore*, is chiefly used.

The iron of commerce occurs in three forms, — cast-iron, wrought-iron, and steel.

Cast-iron is a compound of iron with small and variable quantities of carbon. It is obtained from the ores by heating them in a blast-furnace.

Wrought-iron is nearly pure iron, but still contains a very small proportion of carbon. It is obtained generally from cast-iron by burning out its carbon in a reverberatory furnace.

Steel is also a compound of iron and carbon, containing less carbon than cast-iron, but more than wrought-iron. It is made from cast-iron by burning out its carbon, or from wrought-iron by adding carbon to it.

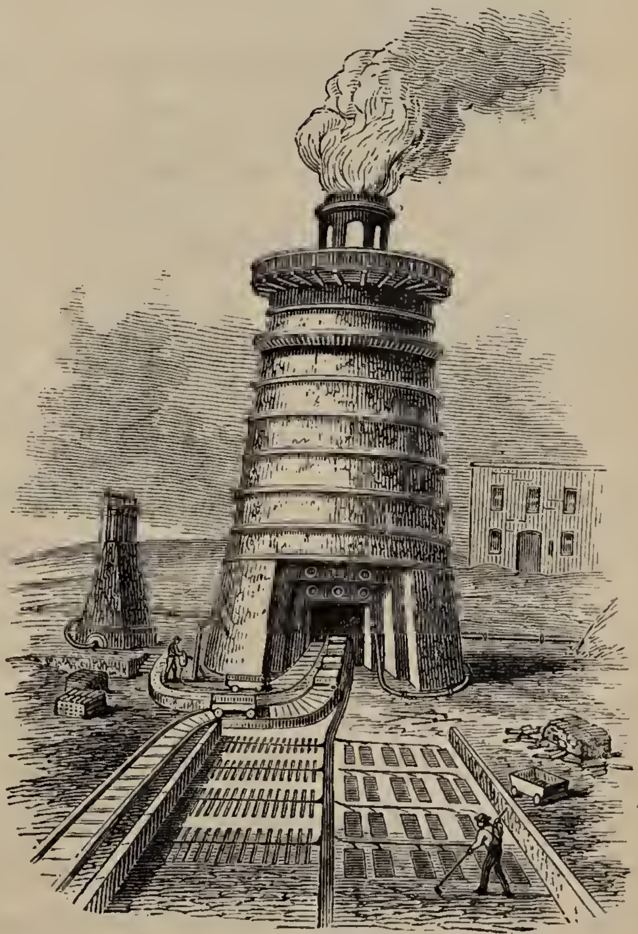


Fig. 95.

242. Cast Iron. — In the most ancient of histories we read of Tubal-Cain, who was the great-grandson of the son of Adam. We are told that he was a cunning workman in brass and *iron.*

Even in such early days this metal must have been extracted from its ores.

The ores are reduced by heat in a *blast-furnace.*

The Blast-Furnace. — The exterior of a blast-furnace is shown in Fig. 95. This, together with Fig. 96, which represents a section of it, will enable us to explain its construction. The interior has the shape of a double cone. It is built of the most infusible fire-brick, and inclosed in solid masonry. It is very large, perhaps fifty feet high by fifteen feet in width at its widest part. The bottom is closed, and the air needed for the fire is forced by steam-engines through pipes, T. The fuel and the ore, with broken limestone or other *flux*, are put in at the top: the metal is drawn off at the bottom.

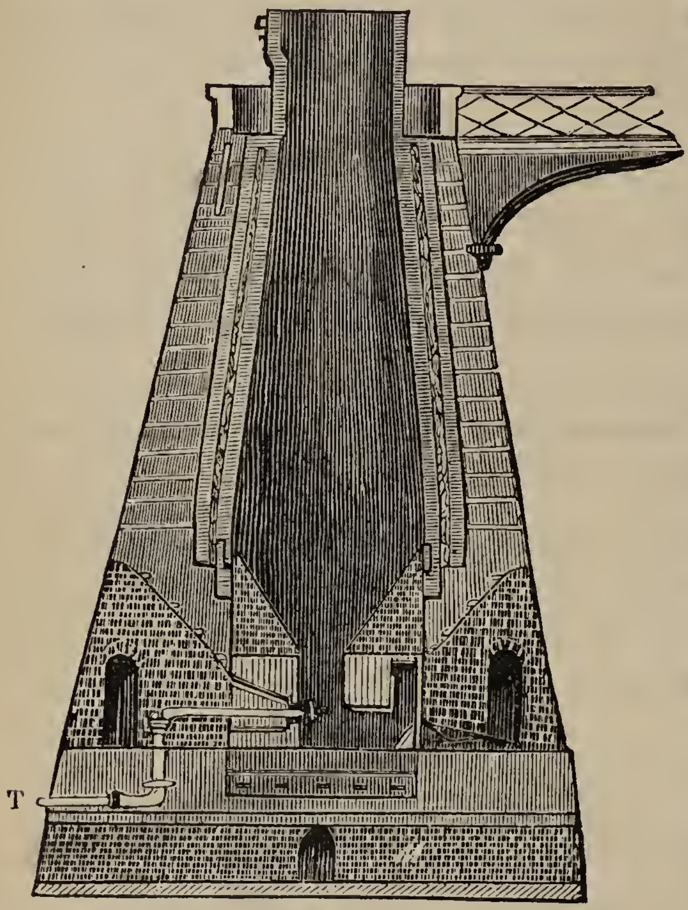


Fig. 96.

The Process. — The ore, if necessary, is first roasted, by which it is changed to the form of *oxide*. The oxide mixed with fuel and flux is made to *fill* the furnace. The fire, having been started, is kept up by the blast of hot air (a cold blast sometimes) driven by the engine. "Where the blast first touches the burning fuel, carbon dioxide is formed: this gas, with nitrogen, rising through the furnace, comes in contact with white-hot carbon, and is reduced to carbon oxide. The layers of solid material thrown in at the top of the furnace gradually sink down; and, as soon as a stratum of ore has gone far enough to be heated by the hot mixture of nitrogen and carbon oxide, it becomes reduced to spongy metallic iron, which, mixed with flux and the earthy impurities of the ore, settles down to hotter parts of the furnace, where it enters into a fusible combination with carbon, while the flux and earthy impurities melt together to a liquid slag. The liquid carburetted iron settles to the very bottom of the furnace, whence it is drawn out at intervals through a *tapping-hole*, which is stopped with sand when not in use." (*Eliot and Storer.*)

The Drawing-off. — In front of the furnace is a large level bed of sand. A channel is scooped through the middle of this bed, and it reaches all the way from the bed to the hearth of the furnace. From the large channel in the middle of the bed of sand, there are smaller ones reaching out each side, and then from these there are other branches about three feet long and three or four inches wide. Look at the picture (Fig. 95), and you will see the arrangement of this central channel with its branches.

Now, about every twelve hours the furnace is opened at the bottom for the melted metal to run out. It has no choice of places: it must flow down through the large channel, and off into all the branches in the bed of sand. There it is allowed to cool.

The iron thus obtained is the *cast-iron* of commerce. The short bars cast in the sand are known as *pigs*. Indeed, the crude cast-iron from the blast-furnace is often called PIG-IRON.

243. Casting. — Large quantities of cast-iron are used for various purposes. In order to make it into useful articles it is melted, and the liquid metal is poured into molds. This operation is called CASTING. The melting of the iron is accomplished in what is called a CUPOLA FURNACE.

The Cupola Furnace. — This furnace is built of fire-bricks, and incased in iron. It is several feet high, and cylindrical, with a door at some distance above the bottom for the iron and fuel to be put in, and a hole, known as the tap-hole, at the bottom, for the melted metal to run out.

In the first place, molds of the desired article are made in sand.

Then, when the melted cast-iron is ready in the furnace, the tap-hole is opened. A fiery stream of liquid iron pours out: it is caught in *ladles* by the workmen, who hurry it away to the molds. The fluid metal runs into every little groove and corner of the mold, and there hardens into the desired form.

244. Wrought-Iron. — Wrought-iron is the purest form of commercial iron, and is generally obtained by treating cast-iron in a reverberatory furnace.

The Reverberatory Furnace. — A section of a reverberatory furnace is shown in Fig. 97. The cast-iron is placed at D upon the hearth. The fire, A, is separated from the hearth by a wall of fire-brick. The roof is arched downward to the chimney, which is forty or fifty feet high, to cause a strong draught.

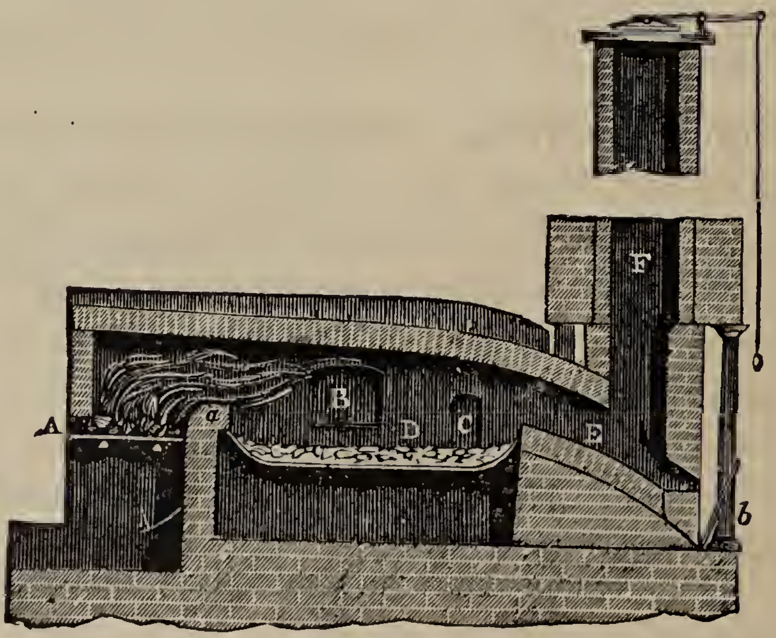


Fig. 97.

The Process. — Flame and hot gases from the fire, striking against the arched roof, are reflected down upon the cast-iron. In a little time the iron begins to melt. When it has been all reduced to a pasty state, the furnace-man unstops an opening (B), through which he puts his paddle. By thoroughly stirring (*puddling*) the pasty mass, all parts are brought in contact with the hot air, during which a part of its impurities, in the form of slag or scoria, is allowed to run off, while its carbon is burned, and its gas escapes to the chimney. The metal is then formed into balls, taken from the furnace, pressed or hammered to remove the remaining scoria, and then rolled into bars or other forms of WROUGHT or MALLEABLE IRON.

245. Steel. — The difference between steel and both cast-iron and wrought-iron is in the quantity of carbon it contains. It contains less than cast-iron and more than wrought-iron. Formerly steel was made by adding carbon to wrought-iron: lately it is largely made by taking a part of its carbon from cast-iron.

Steel from Cast-Iron. — From two to six tons of cast-iron, when melted, is run into a large globular vessel, built of the most infusible substance. Numerous holes in the bottom of this crucible allow a strong blast of air to bubble up through the melted metal. A most violent combustion follows, the heat of which keeps the metal in a fluid state, while its carbon and a small part of the metal itself are burned to oxides. Too much carbon is, by this process, removed, and a quantity of cast-iron is added to restore carbon enough to change the whole mass into steel. The crucible is then tipped upon its pivots, and the melted steel run off into molds. Less than half an hour is enough to change these tons of cast-iron into cast-steel. The process is known as the BESSEMER PROCESS.

From Wrought-Iron. — The older method of preparing steel is called "cementation." Bars of wrought-iron, embedded in charcoal and inclosed in boxes from which air is very carefully excluded, are heated to redness, and kept in this condition for several days. A curious and obscure chemical action goes on, by which these two *solid* substances unite, — the carbon penetrating and combining with all parts of the iron, and thus changing it to steel. To make its composition uniform, this "blistered steel," as it is called, is melted, and cast into large but short bars called ingots.

SECTION VII.

METALS OF THE TIN CLASS.

246. General Description. — Tin is the only useful metal of its class. The others resemble tin in their chemical properties, but they are rare and unimportant. These metals are quadrivalent, and decompose water at high temperatures, forming dioxides.

247. Tin. — Tin is not an abundant element in nature, and yet it is one of the metals longest known to men. The mineral called *tinstone* (stannic oxide, $Sn\,O_2$) is the chief source of the metal. Mixed with powdered coal and a little lime, the ore is spread upon the hearth of a reverberatory furnace. The carbon takes the oxygen from the ore, and the melted metal is run off into iron molds.

In color tin resembles silver. It is soft, malleable, and ductile.

Tin does not easily lose its luster by exposure to air, and on this account it is largely used as a covering for other metals: common tin-ware is sheet-iron, which has been covered with a thin coating of tin.

SECTION VIII.

METALS OF THE ANTIMONY CLASS.

248. General Description. — Of this class arsenic, antimony, and bismuth are the most important metals. They are trivalent, and are very closely related to the trivalent group of non-metals. In bismuth the metallic character is very clear, in antimony it is less evident, in arsenic it is very doubtful, and in phosphorus and nitrogen it is altogether absent. From bismuth to nitrogen, the transition from metal to non-metal is gradual and perfect.

249. Arsenic. — This element has been already described as a non-metal, and its relation to nitrogen and phosphorus pointed out. On the other hand, it is related to antimony and bismuth, whose metallic character is more decided. It would seem to be the bond which links these two divisions of the elements together.

250. Antimony. — Like arsenic, antimony combines with three atoms of hydrogen to form a combustible gas. Treated in Marsh's apparatus, it also forms a stain upon white porcelain, or a metallic mirror in the tube. So great is the resemblance between the stains of antimony and arsenic, that care is to be used not to confound the two metals in the test. The antimony stain is known by its more feeble luster, its blackness, and the high heat needed to volatilize it.

Its most Useful Alloy. — Antimony, with tin and lead, are melted together to make *type-metal*, the most useful of all alloys, since the art of printing depends upon its use. It has the curious property of expanding when it cools from the melted liquid to the solid form. On this account, when poured into the type-molds, and allowed to become cold, it fills every little groove and marking of the mold, and thus takes the perfect shape of the type.

251. Bismuth. — Bismuth forms oxides and chlorides whose composition is analogous to those of arsenic and antimony. It is trivalent, like the others.

The metal itself is of a reddish hue, and, like the other two, very brittle. It melts at 264° C. (507° F.) ; and, when cooling, it crystallizes and *expands*.

With other metals bismuth forms alloys remarkable for the low temperature at which they melt. Its alloy with lead and tin (two parts bismuth, one of lead, and one of tin) is called "fusible metal:" it melts at 93° 7 C. (200° F.). This alloy is used for taking casts from medals and dies: on cooling, it expands, and, filling every crevice or line of the die, makes a most beautiful and faithful copy.

SECTION IX.

METALS OF THE LEAD CLASS.

252. General Description. — This class contains only two metals, lead and thallium. Lead has been known from the earliest times: thallium has been known only since 1861. They are alike in being soft enough to yield to pressure by the finger-nails; in being fusible below red heat; in being malleable and ductile; and in being heavy metals, lead 11.2 and thallium 11.8 heavier than water.

In their chemical properties these metals do not so perfectly agree. Lead is bivalent: thallium is univalent. In other respects, also, thallium resembles the alkaline metals.

253. Lead. — The ore from which the lead of commerce is obtained is a sulphide ($Pb\ S$), called galena. The color and luster of this ore are much like that of the metal itself, but the ore is much harder. It is crystalline, and sometimes occurs in cubes of the most perfect form.

To obtain metallic lead, galena is mixed with lime, and roasted in a reverberatory furnace. By this means, a part of the ore is changed to oxide, another part to sulphate, and some remains a sulphide as it was. The air is then shut off from the furnace, and the heat raised: the compounds are then all decomposed, and metallic lead is produced.

Its Uses. — Metallic lead has numerous and important uses. Among them we may especially notice the construction of water-pipes and cisterns. In cities supplied with water from reservoirs, the conduit-pipes are almost universally made of lead.

Action of Lead upon Water. — But, since the compounds of lead are poisonous, the chemical action of lead upon water is a matter of great importance. Together, especially in the presence of air, they form an oxide which is somewhat soluble in water. But this corrosive action is very much modified by the presence of impurities. It is increased by ammonium nitrate: it is diminished by sulphates and carbonates.

Water containing a solution of calcium carbonate scarcely affects the lead, because an insoluble carbonate is formed on the surface of the metal which prevents any further action; but, if it contains much free carbonic acid, this removes the lead carbonate, which otherwise would protect the surface, and leaves the metal of the pipes constantly exposed to corrosion.

Whether drinking-water may be kept in lead cisterns and pipes without risk to health, depends on the character of the water; and the question can be decided only by learning what substances the water holds in solution, and what is their chemical action on the metals.

SECTION X.

METALS OF THE SILVER CLASS.

254. General Description. — The members of this class, copper, mercury, and silver, are sometimes found native, but they are much more abundant in combination. Their sulphides are their most common ores. They can not decompose water at any temperature. Copper and mercury are bivalent, but silver is univalent.

255. Copper. — Free metallic copper is found in the noted mines of Cornwall and Devon, in England, and in many other parts of Europe. But some of the finest native copper in the world is found in the region of Lake Superior. One single mass of Lake Superior copper weighed over 400 tons.

The native copper has very curious crystalline forms. In some cases it is found in little cubes. In some cases the small crystals are grown together in vast numbers, and thus make large masses; and these masses often show the most singular branch-like shapes (Fig. 98), tempting one sometimes to fancy that the metal had tried to imitate the form of some growing plant.

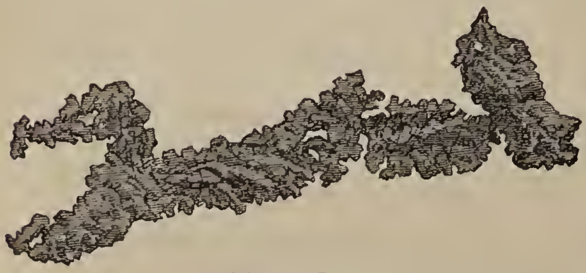


Fig. 98.

Ores of Copper. — *Copper pyrites* is the most abundant ore of this metal. To a great deal of this ore nature has given the form of beautiful cubes, having a color and luster something like brass. This ore is composed of copper, iron, and sulphur, $Cu\,Fe\,S_2$.

Besides this, there are other sulphides of copper, $Cu_2\,S$, and $Cu\,S$. The oxides also are abundant; so also are the carbonates. *Malachite* is one of these.

Malachite is a rich ore of copper, but much less common than many others. It is a stone of most beautiful color. Its rich green surface also takes a fine polish, and on these accounts it is often used for purposes of ornament. It is a compound of copper, represented by the formula $Cu\,C\,O_3 + Cu\,(H\,O)_2$.

Smelting the Ores. — To obtain the metal, the ores of copper are first *roasted* and afterward *melted*. These operations are repeated until the mass contains only two compounds of the metal, the oxide and the sulphide.

The mass is then again roasted; and, during this heating, the two compounds attack each other. The sulphur of one and the oxygen of the other are taken away, leaving the copper of both in metallic form.

Refining. — This crude copper from the ore contains other metals which must be removed. The process of "refining" is for this purpose. It consists in re-melting the metal, and keeping it in the fluid form, exposed to air, for several hours. By this means the impurities are oxidized or burned out. The molten copper is then taken from the furnace in ladles, and poured into molds, which give it the form of solid bars. The furnace, and some of these last operations, are shown in Fig. 99.

Fig. 99.

Properties of the Metal. — Copper is a red metal, very tenacious, very ductile, and malleable.

Copper is among the very best conductors of electricity, and is much used in the construction of electrical apparatus and lightning-rods.

It slowly tarnishes in the air, and is readily acted upon by vegetable acids: the compounds formed are, many of them, poisonous. On this account copper vessels to be used for culinary purposes are usually coated with tin.

Alloys of Copper. — *Brass* is an alloy of copper and zinc: it is made by melting together two parts of copper to one of zinc.

Bronze, gun-metal, and bell-metal are, for the most part, mixtures of copper and tin.

German silver is an alloy of copper, zinc, and nickel.

256. Mercury. — This metal was known in very ancient times; and besides *mercury*, a name which the old chemists gave it, this lustrous fluid was early called *quicksilver*, and this name is still in very common use.

This metal is found in California, in China, in Austria, and to some extent in other countries. Perhaps the most noted among its native countries is Spain: the mines of Almaden in that country have long been worked, and have yielded a rich reward.

Mercury is generally found combined with sulphur. This ore is called *cinnabar*. It is a very pretty mineral, much the color of cochineal. Sometimes this ore is so pure that its color is as rich and brilliant as that of *vermilion*. Indeed, these two things have the same composition: cinnabar is the sulphide which is found in the earth; vermilion is the sulphide which is made in the chemist's laboratory. Both have the composition represented by the formula HgS.

Reducing the Ore. — To obtain the metal, the cinnabar is roasted in a furnace with plenty of air. The oxygen of the air decomposes the ore, and combines with its sulphur, leaving its mercury free. At the temperature of the furnace the metal is in vapor. This vapor is carried over into cold chambers, where it condenses to liquid form.

Properties. — Mercury is a liquid metal about 13.5 times heavier than water. It becomes a silver-white solid at $-39.6°$ C. ($-39.4°$ F.).

Mercury forms two oxides, the mercurous, Hg_2O, and the mercuric, HgO. The last-named substance is a red powder, once known as "red precipitate," and is of historic interest, since by heating it Priestly discovered oxygen in the year 1774. Nevertheless mercury and oxygen have so little attraction at ordinary temperature, that the surface of the metal remains bright on exposure to air.

Its attraction for chlorine is stronger: the two combine at ordinary temperature. Two chlorides are well known: they are the mercurous chloride, $Hg_2 Cl_2$, and the mercuric chloride, $Hg Cl_2$. The former is *calomel*, so long and largely used in medicine; and the latter is *corrosive sublimate*, a virulent poison. Both are white solids; yet they are easily distinguished, because corrosive sublimate is soluble in water, while calomel is not.

Amalgams. — The alloys of mercury are called AMALGAMS. The most familiar example is to be seen on the backs of mirrors: this substance is an amalgam of mercury and tin.

Uses.— Metallic mercury is largely used in "silvering" mirrors, in the construction of thermometers and barometers, and also in extracting silver and gold from their ores.

257. Silver. — Silver, like the other metals of greatest use in the arts, is very widely distributed. It is found native and often alloyed with mercury, copper, and gold; but its sulphide ($Ag_2 S$), either alone or mixed with other metallic sulphides, is its most common form. From these ores the metal is usually obtained. They are found in many countries of Europe, but in greater abundance and richness in Peru, Mexico, and the Pacific slope of the United States.

Obtained by Cupellation. — Galena almost always contains silver, and often in quantities which make it valuable for the extraction of silver as well as lead. "It has been found profitable to extract the silver from lead that contains less than one-thousandth of its weight of the precious metal." From lead rich in silver this metal is obtained at once by *cupellation;* but, when so poor as that just described, the lead is first melted, and then, while cooling slowly, it is stirred. As it cools, it crystallizes; and the crystals may be dipped out of the liquid mass in colanders. Now, an alloy of lead and silver will remain melted when cooled below the temperature at which pure lead solidifies. Hence the crystals dipped out in the colander are crystals of *lead*, while *all* the silver remains in the fluid left behind. By this means the proportions of lead are reduced until the metal is rich in silver, after which it is *cupelled*.

The process of *cupellation* is based upon the fact that lead is rapidly oxidized by air at high temperature, while silver is not.

The alloy, placed in a shallow porous vessel of bone-earth, called a cupel, is melted in a furnace, and its surface is at the same time exposed to a current of hot air. The lead is changed to an oxide: the melted oxide is partly absorbed by the cupel, while another part runs off into other vessels. The silver is not affected by the air; and, when the lead has passed away, the precious metal still remains in the cupel.

Obtained by Amalgamation. — From other ores than argentiferous lead, silver is extracted by a process called *amalgamation*. It is based upon the fact that silver is very soluble in mercury.

For this process, also, the ores require a preliminary treatment. After being thoroughly ground, they are mixed with common salt, and then roasted. By this means the silver is changed to a chloride. The roasted substance is then mixed with water and fragments of iron, and the mixture is violently shaken in revolving casks. Mercury is soon added, and the agitation is kept up for perhaps twenty hours. In the mean time the iron decomposes the chloride, and the silver thus set free is dissolved at once in the mercury. The excess of mercury in the amalgam is removed by filtering it through leather bags under pressure, and the rest by distillation. The precious metal is left behind.

Properties. — Silver is the whitest of metals. It is very malleable and ductile, much harder than gold, and ten and a half times heavier than water. Its surface is not tarnished on exposure to air, not even when heated in oxygen; but it quickly blackens on exposure to sulphuretted hydrogen, because of the strong attraction between silver and sulphur.

Its Alloys. — Silver is too soft for most purposes in the arts. Small quantities of other metals increase its hardness. For coin and for articles of plate, it is alloyed with copper. The English *standard silver* contains 7.5 per cent of copper. The standard is regulated by law, and is not the same in all countries. In the United States the legal silver coin is $\frac{9}{10}$ silver and $\frac{1}{10}$ copper.

258. Action of Light on Silver Compounds. — We have seen that light is an active agent in many chemical changes: we may add that the salts of silver are especially sensitive to its influence. An easy experiment can be made to illustrate this fact.

On the Chloride.— Into a test-tube put a quantity of a solution of silver nitrate, and add a few drops of hydrochloric acid. A heavy white precipitate of silver chloride is at once formed, which speedily settles to the bottom. Let the tube be now placed in the sunlight, and an interesting change of color will be gradually produced. The snow-white chloride becomes pink, violet, brown, and at last dark bronze-black. The chloride, dry and pure, will not show this change on exposure to light: moisture seems to be necessary. Light causes a reaction between water and the chloride, by which hydrochloric acid is formed, and a small quantity of metallic silver is set free. To the presence of this metal in a state of very fine division the darkening is due.

On the Nitrate. — Silver nitrate is not changed by the action of light unless in contact with organic matter: it then blackens like the chloride.

On the Oxide. — When silver iodide is exposed to light no *visible* change occurs, but still a molecular change does take place. This curious fact may be shown by experiments. Let two highly polished plates of silver be held in iodine vapor in the dark: by this means a thin film of iodide will be made over their surfaces. Leave one plate in the dark, and put the other, for a time, in the sunlight: the two plates still look alike; the light has caused no visible change. But now hold both, in a box at moderate temperature, over the vapor of mercury: the one which was exposed to the light immediately blackens; the other is not changed. The darkening in this case is due to a combination of mercury with the silver of the iodide, — the mercury decomposing the iodide after exposure to the light.

The nature of the action of light in this case has been in doubt. The best explanation (Amer. Jour. of Sci., vol. 42, p. 198, and vol. 44, p. 71) supposes it to be a *physical* action, by which the molecules are so disturbed as to yield afterward to the attraction of mercury. That is to say, the molecules of the iodide are so shaken by the impact of the waves of light that their atoms are almost ready to fall apart. In this condition the attraction of the mercury for the iodine is strong enough to complete the decomposition.

259. Photography. — Photography is the art of making pictures by means of light. While there are many different processes, each requiring different substances sensitive to light, still the most interesting process is the one so universally used in portraiture; and this is based on the sensitiveness of the compounds of silver.

The process consists of two parts; viz., the making of the negative, and then the making of a positive.

260. Making the Negative. — Glass is employed as the substance on which the picture taken directly from the object is held. But in order to use it the surface must be made sensitive to light.

Preparation of the Plate. — The object is to give the glass a coating which contains silver iodide.

Cotton, acted on by a mixture of nitric and sulphuric acids, is changed to gun-cotton, or pyroxyline; and this, when dissolved in alcohol and ether, forms the substance called COLLODION. A small quantity of an iodide (potassic), or a mixture of two or three iodides, is put with the collodion, and the solution is then rapidly poured from a wide-mouthed vessel over the very clean and dry surface of the glass: the alcohol and ether quickly evaporate, and leave a thin film upon the glass, which is then plunged into a bath of silver nitrate. While in this bath for a few minutes, the iodine in the film takes silver from the nitrate, and forms silver iodide ($Ag\,I$). The plate being taken from the bath, thoroughly washed and dried, is ready to receive the action of light.

Exposure in the Camera. — The plate, with the sensitive film on its surface, is exposed in a camera for a short time, during which the light from the object performs its curious action upon the iodide, by which no picture is made visible, but by which one is prepared to be developed. On pouring over the film of iodide a solution of pyrogallic acid ($C_6\,H_6\,O_3$) or of ferrous sulphate ($Fe\,S\,O_4$), containing a few drops of silver nitrate, the picture is brought out, or "developed" as the process is technically described.

Fixing. — The next step is to dissolve away the undecomposed iodide, in order to make the picture permanent; for it is evident that if any of this remains in the film the light will continue to work, and thus destroy the picture. This is done by washing the plate in a saturated solution of sodium hyposulphite. After this the plate is thoroughly washed in water, dried, and varnished on the picture side, to preserve the film from injury.

The Lights and Shades are Inverted. — When a photograph on glass is viewed by transmitted light, it will be seen that the shades of the object are lights on the picture, the lights and shades being reversed. Such a picture is called a NEGATIVE. If the "development" be pushed farther, more of the iodide in the film will be decomposed, and the shades will more completely intercept the light. From such a negative a picture may be printed upon paper. This picture of the negative reverses the lights and shades again, and hence brings them true to the object. Such a picture, in which the lights and shades are the same as those of the object itself, is called a POSITIVE.

261. Making the Positive. — The paper for this purpose is prepared by floating it first upon a solution of salt, and afterward upon one of silver nitrate. Silver chloride ($Ag\,Cl$) is thus formed in the paper. This paper, placed under the negative, is exposed to light; the light is stopped by the shades of the negative, but passes freely through its lights: upon the paper, therefore, the lights and shades correspond with those of the object from which the negative was taken.

The photograph is next to be thoroughly washed and then "toned." The toning-bath is a solution of hydrosodium carbonate (bicarbonate of soda), containing a little gold chloride ($Au\,Cl_3$). The change produced by this bath is beautiful to witness. From an unpleasant, dull, reddish color, the picture visibly changes to a rich blue-black. The picture, having next been washed in water to remove all traces of the gold bath, is soaked for some minutes in sodium hyposulphite, to wash out all of the undecomposed chloride, and afterward for twenty-four hours in water, to remove all traces of the hyposulphite.

SECTION XI.

METALS OF THE GOLD CLASS.

262. General Description. — The metals of the gold class are always found in the metallic state. They are not tarnished by air at any temperature short of their melting-points, nor can they be dissolved by nitric acid. Only chlorine or aqua-regia can dissolve them. Gold and platinum are the most familiar and important members of the class.

263. Gold. — In small quantities gold is very widely distributed in nature. Fine grains of it occur in the sands at the bottom of many rivers, in the crystalline rocks and the soils derived from them. Iron pyrites, found almost everywhere, often contains traces of gold; and silver is never found entirely free from it. Few places, however, seem to possess the precious metal in quantities that will pay for the laborious work of separating it from the sands or rocks in which it is found.

Separation of Gold. — Except osmium, iridium, and platinum, gold is the heaviest of metals. It is very much heavier than sand; so that when gold-bearing sand is violently shaken in water, the fine and precious grains quickly sink to the bottom, while the sand may be poured off with the water. By repeating this process, called WASHING, the sand or other loose material is finally all washed away. The fine metallic grains left behind are seldom pure gold. The baser metals — silver, copper, and others — are taken out by sulphuric or nitric acid, in which they will dissolve, but which can not dissolve the gold. This process of separating gold from the metals with which it is alloyed is called REFINING.

When the native gold is scattered in fine grains through solid rock, it is extracted by AMALGAMATION. Mercury mixed with the crushed ore dissolves the gold: the amalgam is then separated from the excess of mercury by filtration, and the remainder of the mercury is afterward driven away completely from the gold by distillation.

Properties of Gold. — Gold is remarkable for its fine yellow color and beautiful luster. This metal is about nineteen times heavier than water; its specific gravity is 19.33. It is the most malleable of metals: it is said that leaves have been beaten so thin that 280,000 would be needed to make an inch in thickness! There is a curious fact about the color of gold-leaf; it is this: looked at in the usual way it is yellow; but let a leaf be spread upon glass so that it may be viewed in transmitted light, and it exhibits a fine green color.

264. Platinum. — Platinum is a rare metal, not so widely distributed as gold. The slopes of the Ural Mountains, Brazil, and Peru are among the localities richest in this metal. It is always found in the metallic state, but never pure. It is heavier than gold, and is obtained from loose sands or soils by washing in the same way. Its commercial value is about one-half that of gold and about eight times that of silver.

Pure platinum is almost as white as silver, and can not be tarnished in air at any temperature. The most intense heat of the blow-pipe is needed to melt it. On these accounts platinum vessels are much used in the laboratory. Few indeed are the chemicals which can affect it. Aqua-regia and the caustic alkalies are among those which can. On the other hand, it readily forms alloys with most other metals; and these are more easily melted than the metal itself, and are soluble in acids.

REVIEW.

I. — SUMMARY OF PRINCIPLES.

265. Malleability, ductility, conductivity, and luster are the most characteristic properties of the metals.

They occur in nature in the form of *ores*, in which they are combined with non-metals; as *alloys*, in which they are combined with one another; and as *native* metals, in which condition they are uncombined.

The art of extracting metals from their ores is called *metallurgy*.

Roasting and *reducing* are two important operations in metallurgy. Roasting consists in heating the ore in the presence of air: its object is to change the ore to an oxide of the metal. Reducing consists in heating the oxide with carbon, or some other reducing agent: its object is to decompose the compound, and set the metal free.

The metals of the alkalies are lighter than water, soft as wax, univalent, and have so strong attraction for oxygen that they decompose water at low temperatures. Potassium and sodium are the leading members of the group.

Potassium is a constituent in plants; and when wood, for example, is burned, the potassium is changed into potassium carbonate.

The carbonate is obtained in solution from the wood-ashes by treating them with water. This *lye* yields the solid *potash* by evaporation, and the potash becomes *pearlash* by heat. A solution of pearlash yields crystals of pure potassium carbonate.

Sodium chloride is the most abundant compound of sodium. It is largely used in making other useful compounds of sodium.

Treated with sulphuric acid this salt yields sodium sulphate, or *Glauber's salts.*

The sodium sulphate is reduced by carbon to sodium sulphide.

Sodium sulphide is changed by calcium carbonate into sodium carbonate, or *black ash.*

Black ash, by leaching and evaporation, yields the *sodium carbonate* ($Na_2 C O_3$) of commerce. The manufacture of sodium carbonate from sodium chloride ($Na Cl$) is an important industry.

Treated with carbon dioxide, sodium carbonate becomes *hydrosodium carbonate.* This is the substance known as baking-soda.

The metals of the alkaline earths are bivalent, with strong attraction for oxygen. They oxidize rapidly on exposure to air, and decompose water at ordinary temperatures.

Calcium is the most common metal of this class. Its compounds are abundant in nature and useful in the arts.

Aluminium and indium are called the "metals of the earths."

Aluminium is very abundant in the earth: it is a constituent of all clay soil and rocks, in which it exists chiefly as the silicate.

The metals of the zinc class can be easily melted and vaporized, and burn freely when heated in air. They decompose water at high temperatures, and evolve hydrogen from dilute acids.

Zinc is the leading member of the class. It is a bluish-white metal, malleable at temperatures between 100° and 150° C. and not easily oxidized by exposure to air.

The metals of the iron class melt with difficulty, and are not volatilized by the intense heat of the furnace. Iron is the leading member of the class.

In commerce iron occurs in three forms; viz., cast-iron, wrought-iron, and steel.

These differ in the quantity of carbon in combination with iron.

Iron ore is reduced in the blast-furnace, which yields the metal in form of *pig-iron.*

The cast-iron is changed to wrought-iron in the reverberatory furnace by the action of the oxygen of the air.

Steel is made by burning the carbon out of cast-iron in the *Bessemer* process, or by causing carbon to combine with wrought-iron in the *cementation* process.

Of the tin class, tin is the only useful metal.

Tinstone, which is an oxide of the metal, is the only ore from which tin is obtained.

Tin is a bluish-white, malleable metal, not easily oxidized, retaining its luster on exposure to air.

Common "tin" is sheet-iron coated with a film of tin.

Of the metals of the antimony class, antimony, arsenic, and bismuth are the most interesting. These metals are closely related to the non-metals of the trivalent group.

The metals of the lead class are lead and thallium.

Lead is a soft metal, 11.2 times heavier than water.

The principal ore of lead is galena, from which it is obtained by heating with lime in a reverberatory furnace.

The compounds of lead are poisonous, and hence the risk which attends the use of lead for cisterns.

The metals of the silver class, copper, mercury and silver, are often found native. Their sulphides are their most abundant ores.

Copper is a red metal, very tenacious, and very ductile. It is an excellent conductor of electricity. Its surface is tarnished by exposure to air, and it is readily attacked by acids.

Mercury is a liquid metal, 13.5 times heavier than air. It does not combine with oxygen on exposure to air. It forms two oxides and two chlorides. Calomel is the mercurous chloride, and corrosive sublimate is the mercuric chloride.

Alloys of mercury are called amalgams.

Silver is a white metal, very malleable and ductile, not affected by air or oxygen, but blackened quickly by sulphuretted hydrogen.

Silver is obtained from argentiferous galena by cupellation. It is obtained from other ores by first changing them to oxides, then reducing the oxides by iron, and finally separating the silver by amalgamation.

Silver coin of the United States is an alloy with copper, containing ninety per cent of silver.

The metals of the gold class are always found native. They are not affected by oxygen, and can not be dissolved in strong acids. They have a strong attraction for chlorine, and dissolve in aqua-regia.

Gold is a soft and heavy metal, yellow by reflected light, and green by transmitted light. It is malleable in the highest degree.

Platinum is almost as white as silver, not affected by oxygen, but soluble in aqua-regia.

II. — EXERCISES.

Name the characteristic properties of metals. What of their melting? Of their weight? In what condition are they found in nature?

Into how many classes are they grouped?

Describe the metals of the first class.

Illustrate the class properties. How does potassium act upon ice? Give the reaction. What difference between alkalies and other hydrates?

Describe the manufacture of pearlash. Of pure potassium carbonate.

Describe the manufacture of potassium hydrate. What are its properties?

What is the composition of sodium chloride? How is it obtained? For what is it used?

Show how sodium chloride is changed to sodium sulphate. What are Glauber's salts?

Show how sodium sulphate is changed to sodium sulphide.

Show how sodium sulphide is changed to sodium carbonate. What is the final process in this manufacture?

Give a brief description of hydrosodium carbonate.

Give a brief description of the metals of the alkaline earths.

Illustrate the bivalent character of calcium.

What is slaked lime? For what is it used? How is mortar made? What class of salts do the members of this group form? Are they soluble in water? How are bicarbonates formed? Explain the production of stalactites.

What is said of the metals of the earths?

Describe aluminium.

What is said of the metals of the zinc class?

How is zinc obtained? What are its properties?

Illustrate the class properties, as to melting-point; as to vaporization; as to combustion; as to their action on water.

What is said of the metals of the iron class?

In what ores chiefly is iron found? Name its three commercial forms.

What is cast-iron? How obtained?

Describe the blast-furnace.

Describe in full the process of getting cast-iron from its ores.

What is wrought-iron? How obtained?

Describe the reverberatory furnace.

Describe in full the process of making wrought-iron.

What is steel? How made?

Describe the "Bessemer process."

Describe the process of cementation.

What is said of the metals of the tin class?

How is tin obtained? What are some of its properties? For what is it used?

What is said of the metals of the antimony class?

What is said of arsenic?

Describe antimony.

What is type-metal?

Describe bismuth.

What is said of its alloys?

What is said of the metals of the lead class?

What of the properties of lead and thallium?

From what ore is lead obtained? Describe the process.

What are the uses of lead? What is said of its use for water-pipes?

What is said of the metals of the silver class?

How does copper occur in nature?

Describe the smelting of copper ores. Also, the process of refining.

What are the properties of copper?

Describe its alloys, — brass, bronze, gun-metal, German silver.

In what condition is mercury found? What is a peculiar property of this metal? What are its uses? What is calomel? What is corrosive sublimate? What are amalgams?

In what condition is silver found in nature? Name some important localities.

By what process is silver obtained from galena rich in silver? With poor galena, what process is first used? Upon what principle is cupellation based? Describe the process.

With what ores is amalgamation used to obtain the silver? What preliminary treatment needed? Describe the process of amalgamation.

What is said of the alloys of silver?

What art depends upon the chemical action of light?

What are photographs?

How is collodion made? How is a plate of glass coated with it? What is the effect of light upon the coated plate? How is the picture developed? Describe the rest of the process.

What is a "negative"? How is the paper prepared, and the "positive" made? What other processes are gone through with to finish the picture?

What is said of the gold class?

Where is gold found, and in what condition? How is it separated from sand? From baser metals? From rocks in which fine grains are scattered?

How does platinum occur in nature? How is it obtained? What are some of its properties?

APPENDIX.

EASY EXPERIMENTS FOR THE CLASS-ROOM.

It is recommended that the following experiments be made, *in addition* to those described in the text. The numbers in heavy type refer to the paragraphs of the book which the experiments more fully illustrate or extend.

No attempt is here made to furnish a manual of chemical manipulation. Remembering the conditions which surround the teachers of chemistry in a majority of our schools, it is believed that no genuine service would be done by calling their attention to costly apparatus, lengthy or delicate processes, or even to a large number of experiments. In the selections which follow, the teacher will find those *which he can use*, and which he can use with the smallest expenditure of money, time, and patience, because they combine in the highest degree, —

Simplicity of details,
Certainty of results,
Cheapness of materials,
Adaptation to the subject.

1. Physical Changes. — Ex. 1. Provide a Bunsen's burner, a piece of platinum wire about two inches long, a piece of magnesium wire or ribbon about six inches long, and a pair of forceps.

With the forceps, hold the platinum wire in the flame of the lamp, and notice that it glows with a bright red heat, but that it suffers no change in its nature.

With the forceps, then hold the magnesium with one end in the flame, and notice that it glows with a red heat, then takes fire, and burns with vivid brightness, and that it is at the same time changed into an entirely different kind of matter.

3. Chemical Change. — Ex. 2. Provide a test-tube, a saturated solution of calcium chloride, and some dilute sulphuric acid (1 of acid to 4 of water). Fill the tube to the height of two inches with the calcium chloride, and then add, all at once, the dilute acid two inches more, and shake quickly, or stir with a glass rod. These colorless liquids combine, and produce a white solid, which will not fall out of the tube, even when held bottom upward.

Ex. 3. Fill a cylinder two-thirds full of water, and add about fifty cubic centimeters of a solution of lead acetate. Then add, little by little, some potassium chromate. These *colorless* liquids yield a *rich yellow* solid.

10. Mechanical and Chemical Attractions. — Ex. 4. Take two pieces of plate-glass, very clean and smooth, and slide one upon the other, gently pressing them together at the same time. *Notice* that they cling firmly. The lower one may be lifted by the upper. They are held together by *cohesion*, or, as we are accustomed to say, by adhesion. Cohesion is the generic name given to all attraction among molecules.

Ex. 5. Use a magnet with small nails, and *notice* the effect of *magnetic* attraction.

No permanent effect is left on bodies by the action of these attractions.

Ex. 6. Take two wide-mouthed bottles of equal size. Moisten the inside of one with ammonia, and the inside of the other with hydrochloric acid. Invert one, and stand it upon the other, mouth to mouth. *Notice* that, whereas the contents of both were colorless, both are now filled with white fumes. Two colorless gases, ammonia and hydrochloric acid, have united to form a white solid. They are held in combination by chemical attraction.

11. Influence of Cohesion. — Ex. 7. Mix a little coarsely powdered copper sulphate with an equal quantity of coarsely powdered potassium ferrocyanide in a mortar, and *notice* that no chemical change occurs. Grind them together: still no chemical change. Sprinkle a little of the fine powder into a cylinder of water, and *notice* the dark red-brown solid produced.

Ex. 8. Mix a little ferrous sulphate, in powder, with a little powdered potassium ferrocyanide: no chemical action occurs. Sprinkle a very little of the mixed powder upon water in a cylinder, and notice the fine blue compound produced.

14. Influence of Light. — Ex. 9. Into a cylinder of water put a small quantity of silver nitrate, and then add a little hydrochloric acid. Observe the dense *milk-white* precipitate which forms. Place the cylinder in the strong sunlight, and notice the change from white to a *purplish-black.*

17. No Loss nor Gain. — Ex. 10. Provide two beakers. Into one put about a hundred cubic centimeters of moderately strong solution of lead nitrate, and into the other put half as much solution of potassium chromate. Then put the two upon one pan of the balance, and weigh them carefully. Next, pour the chromate into the nitrate, and restore the beaker to its place. The weight will be found unchanged, notwithstanding the production of a very abundant yellow solid.

24. The Law of Constant Proportions. — Ex. 11. Fill a test-tube about one-third full of water, and add about a quarter of an inch of silver nitrate solution. To this add a few drops of hydrochloric acid, and shake it vigorously. The precipitate will soon settle. To the clear liquid above the precipitate, again add drops of the acid, and repeat the operation until the last drops added give no precipitate at all. The nitrate is then all used up by the acid to make the white chloride.

Then add drops of the nitrate: a white precipitate shows that the acid had been added in excess of what was needed. Add nitrate as long as a precipitate is formed: it will require but few drops.

The slightest quantity of either, beyond a certain definite proportion, remains unchanged.

Ex. 12. Into a small beaker pour about fifty cubic centimeters of hydrochloric acid, and drop into it, little by little, powdered sodium carbonate until the last small quantity produces no effervescence. The acid is then all neutralized. Then stir in drops of the acid patiently until with the last drop the last of the small quantity of solid carbonate disappears.

Each of these substances requires a certain definite proportion of the other: any excess remains unchanged.

Ex. 13. Evaporate the clear liquid obtained in the foregoing experiment, and you will get a good sample of common salt which was produced in the chemical action.

30. The Molecule. — Ex. 14. Dissolve a piece of aniline-red, no larger than the head of a small pin, in a thimble-full of alcohol, and then pour the solution into a half-gallon jar of clear water. Notice the crimson color imparted to the whole. Let us estimate the half-gallon to contain 30,000 drops of water, and that to color a single drop so uniformly must require one hundred of the minute pieces into which the aniline-red is divided, and we must then infer that the small piece of the coloring matter has been broken into no less than *three millions of pieces.* Such an experiment illustrates the exceeding great divisibility of matter.

Ex. 15. Fill a cup to the brim with hot water: sprinkle into it slowly some finely ground loaf-sugar. Notice that the sugar disappears, and further that the cup is no fuller than before. Two or three tea-spoonfuls of sugar may be added before the water overflows.

Ex. 16. Take a cylinder, and fill it with alcohol to a height carefully marked with a rubber ring, or even a thread tied around it. Take cotton, and pick it out into fine shreds, and then introduce it, little by little, into the alcohol, carefully pressing it down to the bottom with a glass rod. A large quantity of cotton may be introduced without raising the level of the liquid.

These two experiments can be explained only by supposing that the minute parts of the water or the alcohol are not in actual contact, and that the minute parts of the sugar or cotton enter into the spaces between them.

Such experiments as these fall far short of giving a demonstration of the existence of molecules: nevertheless they are useful, since they pave the way to a clearer conception of the molecule and of molecular spaces.

35–41. Chemical Nomenclature.—A useful exercise *to precede* the study of the nomenclature may be conducted as follows: —

Ex. 17. Place upon the table, in a promiscuous group, bottles containing acids, bases, and neutral substances, two, three, or more of each. Provide also two cylinders, a solution of blue litmus, and a pitcher of water.

Nearly fill the cylinders with water, and add litmus enough to give a distinct blue color to both.

Then proceed to test the effect of each one of the substances by adding a few drops to the litmus. Those which redden the blue color, place together in a group at one side; those which restore the blue color of reddened litmus, place in a group at the other side; and those which have no effect upon either the blue or the red litmus, place together in a third group.

The classification of substances as acids, bases, and neutral bodies, is, in this way, clearly illustrated.

If you have selected binary compounds, as, for example, water, potassium iodide, mercuric chloride, to represent neutral bodies, you can easily make these an introduction to the study of paragraph 36, by writing their formulas on the blackboard, to show the meaning of the term *binary*.

They will be useful again, when you reach the topic "Other Binary Compounds," to illustrate the nomenclature.

39. Salts.—Ex. 18. Into a beaker put some dilute sulphuric acid, and drop into this some fragments of zinc. Violent effervescence soon begins, atoms of zinc taking the place of atoms of hydrogen; and this may be kept up by adding fresh pieces of the metal until the acid is exhausted.

Ex. 19. Filter the fluid just obtained, while hot, and then allow it to stand quietly until cold. A mass of the white crystalline salt, *zinc sulphate*, will be obtained.

Ex. 20. Repeat the experiments, using iron in the form of small nails instead of zinc, and the green crystalline salt, *ferrous sulphate*, will be produced.

Ex. 21. A small piece of sodium may be dropped upon the surface of dilute sulphuric acid. The salt *hydro-sodium sulphate* will be formed.

36–41. A very useful exercise to follow the study of the nomenclature may be conducted as follows: Place upon the table bottles containing substances representing all classes whose names have been studied. Their labels should contain their names only. Let a student be given a bottle, and let him, after reading the name of the substance in it, be asked to state, to what *class* of compounds it belongs. If a binary, then what are its constituents, and in what proportions, and what is its formula?

If an acid, then of which class, and what are its constituents?

If a base, then what are its constituents?

If a salt, then from what acid may we consider it to be derived? By what substitution? What are its constituents?

50. Hydrogen. — In experimenting with hydrogen the greatest care should be taken to have the gas unmixed with air, otherwise unexpected explosions may occur. Before collecting the gas in receivers, all air should be driven out of the apparatus. To know whether the gas is coming off free from air, it may be tested by collecting a test-tube full, and, keeping its mouth downward, apply the flame of a match. If the gas continues to burn quietly after the first slight explosion, it is pure enough. If the explosion is sharper, and no flame at all survives it, the gas is dangerous.

51. Reaction with Iron. — Ex. 22. Put a few small nails into a test-glass, and cover them with dilute sulphuric acid, made by mixing one part of the acid with three parts of water. Cover the glass with a glass plate. Light a match after the effervescence has gone on rapidly half a minute, and, removing the glass cover, bring the flame to the mouth of the glass. Notice the slight explosion which follows. Perhaps the flame will continue to play above the foaming liquid in the glass. Hydrogen is evolved by the chemical action as follows: —

$$Fe'' + H_2 S O_4 = Fe S O_4 + H_2.$$

The *bivalent atom* of iron takes the place of two atoms of hydrogen in the molecule of acid.

52. Lightness of Hydrogen. — Ex. 23. Fill a small jar with hydrogen. Lift it from the cistern of water, and turn its mouth upward. Carry it two or three paces away, and then bring a lighted taper into the jar: no hydrogen will be detected. Its quick escape from the jar shows it to be lighter than air.

Ex. 24. Lift another jar full of hydrogen from the cistern, and keep its mouth downward while it is carried away in the same manner. Bring a lighted taper to its mouth: the hydrogen takes fire, showing that the gas has remained in this inverted jar, proving again that it is lighter than air.

Ex. 25. **Hydrogen Soap-Bubbles.**—Remove the delivery tube of the hydrogen-generator (Fig. 23) from the cistern, and attach it to the stem of a tobacco-pipe by rubber tubing. Having the soap solution in a capsule, dip the bowl of the pipe into it in the usual way, and let the gas, as it comes over from the bottle, blow the bubble. While the bubble is still small, turn the mouth of the pipe upward. The bubble, having attained a diameter of three or four inches, will break away, or else it may be easily detached by a sudden movement of the pipe downward. It will then rise rapidly.

53. Explosive Bubbles.—Ex. 26. Cover the bottom of a dinner-plate with the bubble solution. Place the mouth of the pipe in the solution, slowly moving it from place to place, until a number of small bubbles rest upon the surface. Remove the pipe, and shortly afterward touch the bubbles with the flame of a match, which for this purpose may be tied upon the end of a long wire handle.

Air passes through the thin film, and mixes with the gas to form the explosive mixture.

Ex. 27. The explosive character of such a mixture may also be shown as follows:—

Into a wide-mouthed bottle put some fragments of zinc, and just cover them with water. Select a thin cork, which fits the bottle not too tightly, and make a hole through it large enough to admit a match easily. Take a long wire, and bend it near one end at right angles to itself. Bind a match upon this short bend of the wire. All this having been arranged, pour enough sulphuric acid into the bottle to liberate hydrogen with some rapidity. Insert the perforated cork, and wait about half a minute. Then fire the match, and by means of its wire handle insert its flame into the perforation of the cork. A violent but harmless explosion instantly follows, ejecting the cork from the bottle.

Ex. 28. Put a quantity of the soap-bubble solution into the hand, held slightly cup-shaped to retain it. Let the gas from the hydrogen bottle blow a bubble on this solution, as it did on the solution in the plate, Ex. 26. Remove the pipe, and touch the bubble with a match or taper flame, held in readiness by an assistant. An explosion follows, while the hand on which it occurs scarcely feels the slightest jar. When occurring in open air the explosion expends its violence in sound; when confined, the exploding mixture shatters the vessel which holds it, as would a charge of gunpowder.

57. Chlorine. — Chlorine is such a peculiarly suffocating gas, that great care should be taken to prevent its escape into the room. All joints in the apparatus should be perfectly air-tight. The receivers should not be filled quite to the brim with the gas. In changing the delivery-tube from one vessel to another, the transfer should be made as quickly as possible. The vessels of gas should be kept covered. Bottles of white glass are very good chlorine receivers, because they can be tightly closed with a cork until the gas is to be used.

If a little strong sulphuric acid is put into the bottle, so that the gas must bubble through it, the chlorine will be dried thereby: this is a point of importance in some experiments.

Ex. 29. Chlorine may be easily prepared, and in sufficient purity for some purposes, by covering the bottom of a jar or bottle with bleaching-powder, and adding a little sulphuric acid.

58. Solubility of Chlorine. — Ex. 30. Provide a bottle with a perforated cork, through which passes a short piece of glass tube, or even a piece of pipe-stem. Fill the bottle about one-third full of water, and the remaining two-thirds with chlorine from the apparatus in Fig. 29. Close the bottle with its cork tightly. Cover the end of the tube with the finger closely, then shake the whole violently for a few moments, and afterwards insert the tube in a vessel of water, and remove the finger. Water will be seen to rise into the bottle to supply the place of the gas which has been dissolved.

59. Chemical Properties. — Ex. 31. Tie a small tuft of cotton on the end of a wire, and wet it with ether. Set it on fire, and lower it into a jar of chlorine. The combustion will continue, with much smoke.

The chlorine combines with the hydrogen of the ether, but not with its carbon, which is therefore set free.

Ex. 32. **Bleaching.** — Into a bottle of chlorine pour a little litmus solution: cork the bottle, and shake it; the color of the litmus is discharged.

Ex. 33. Black, or other colored ink, may be used instead of litmus solution.

Ex. 34. Moisten a piece of paper containing ordinary writing, such as a part of a letter, and hang it awhile in a jar of chlorine. The writing will disappear.

Ex. 35. Try the experiment with a piece of newspaper or other printed paper. Notice that the characters are not destroyed. The coloring matter of printer's-ink is carbon, for which chlorine has little attraction.

Ex. 36. Insert a rose, or other colored blossom, in a jar of moist chlorine: it will be speedily bleached.

62. Iodine. — Ex. 37. Put a few crystals of iodine into a clean and dry flask, and gently heat it: notice the fine violet vapor.

Ex. 38. **Solubility of Iodine.** — When the flask is sufficiently cool, pour water in, to fill it two-thirds full. Shake it vigorously, and notice the brownish-yellow color imparted to the water. Iodine is very slightly soluble in water; 1 part of iodine requires 7,000 parts of water.

Ex. 39. Pour off this solution (iodine-water), and put a little alcohol into the flask. Notice the deep reddish-brown color which the liquid quickly assumes. Iodine is very soluble in alcohol. A solution of iodine in alcohol is called "tincture of iodine." It is used for medicinal purposes.

Ex. 40. **Test for Iodine.** — Boil a little starch in a beaker of water, and pour some of the liquid into a jar of water. Then add a few drops of "iodine-water," and notice the fine deep-blue color produced. This result is a very delicate test for the presence of free iodine.

71. Oxygen. — Certain precautions should be observed in preparing oxygen (Fig. 36). The materials should be dry and well powdered. The heat should be applied gently at first, and afterward so regulated that a steady and not too rapid stream of gas will be evolved. It will often be necessary to withdraw the flame altogether, and restore it again when the stream of gas slackens.

The end of the delivery-tube must be taken out of the water before the gas wholly ceases to issue, else, as the flask cools, atmospheric pressure forces water over into the hot flask, which may be then blown to pieces by the sudden evolution of steam within.

The quantity of material needed may be estimated by remembering that about two gallons of gas will be obtained from an ounce of potassium chlorate.

When large quantities of oxygen are desired, metallic flasks must be used instead of glass flasks, which are too frail.

Ex. 41. Fold a piece of tough paper into a narrow band. Bend this band around the upper part of a hard glass test-tube, and grasp both branches of it between the thumb and finger. The tube is thus provided with a convenient handle.

Put about a quarter of an inch of finely powdered potassium chlorate into the tube, and apply heat. The chlorate will first melt, and afterward appear to boil. Then insert a lighted match into the mouth of the tube, and notice that the vigor of its combustion is increased. Let the charred end of the match, with a spark upon it, fall upon the fused mass at the bottom, and notice the brilliant deflagration which follows. The chlorate is decomposed by heat, yielding oxygen.

72. Ex. 42. **Oxygen heavier than Air.** — Having two small jars of oxygen, stand one of them with its open mouth upward, and leave it uncovered; and the other with its open mouth downward, and with its edges resting on blocks to support it above the table. Then test the jars by bringing a lighted taper into each. It will be found that the taper burns more brilliantly in that which has been standing mouth upward, but is not affected by the other. Evidently the oxygen has fallen out of the latter. Both jars, equally, show that this gas is heavier than air.

73. Brilliant Combustion. — Ex. 43. Take a piece of crayon, and provide for it a long wire handle, so that it may be lowered into a jar of gas. This may be done by winding one end of the wire two or three times spirally around the crayon. Scoop out the upper end of the crayon, making a little cup. Into this cup put a piece of camphor. Set fire to the camphor, and quickly lower it into a jar of oxygen: it will burn with intense white light.

Ex. 44. **Combustion of Zinc.** — Cut a long and narrow strip from a sheet of zinc, and coil it by wrapping it spirally around a lead-pencil. Remove the pencil, and wind one end of the strip of zinc with thread, and immerse this in melted sulphur. Next set fire to this sulphur, and thrust it down into a jar of oxygen. The burning sulphur will heat the zinc, which will quickly take fire, and burn brilliantly.

75. Ozone. — To prepare the test-paper for ozone, take two hundred cubic centimeters of water, and add one gram of potassium iodide. When solution is complete, add ten grams of finely powdered starch, and heat gently until the fluid is thickened by the starch. Let narrow strips of paper be drawn through this mixture. The paper may be dried, and kept for a long time in closely stoppered bottles; but it must be moistened when used.

Heat the glass rod quite too hot to be handled.

When phosphorus is used, for the preparation of ozone, let the stick be first held under water while its surface is gently scraped: this cleaning of its surface is necessary if the phosphorus has been long exposed to light, and is thereby covered with red coating.

80. Water as a Solvent. — Ex. 45. Fill a test-tube, or small beaker, about half full of water; sprinkle into it a small quantity of finely powdered copper sulphate, and shake or stir it vigorously. The blue salt will impart its own color to the water; and, if too much has not been added, the solid will wholly disappear. A blue transparent liquid remains: this is a *solution* of copper sulphate.

Ex. 46. To the solution just made, if it be quite transparent, add a little more of the salt, and again agitate it. Repeat this operation, if necessary, until the last portion added remains undissolved. The clear blue liquid is then a *saturated* cold solution of copper sulphate.

Ex. 47. Apply the heat of a lamp-flame to the vessel containing the saturated cold solution. It will be found that the undissolved sulphate in the bottom will disappear. Then add more: in a little time that too will be dissolved, showing that heat increases the solubility of copper sulphate.

Ex. 48. Continue to add the powdered sulphate to the hot liquid until it, at length, refuses to dissolve. The quantity to be added may be surprising; but finally the liquid will take no more, and then it is said to be a saturated hot solution. Let it stand until cold.

82. Filtration.— Ex. 49. Into a small beaker of water put enough of the copper sulphate solution, obtained above, to give a slight, but distinct, blue tint; and then add some solution of potassium ferrocyanide, as long as the precipitate is increased by the addition. Notice the abundant dark-brown precipitate suspended in the liquid. Treat this fluid as described in paragraph 82, Fig. 44. A clear and colorless liquid falls into the bottle.

Ex. 50. **Distillation.** — A small quantity of water may be distilled in the apparatus represented in Fig. 50. Let water be poured in until the retort is nearly half full, and then heat it to boiling. Keep the flask, into which the beak of the retort is thrust, cold by a stream of cold water, as represented, or by pouring cold water from a pitcher. The distilled water will collect in the flask.

91. Effects of Heat on Sulphur. — The statements made in the text may be very easily verified in the manner described. Apply the heat carefully, and raise the temperature gradually.

92. Crystals by Solution. — Ex. 51. If the hot saturated solution of copper sulphate be examined when cold, a mass of blue crystals will be found. The clear blue liquid which remains, is a cold saturated solution, and the crystals are the excess which hot water can dissolve. If merely warm water instead of boiling water be used, this excess will be smaller, and the crop of crystals, on cooling, will be also smaller. In this case they will be more distinctly defined.

Ex. 52. Make a warm saturated solution of mercuric chloride. On cooling, it will deposit a crop of needle-shaped crystals.

Ex. 53. Make a warm saturated solution of potassium nitrate. On cooling, fine prismatic crystals will appear.

Ex. 54. Dissolve a little copper chloride in alcohol, making a saturated solution. Pour this solution over a *perfectly clean* glass plate, and let the excess run off, leaving only a film upon the surface. In a few moments a beautiful crystallization will begin, and spread rapidly over the plate.

Ex. 55. Use a saturated solution of ammonium chloride in water, to flow over the plate. Then *gently* warm the glass over the lamp, and afterward watch the growth of crystals.

93. Sulphur and Iron. — Ex. 56. Mix very intimately four grams of sulphur with seven grams of the finest iron-filings, and put the mixture into an ignition-tube, that is, a test-tube made of hard glass. Then apply heat to the lower end of the tube. Shortly the mixture will begin to glow, on account of the chemical action between the sulphur and the iron. After the action has well started, withdraw the tube from the flame, and the ignition will continue until the contents have been completely changed into iron sulphide.

98. Precipitates by Sulphuretted Hydrogen. — Ex. 57. Into one test-glass put a dilute solution of copper sulphate, into a second put a dilute solution of arsenious oxide, and into a third put a dilute solution of zinc sulphate. Add a few drops of hydrochloric acid to the first two solutions, and some ammonia to the third. Finally, add to each some of the solution of sulphuretted hydrogen, made in the way represented by Fig. 47.

In the first will appear a *black* precipitate of copper sulphide,
In the second " " *yellow* " arsenious sulphide,
In the third " " *white* " zinc sulphide.

103. Nitric Acid an Oxidizing Agent. — Ex. 58. Take a piece of tin-foil, about two inches square, fold it loosely, and place it in a porcelain cup or small beaker. Then add a small quantity of nitric acid. Very soon a violent action will begin, resulting in volumes of red fumes, which escape into the air, and a moist white powder in the dish. This white powder is the tin oxide into which the tin has been converted by the acid.

105. Sulphuric Acid and Water. — In mixing strong sulphuric acid and water, always pour the acid into the water, never the water into the acid. The fluid should be constantly stirred while the acid is being added.

111. Preparation of Nitrogen. — Ex. 59. Let the top of a cork be made slightly concave with a sharp knife, and then let it be thoroughly rubbed with powdered crayon or chalk. Cut from the end of a stick of phosphorus under water a piece as large as a small pea, and dry it completely by very gently pressing it for a moment between the folds of a piece of blotting or filter paper. Place the prepared cork on the water over the shelf of the cistern, and lay the phosphorus upon it. Next touch the phosphorus with a hot wire, and immediately invert over it a gallon jar. See Fig. 49, and the text accompanying it.

It will be well to have a jar of the gas prepared beforehand, with which to illustrate the properties of the gas. This will obviate the necessity of waiting for the absorption of the phosphoric vapors.

Ex. 60. **Lightness of Nitrogen.** — Fill a small jar with the gas, slip a glass plate under its mouth, and lifting it from the water, keeping its mouth downward, thrust a lighted taper up into it. The flame will be extinguished.

The result not only shows that nitrogen will not support combustion, but also that it is lighter than air.

Ex. 61. Fill the jar again, and, slipping the glass plate under it as before, lift it from the water, place it mouth upward on the table, and remove the plate. Leisurely light a taper, and finally thrust it into the jar. The flame will not be extinguished this time, which shows again that nitrogen is lighter than air.

126. Nitrous Oxide. — Ex. 62. Having filled some small jars with nitrous oxide, by the experiment described in connection with Fig. 51, remove one to the table, on which let it stand mouth upward. Slide its cover to one side, and insert a lighted taper. The flame instantly enlarges, continues with vigor, and is surrounded by a hazy envelope.

Ex. 63. Lower into another jar of the gas a combustion-spoon, containing a bit of ignited phosphorus, and notice that the combustion will proceed with almost as great brilliancy as in oxygen.

127. Nitric Oxide. — Ex. 64. Remove a small jar of this gas to the table, and plunge a lighted taper into it: the flame will be extinguished.

Ex. 65. Place a bit of phosphorus in a combustion-spoon, and ignite it. When the combustion is well started, plunge the phosphorus into a jar of nitric oxide: the phosphorus will continue to burn with great brilliancy.

129. Nitric Peroxide. — Ex. 66. Let a jar be about half filled with nitric oxide, and place it so that its mouth projects over the edge of the shelf on which it stands in the cistern. Pour air from another small jar up into it, and notice the cherry-red vapor which instantly appears. The greater part of this red vapor is nitric peroxide.

Or the jar containing the nitric oxide may be simply lifted a little to let air bubble under one edge into it. The red vapors will increase with every bubble.

137. Diffusion of Liquids. — Ex. 67. Make the experiment represented by Fig. 43, and mark the height of the vial in the jar. Then leave the whole standing quietly for twenty-four hours, and notice that the vial stands higher. The salt water has diffused upward, lifting the vial with it.

Ex. 68. **Osmose of Liquids.** — Fill a wide-mouthed vial with a strong solution of potassium chromate, and tie a piece of bladder over it, so as to close it completely. The vial should be full, and the bladder in contact with the fluid; but not the least portion of the solution should be on the outside. Stand this bottle on the bottom of a large beaker, and then nearly fill the beaker with water. Let the apparatus stand for twenty-four hours, when it will be found that the water is colored throughout, showing that the salt has passed through the membrane, to diffuse through the fluid outside.

Other salts may be treated in the same way, and their diffusion through the membrane compared.

141. Phosphorus. — Phosphorus should never be handled without the greatest care being taken to guard against igniting it. Handle it with forceps, and cut it when under water.

Ex. 69. **Phosphorescence.** — Place a clean stick of phosphorus in a dark room: it will emit a pale, pearl-like light. If the phosphorus is old, and covered with a red coating, it should first be immersed in water, and the red surface removed by scraping, to expose the translucent substance beneath.

Ex. 70. **Solubility in Ether.** — Put a little ether into a small vial, and add a few small pieces of phosphorus. Allow it to stand, shaking it occasionally, for several hours. Much or the whole of the phosphorus will dissolve.

Ex. 71. **Attraction for Oxygen.** — Expose a clean stick of phosphorus to the air, and notice the clouds of white vapor which fall away from it. Phosphorus and oxygen combine at ordinary temperature to form this phosphorus oxide.

Ex. 72. Saturate a strip of filter or blotting paper with the solution of phosphorus in ether, made in Ex. 70, and hang it conveniently exposed to air. The ether soon evaporates; and then the phosphorus combines with oxygen of the air, yielding clouds of white vapor, and sometimes evolving heat enough to set the paper on fire.

Ex. 73. **Fire in Water.** — Cover the bottom of an ale-glass or a test-glass with potassium chlorate, and add a small piece of phosphorus. Let water be introduced, enough to fill the glass two-thirds full. Next fill a pipette with strong sulphuric acid; close the top of it with the finger, and thrust the lower end into the water, and down upon the chlorate. Then remove the finger, and allow some of the heavy acid to flow. The acid decomposes the chlorate, and liberates chlorine peroxide. This is immediately decomposed by the phosphorus, which takes its oxygen, and enters into a brilliant combustion in the water.

153. Arsenic Compounds. — Ex. 74. Powder some arsenious oxide, As_2O_3, and heat it with water in a beaker. It is quite soluble in hot water. Saturate the hot solution, and then allow it to cool quietly. Crystals of arsenious oxide will be deposited, showing that this substance is less soluble in cold water.

Test the solution with litmus paper: it is an acid solution. The As_2O_3 is therefore an anhydride.

Ex. 75. **Arsenious Oxide is Volatile.** — Take a piece of glass tubing, about eight inches in length and three-sixteenths of an inch in diameter; hold its middle part in a gas-flame, turning it constantly to heat all sides equally, and when it softens pull it apart. Two short tubes, closed at one end, are thus obtained.

Into one of these tubes drop a very little arsenious oxide to the bottom, and apply a gentle heat. The arsenious oxide will soon leave the bottom of the tube, and afterward will be found, as a ring of white crystals, in the upper cold parts of the tube.

Ex. 76. **Scheele's Green.** — Put some of the arsenic solution into a test-glass containing water. Next prepare some ammonio-copper sulphate, as follows: To a dilute solution of copper sulphate add ammonium hydrate until the blue precipitate which is at first formed is again dissolved. The rich blue solution is the ammonio-copper sulphate. Add this solution, little by little, to the arsenical solution, and notice that a fine green precipitate is produced. This is Scheele's green, or copper arsenite.

Ex. 77. **Reinsch's Test.** — Add some of the arsenical solution to a test-glass of water, and add also a drop or two of hydrochloric acid. Next insert a piece of bright copper wire. In a little while the copper will be found to be covered with a dark gray coating. The copper decomposes the arsenical compound, and arsenic itself is deposited upon it.

166. Carbon Oxide. — The preparation of this gas may be accomplished with an apparatus fitted up as shown in Fig. 29, except that the delivery-tube should reach over to the water-cistern, so that the gas may be collected over water.

Ex. 78. Put about one hundred and eighty grains of finely powdered potassium ferrocyanide into the flask, and add about ten times this weight of strong sulphuric acid. Then cork the flask, and apply heat. The gas will be given off abundantly; and, after the air has been driven out of the apparatus, several small jars or wide-mouthed bottles may be filled with it.

Ex. 79. **Its blue Flame.** — Lift a jar from the cistern, keeping its mouth downward, and thrust up into it the flame of a taper. The taper will be extinguished on entering the gas, but the gas will itself take fire, and burn with a blue flame.

Ex. 80. **Lighter than Air.** — Lift another jar of the gas, and turn its mouth upward. Then introduce the flame of the taper: no combustion of gas follows, showing that it has escaped upward.

167. Carbon Dioxide. — There is a more simple method of obtaining carbon dioxide than that described in the text, which will be good enough for some purposes. Thus: —

Ex. 81. — Cover the bottom of a jar, or wide-mouthed bottle, with sodium carbonate, or with small fragments of marble, and pour a little hydrochloric acid upon this material. Vigorous effervescence occurs: carbon dioxide is set free. It lifts the air, and finally itself fills the jar.

Ex. 82. **It extinguishes Flame.** — Insert into the jar containing the gas the flame of a taper: it is instantly extinguished.

185. Organic Substances. — Ex. 83. Take a piece of common paper, and roll it into the form of a compact ball the size of a small nut. Drop this to the bottom of a test-tube. and heat it in the lamp-flame. The paper will give off white vapors, and a brownish liquid will condense on the upper and cold parts of the tube; but notice particularly that the paper becomes black. When no more vapors arise, remove the black mass from the tube, and notice that it is brittle, and in all other respects shows the properties of *carbon.*

The experiment proves that carbon is a constituent of this organic substance.

Ex. 84. Put about two inches of thick sirup of sugar into a cylinder, and stir into it a nearly equal volume of concentrated sulphuric acid, or "oil of vitriol." The acid decomposes the sugar, removes its hydrogen and oxygen, but leaves its carbon. Notice the bulky, coal-black residue. Carbon is a constituent of this organic substance also.

Ex. 85. Burn a bit of camphor, and hold a white plate in its flame. Notice the large deposit of carbon in the form of soot. Camphor is an organic substance, and carbon is one of its constituents.

187. Marsh-Gas.—Ex. 86. Marsh-gas may be prepared with the apparatus represented in Fig. 5. The ignition-tube should be about eight inches long. It should be charged with as much as needed of a dry powder, made by mixing two grams sodium acetate, four grams caustic soda, and eight grams slaked lime, and gently heating on a plate until the water of crystallization of the acetate is wholly driven off. The tube is then to be heated. The gas will be collected in the first flask.

$$Na\,C_2\,H_3\,O_2 + Na\,H\,O = Na_2\,C\,O_3 + C\,H_4.$$

The lime is used to render the mixture porous, and prevent its fusion.

Ex. 87. If a larger quantity of the gas is desired, the experiment may be made with the apparatus shown in Fig. 48. A larger quantity of the mixture of acetate, hydrate, and lime, is placed in the retort, and the gas is collected over water in the cistern.

When experiments are to be made with the gas, this method of preparation is better than that given above. Several small bottles may be filled for examination.

Ex. 88. Lift a bottle of marsh-gas from the cistern, keeping its mouth downward, and bring a lighted taper beneath it. The gas takes fire, and burns with a yellow flame.

Ex. 89. Lift a bottle of the gas from the water, turn its mouth upward, and, a moment or two afterward, insert a lighted taper. No combustion of the gas follows. The gas has escaped, showing its lightness.

Ex. 90. Lift a bottle of the gas from the cistern, and, holding a lighted taper in the other hand, turn the bottle mouth upward just under the flame. The gas rising out of the jar comes in contact with the flame, and burns.

191. Alcohol. — Ex. 91. Dissolve ten grams of honey in a liter of water, and add a little brewer's-yeast. Fill a small flask with the solution. Close the flask, and invert it in a dish containing enough of the same sirup to cover the mouth of it. Open it, and leave it standing in a warm place for twenty-four hours. Notice then that a colorless gas has collected in the upper part of the flask.

Ex. 92. Cork the flask while its mouth is still under the surface of the sirup: then lift it away, and place it upright on the table. The gas will now be in the neck of the flask. Remove the cork carefully, and insert the flame of a match or taper: it will be instantly extinguished. This, together with the fact that it is heavier than air, as shown by its remaining in the open flask, shows that it is *carbon dioxide.*

Ex. 93. Taste the fluid in the flask: it will be found to have the flavor of *alcohol.*

Ex. 94. Close the flask with a cork provided with a delivery-tube, such as may be seen represented in Fig. 59; but instead of letting the end of the tube dip into water, as there shown, let it pass into a second small flask resting on the water. Then apply heat, as in the figure, and keep the second flask cold by means of water. At a temperature of about 90° C. the dilute alcohol will distill over, and be condensed in the second flask.

194. Ether. — Ex. 95. Upon a plate first pour water enough to well cover the bottom, and then carefully pour ether upon one edge of the liquid, *letting no drops fall outside the plate.* The ether will spread over the surface of the water. Next touch a match-flame to the surface at one edge, and notice the instantaneous ignition. The flames will continue until the ether is all burned away. This shows very prettily that ether is lighter than water, and very combustible.

Ex. 96. Pour two or three cubic centimeters of ether into a wide-mouthed bottle or tumbler, and cover it loosely. After a few moments remove the cover, and apply the flame of a match. The vapors of ether will take fire, and burn with a flash. Hence ether is very volatile, and its vapors are heavier than air.

Ex. 97. Pour a few drops of ether on the bulb of apparatus (Fig. 1), and notice the rise of the liquid in the tube below. The rapid evaporation of the ether absorbs heat, and cools the bulb and the air within.

196. Olefiant Gas. — Ex. 98. For the preparation of this gas, the apparatus shown in Fig. 36 may be used. The flask should be large: one holding a liter may be used with the following quantity of materials, but in any case it should not be more than one-third full. Into the flask put fifty cubic centimeters of alcohol, and add two hundred cubic centimeters of strong sulphuric acid. Shake the mixture well, and then place the cork in the neck of the flask, and apply a gentle heat. The chemical action quickly begins. Ether is set free at first, but ethylene soon takes its place. Hence the first portions of gas may be allowed to escape before the receiver is brought over the end of the delivery-tube to catch the ethylene.

Watch and regulate the heat, so that the liquid may not froth up too much in the flask, and do not keep the operation going too long, since sulphurous acid is produced in the last stages of the reaction.

Ex. 99. Lift a jar of the gas from the cistern, mouth downward, and thrust a lighted taper up into it: the gas takes fire at the mouth of the jar, and burns with a bright flame. Turn the jar mouth upward while the gas is burning, and the flame will be fed by the gas escaping upward. It is a little lighter than air (.978).

227. Flame Tests for Potassium and Sodium. — Ex. 100. Bend the end of a platinum wire into a loop, moisten it, and gather some powdered potassium compound, as potassium chloride, upon it, and then thrust it into the edge of a colorless gas-flame, or even of an alcohol-flame. Notice the fine *violet* color imparted.

Ex. 101. Clean the platinum wire, and use it again in the same way, with some compound of sodium, — sodium chloride for example, — and notice the rich *yellow* color imparted to the flame.

Ex. 102. Make a strong solution of sodium chloride in alcohol. Fix a small tuft of cotton on the end of a wire, and wet it with this solution. By touching a match-flame to the tuft, a fine, large flame will be produced, which shows the yellow color to good advantage.

233. Lime-Water. — Ex. 103. Into a bottle of clear water put a small quantity of lime, and shake it thoroughly. A little of the lime will dissolve. Pour the milky liquid upon a filter: the clear and colorless filtrate is *lime-water*.

235. Calcium Carbonate. — Ex. 104. Fill a test-glass one-fourth full of lime-water, and dilute it with an equal quantity of water. Pass carbon dioxide through this solution, by means of the apparatus, Fig. 69, by putting the end of the delivery-tube into the test-glass. Notice that the fluid soon becomes milky: this is owing to the production of calcium carbonate, which is insoluble in water, and appears as a white precipitate.

$$Ca\,(H\,O)_2 + C\,O_2 = Ca\,C\,O_3 + H_2\,O.$$

Ex. 105. **Solubility of the $Ca\,C\,O_3$.** — Continue the stream of carbon dioxide, and after a time notice that the whiteness of the liquid is diminished. Indeed, after a while the solution may become again quite clear. The carbonate which formed at first is dissolved.

Ex. 106. **To restore the Carbonate.** — Put this clear solution of the carbonate in a flask, and boil it for some time. Notice that it becomes turbid again. The carbon dioxide is driven away by the heat, and the water can no longer hold the carbonate in solution. Hence it re-appears.

Ex. 107. **A Class-Test for the Group.** — Prepare three test-glasses, one with a solution of calcium chloride, another with strontium chloride, and the third with barium chloride. Into each pour a little ammonium carbonate. Notice that a white precipitate at once forms. The compounds of these metals behave alike toward ammonium carbonate: they are converted into carbonates, which are precipitated.

Ex. 108. **The Flame-Tests.** — Prepare a strong solution of each, barium chloride, strontium chloride, and calcium chloride. Soak a piece of pumice in each, and then with a pair of forceps hold them, successively, in the colorless gas-flame. The barium compound yields a delicate green flame; the calcium compound, a pale rose-red; and the strontium compound, a fine crimson.

237. To precipitate Aluminium Hydrate. — Ex. 109. Prepare a solution of aluminium sulphate or of alum in a test-glass, and add ammonia, little by little. A gelatinous white precipitate falls: this is the aluminium hydrate, $Al_2\,(H\,O)_6$.

Ex. 110. **Carmine Lake.** — Boil a little cochineal in a small flask of water until the coloring matter is extracted. Filter this solution into a test-glass, or beaker. Add to this colored water about an equal volume of aluminium sulphate or alum. Finally add ammonia. A colored precipitate is at once produced, which, on settling, leaves the solution quite or nearly colorless.

The ammonia precipitates the aluminium hydrate as in the preceding experiment. But this hydrate has a strong attraction for the coloring matter of the cochineal, and carries it along with itself. The colored precipitate is called CARMINE LAKE.

Ex. 111. **Other Lakes.** — Almost any other organic coloring matter may be used instead of the cochineal. A colored precipitate will be produced: all such are called LAKES.

It is this power of aluminium hydrate to precipitate coloring matter, that renders aluminium sulphate and alum useful in the arts of dyeing and calico-printing.

240. Compounds of Iron. — Iron forms two large and important classes of compounds, called respectively the FERROUS compounds and the FERRIC compounds.

Ex. 112. **The Ferrous Hydrate.** — Boil a little water in a beaker, to expel its dissolved air, and then dissolve in it a few small crystals of ferrous sulphate (green vitriol). Select those which are green, without white spots. Into another beaker put a little solution of potassium hydrate, and boil it also, to expel the air which it holds in solution. Finally pour the two solutions together, and notice the production of a copious white or greenish-white precipitate: this is *ferrous hydrate*, $Fe(HO)_2$. When pure, it is white: in presence of air, it becomes green.

Ex. 113. **The Ferric Hydrate.** — Make a solution of ferrous sulphate in water, add a few drops of strong nitric acid, and then heat to the boiling-point. Finally add potassium hydrate. Instead of the whitish-green precipitate obtained before, a copious red-brown precipitate is produced: this is the *ferric hydrate*, $Fe_2 (H O)_6$.

The nitric acid changed the ferrous into the ferric sulphate.

Ex. 114. **Test for Ferrous Compounds.** — Select a perfect blue-green crystal of ferrous sulphate, and dissolve it in recently boiled water. Divide the solution into two parts. To one part add a little potassium ferrocyanide, and notice the *pale-blue precipitate* which appears.

To the other part add a little potassium ferricyanide, and notice the *fine dark-blue precipitate* of ferrous ferricyanide.

Pale-blue precipitates with ferrocyanide, and dark-blue precipitates with ferricyanide, are given by the *ferrous compounds.*

Ex. 115. **Test for Ferric Compounds.** — Add a few drops of nitric acid to a solution of ferrous sulphate, and heat it to boiling, in order to change it into ferric sulphate, and divide the solution into two parts.

Add to one part a little potassium ferrocyanide, and notice the rich blue precipitate of "Prussian blue" which instantly appears.

Add to the other part a little ferricyanide, and notice, that while the liquid becomes colored, yet no precipitate is produced.

Deep-blue precipitates with ferrocyanide, and no precipitates with ferricyanide, are given by the *ferric compounds.*

Ex. 116. **To dye Cloth Blue.** — Make a solution of ferric sulphate by adding a little nitric acid to ferrous sulphate, and heating the mixture to the boiling-point. In another beaker make a solution of potassium ferrocyanide. Dip a piece of white cotton cloth into the ferric sulphate, and afterward immerse it in the ferrocyanide. Prussian blue will be precipitated upon every fiber of the cloth, and color it permanently blue.

Prussian blue is largely used in the arts of dyeing and calico-printing.

Ex. 117. **To dye Cloth Black.** — Make a weak solution of tannic acid in water, in one beaker, and a solution of ferrous sulphate in another. Bring a little of the two together in a test-glass, and notice the dark-colored precipitate formed. This precipitate becomes black on exposure to air. It is the ferric tannate.

Ex. 118. Saturate a piece of cotton cloth in the solution of tannic acid, and let it dry. Immerse the dried cloth in the ferrous sulphate, and hang it up exposed to air. Ferric tannate will be precipitated upon the fibers of the cloth, and dye them black.

Ex. 119. **Tannic Acid in Tea.** — Let a few tea-leaves be boiled in a small quantity of water. Pour the clear solution into a test-glass, and add a few drops of ferrous sulphate. The liquid blackens, and on standing it will deposit a precipitate of ferric tannate.

In the same way one may detect the existence of tannic acid (tannin) in coffee, in the husk of the horse-chestnut, in oak-bark, or in sumach.

253. Compounds of Lead. — Ex. 120. Into a test-glass, containing strong nitric acid, put some clippings of metallic lead. A violent chemical action soon sets in, with the evolution of volumes of red vapors. The lead is converted into *lead nitrate*, which remains dissolved in the liquid.

Ex. 121. Add a considerable quantity of the lead, enough to use up all the acid, and put the test-glass outside the window, so that the fumes may not fill the room, and let the action go on as long as it will.

The liquid which may be poured off clear, or filtered if necessary, contains the nitrate, which may be used in other experiments. It will be well to make this experiment beforehand, so that the nitrate may be ready for use when wanted.

Ex. 122. **Lead Chloride.** — Into a beaker put a teaspoonful of the nitrate solution, and add one hundred cubic centimeters of water. Then add a little hydrochloric acid, as long as a precipitate forms. This white precipitate is the lead chloride.

Ex. 123. **Obtained in Crystals.**—Heat the beaker until its contents boil, and notice that the chloride dissolves to a clear solution. Then place the beaker aside, where it will be undisturbed, and let it cool. By and by examine it, and notice the needle-shaped crystals of chloride which have been deposited.

This compound is very soluble in hot water, slightly in cold water, and the excess crystallizes.

Ex. 124. **Lead Iodide.**—Into a beaker containing water put another small portion of the lead nitrate solution, and add to this, drop after drop of potassium iodide. Each drop produces an additional quantity of the rich yellow precipitate of LEAD IODIDE.

Ex. 125. **Obtained in Crystals.**—Add a few drops of hydrochloric acid to the contents of the beaker, and heat to boiling. The iodide dissolves. Stand the beaker in cold water, or let a stream of cold water run upon it to hasten its cooling, and watch the result. The iodide separates from the liquid; and, on holding the beaker up to the light, multitudes of brilliant crystalline scales will be seen reflecting all the colors of the rainbow.

Ex. 126. **Lead Chromate.**—Fill a tall cylinder three-fourths full of water, and add fifty cubic centimeters of lead nitrate. Add a solution of potassium chromate, little by little, and notice the fine yellow cloud-like masses of LEAD CHROMATE which roll downward towards the bottom of the jar.

Ex. 127. **Action of Lead on Water.**—Two days before this subject is reached, prepare the experiment as follows:—

Into each of two bottles put some clippings of lead: the surfaces should be bright. Fill one, two-thirds full of rain-water, and the other with spring-water. It is likely that the rain-water will be found to be turbid, and that the spring-water will remain clear.

Or a little ammonium nitrate may be purposely added to the rain-water, and a little potassium carbonate to the spring-water, by which these results are facilitated.

255. Compounds of Copper. — Ex. 128. Fill a dinner-plate nearly full of water, and in the middle of it stand a test-glass. Put some clippings of copper into the glass, and pour upon them moderately dilute nitric acid, a little more than enough to cover them. At once invert over the glass a large receiver (shown in Fig. 20). A violent action quickly begins. The air in the receiver becomes cherry-red. The fluid in the test-glass becomes intensely blue. The red fumes will, for the most part, be dissolved in the water, and thus be prevented from escaping into the room; and, if a little blue litmus is put into the water, another change of color will be witnessed, from blue to red.

The blue liquid in the test-glass is a solution of *copper nitrate.* Pour it off from the residue of copper into a small beaker.

Ex. 129. **Copper Hydrate.** — Add a few drops of this nitrate solution to a test-glass of water, and then, drop by drop, ammonia, and observe the pale-blue precipitate: it is copper hydrate.

But continue to add the ammonia, and the pale-blue precipitate will begin to disappear. It dissolves, a fine rich blue liquid being obtained.

Ex. 130. To a test-glass of water, first add some drops of solution of copper sulphate, often called "blue vitriol," and then carefully add ammonia. The same pale-blue hydrate is precipitated. Continue the addition of the ammonia, and the hydrate dissolves in the excess as before, yielding a transparent liquid with beautiful azure-blue color. The richness of this color may appear to better advantage by diluting the liquid. For this purpose pour it into a beaker, and add water until the light can be seen through the solution.

This is a very sensitive test for copper salts. A very minute quantity can be detected by this ammonio-copper sulphate solution.

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