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ELEMENTARY CHEMISTRY

BY

THE REV. A. J. HAAS, S.J.

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SAP. xi-21.



FOURTH EDITION.

ST. JOSEPH'S COLLEGE,
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—
1908.

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PREFACE TO SECOND EDITION.

The first edition of this little work has become exhausted much sooner than was expected. This is owing to the fact that, though originally intended for our own High-School students only, it has been adopted as a text-book in several other Schools and Colleges.

In this edition, some new matter has been introduced, without, however, disturbing the original plan; new illustrations, partly chosen from among the stock of MESSRS. VIEWEG, partly made to order by the same artists, have been interspersed in the text; and a series of additional questions has been added and arranged, as far as possible, in the same order as the subject-matter of the book.

March 1900.

A. J. H., S.J.



PREFACE TO THIRD EDITION.

Beyond a few verbal corrections and some additional notes, no changes have been introduced into the present edition. A distinction has, however, been made here and there between portions of greater and less importance, by means of a difference in type, and three new illustrations have been added.

February 1907.

A. J. H., S.J.

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ELEMENTARY CHEMISTRY.

PART I.

GENERAL NOTIONS.

✓ CHAPTER I.

Introductory—Various changes in nature—Physical and Chemical Phenomena—Chemistry defined—Compound and Elementary Bodies—Mechanical Mixture and Chemical Combination.

1. Introductory.—Chemistry is a science which has at all times excited the interest and curiosity of man. It is no matter for wonder, indeed, that it should stimulate his curiosity, as it is of such a mysterious nature—mysterious on account of its strange and wonderful effects and phenomena, the causes of which were, for so many centuries, either entirely ignored, or wrongly interpreted; or that it should excite his interest, as there are so many precious advantages to be derived from it. Of course, now-a-days it is no longer with the vain hope of transmuting the baser metals into gold, that man devotes so much attention to the study of this science, but on account of the numerous and real benefits

which Chemistry confers upon mankind ; for medicine, hygiene, metallurgy, agriculture, as well as commerce and the fine arts themselves, all come in turn to call in its assistance and profit by its teaching.

2. Various kinds of changes in nature.—A very superficial observation of nature suffices to acquaint us with the fact that there are a vast number of phenomena or changes of every description going on around us daily. Thus, the sun rises and disappears again behind the horizon, light dissipates darkness, and cold succeeds heat. We see water flowing, clouds forming in the sky, and showers of rain pouring down. Animals and plants grow and decay ; fruits ripen, fall to the ground, and putrefy. Minerals heated with charcoal leave metallic iron or copper behind. A candle or a piece of wood burns away, giving off gases, vapours or smoke. The atmosphere of a crowded room becomes vitiated.

Now these and other numberless familiar changes in material substances may, for our present purpose, be roughly divided into two large classes—namely, *physical* and *chemical* changes. These form the subject-matter of the study of Physics and Chemistry.

The physical phenomena were known and their laws investigated at a comparatively early period, whilst the true nature of chemical phenomena and their laws did not come under notice till scarcely more than a hundred years ago. Hence Chemistry is sometimes called the younger sister of Physics.

3. Difference between physical and chemical changes.—Let us precise the difference between physical and chemical changes or phenomena by a few common laboratory examples.

Suppose we heat an iron rod. We observe that the rod expands and becomes appreciably longer; if we heat it sufficiently, it becomes white-hot or incandescent, and emits light and heat. On being allowed to cool down, it regains its original length; it becomes cold and dark again; in fine, it is the same iron rod as before. If now we fasten it by one end and make it vibrate, it will emit sound; if an electric current be made to pass around it, it acquires magnetic properties. These new properties disappear as soon as the vibrations stop or as soon as the electric current is interrupted, and the iron becomes exactly the same as it was before.

Again let us heat some sulphur in a glass flask. The sulphur melts and eventually turns into vapour. If we allow it to cool, it solidifies again. Thus the sulphur has passed from the solid into the liquid, and from the liquid into the gaseous state, and after cooling, has finally become the same sulphur with the same properties that it originally had.

A piece of lead when heated melts, and solidifies once more on cooling; a lump of sugar dissolves in water; when the water is evaporated, the solid sugar is again left behind.

In all these instances, the iron, the sulphur, the lead, and the sugar, undergo but temporary and

passing modifications, which do not change the nature of the substances: these modifications are called **Physical Changes** or **Physical Phenomena**.

But if we now leave the iron rod in moist air, it will soon be covered with a reddish-brown layer of iron rust. This iron rust is something new, altogether different from iron; it is no longer iron. If we heat the sulphur in the open air, it burns away with a blue flame, giving off a gas of a well-known suffocating odour, which will bleach coloured flowers when they are moistened, or will turn moistened blue litmus paper red and finally bleach it altogether. Nothing of the kind is observed when sulphur is heated in a glass flask. Moreover, the new gas formed does no longer condense into solid sulphur as previously. Hence the sulphur in the second case has undergone a lasting change; it has been transformed into something new, quite different from sulphur; it is no longer sulphur.

Similarly, if we keep molten lead for some time in the open air, it becomes covered all over with a yellow powder, which is something entirely different from the original lead; and if white sugar be placed on a piece of tin and heated over a lamp, gases are seen to be given off, and a black mass of a charcoal-like substance remains behind which indeed has none of the qualities of sugar. Here the iron, the sulphur, the lead, and the sugar, have undergone permanent modifications which change their nature entirely.

Such changes are called **Chemical Changes** or **Chemical Phenomena**.

4. Difference between Physics and Chemistry still more precised—Chemistry defined.—There is another important point of difference between Physics and Chemistry. Taking water as an example, the Physicist teaches us that water is a liquid, and shows how it can pass into the solid or into the gaseous state; he tells us its density, he describes its power of transmitting pressure, of finding its level, and so on. But he never inquires into the nature of the substance. The Chemist, on the other hand, not satisfied with what the Physicist teaches him about water, goes deeper and investigates the very nature and composition of water; he submits water to analysis; he splits it up into its elementary constituents, and to our great surprise, tells us that this well-known inert liquid consists of two gases, oxygen and hydrogen, chemically combined, and moreover, that for every given proportion of one gas, there is an invariably fixed proportion of the other.

From what has been stated it follows that *Chemistry is the science which teaches us of what substances bodies are composed, in what proportions their constituents are combined, and what are the causes which produce in bodies complete and lasting changes.*

5. Elementary and compound bodies.—The study of the composition of bodies has led chemists to point out a fundamental distinction between the various

substances with which they have to deal. Some substances—for example, mercuric oxide, chalk, water—can be decomposed into two or more different new substances, whilst from others, nothing new can be got, nothing different from themselves. If we heat in

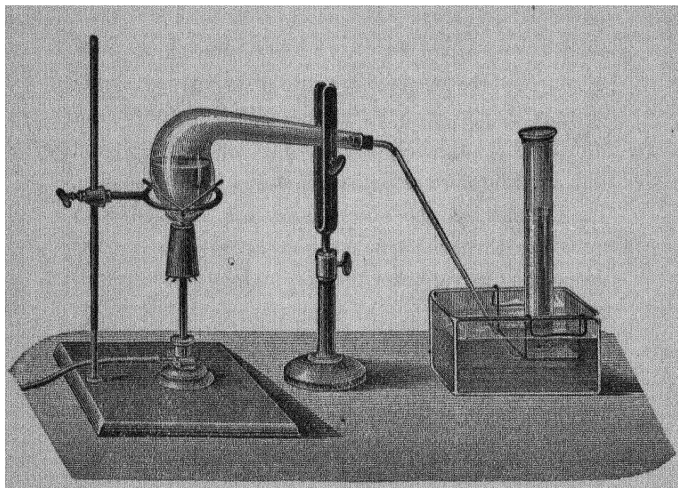


Fig. 1 —Decomposition of mercuric oxide by heat.

a glass flask or a glass retort (Fig. 1) some red oxide of mercury, we notice the escape of a colourless gas, which, on experiment, is found to be oxygen, whilst in the upper and cooler part of the flask or retort, there collects a shining deposit of a metallic substance which is known to be mercury. If we continue the experiment sufficiently long, all the red powder will disappear. We conclude, therefore, that the original

red powder is made up or *composed* of two substances, oxygen and mercury. Mercury oxide is, therefore, said to be a *compound body*. Up to now it has been found impossible to split up in their turn either oxygen or mercury into anything *different* from oxygen and mercury: for this reason oxygen and mercury are called *simple* or *elementary* substances.

Chemists have never yet been able to resolve or break up sulphur into two or more dissimilar sub-

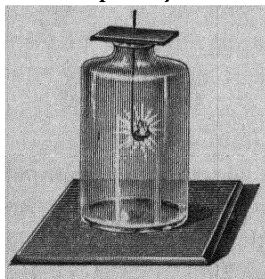


Fig. 2

Sulphur burning in oxygen.

stances, they have never obtained from it anything different from sulphur; sulphur is therefore said to be an *element*. If we burn sulphur in oxygen (Fig. 2), both the oxygen and the sulphur soon disappear entirely, and in their place there is left another substance, a combination of sulphur and oxygen, called sulphur dioxide;

this substance is a *compound body*.

Therefore a **simple** or **elementary body**, or an **element**, is a substance which it has never yet been possible to separate into two or more dissimilar substances or from which we cannot get more than one kind of matter. *Chemical Language*

A **compound body** is a substance out of which we can get two or more different kinds of matter, or which

results from the union of two or more substances into one single new substance.

6. Occurrence of the elements.—The elementary substances that exist in this world seem to be comparatively few in number. At present some 70 or more are known to chemists. Moreover many among these are very rare; others, though more abundant, do not seem to be so very useful. Thus gold and silver, valuable as they are in other respects, so far as we know at least, are not of great use in the economy of nature. Comparatively few metallic elements are practically useful for our daily life. These are iron, copper, tin, zinc, lead, and their alloys; a small number also of non-metallic elements are the chief constituents of all vegetable and animal bodies, as well as of all kinds of food. They are carbon, hydrogen, oxygen, and nitrogen. The elements of which the greater portion of the earth's crust, with all the living beings upon it, is built up, amount to about 12 or 15 only. Thus the solid portion, comprising stones and rocks, consists chiefly of carbonates and silicates (that is, compounds containing oxygen along with carbon or silicon respectively) of aluminium, iron, calcium, magnesium, sodium, and potassium.

The liquid portion, or water, consists of hydrogen and oxygen.

The gaseous atmosphere—of oxygen and nitrogen chiefly.

Plants—chiefly of carbon, hydrogen, and oxygen.

Animal bodies—chiefly of carbon, hydrogen, oxygen, and nitrogen, along with calcium phosphate or carbonate.

7. Difference between a chemical combination and a mechanical mixture.—(1) When two or more substances are simply mixed, no heat is evolved and no light perceived, whilst when they combine chemically, heat is evolved,* and sometimes light is perceived.—For example, when copper or iron filings are merely mixed up with flowers of sulphur, no heat is perceived and no light appears, whilst, when they are heated together, they combine with evolution of heat and light.

(2) The mixture does not acquire properties entirely different from the original substances, but these properties are intermediate; whilst when substances combine chemically, the resulting compound is something altogether different, having no resemblance at all with the combining substances.

Thus the properties of a mixture of iron and sulphur are intermediate between those of iron and sulphur taken separately; for the sulphur of the mixture still burns with a blue flame, and the iron of the mixture is still attracted by a magnet; but these two substances having chemically combined, the iron

* Though this is correct as a general rule, it is also true that in some cases of chemical combination heat is absorbed.

sulphide which is formed, bears not the least resemblance to sulphur, it no longer burns like it; nor is it any longer attracted by the magnet like iron.

(3) Whatever be the proportions of either one or the other of the substances in a mixture, it is always a mixture. But in a chemical combination, substances unite in fixed proportions by weight, *i.e.*, for a given weight of one substance, a given weight of another substance is required.—For example, whether we mix 1, 10 or 100 ounces of iron with 100, 10 or 1 ounce of sulphur, the resulting substance will always be a mixture; but when iron and sulphur combine chemically, for every 56 parts by weight of iron, always and invariably 32 parts by weight of sulphur are required to form 88 parts by weight of iron sulphide. If these proportions be not kept, and if there be any excess of either the one or the other, this excess will be left free.

(4) Mixed substances can, in many cases, be easily separated by physical means, whilst in a chemical combination it is not so. For example, the iron in our mixture of iron and sulphur may be separated from the sulphur by any of the following methods:—By gently blowing over the mixture, the lighter sulphur flies off; by passing a magnet over it, the iron alone is attracted; by throwing the mixture into water, the heavier iron sinks at once to the bottom, and the sulphur remains suspended in the water, which may be poured off; by treating the mixture

with carbon disulphide, the sulphur alone is dissolved and may be separated by filtering the liquid and allowing it to evaporate.

On the other hand, to separate substances that have combined chemically, more difficult, indirect and delicate means must be employed.

CHAPTER II.

Molecules—Atoms—Elementary and Compound Molecules—Chemical Affinity—Cohesion—Adhesion—Symbols and Formulæ—Atomic and Molecular Weights.

8. Molecules and Atoms.—As the edifice of chemical and physical science rests upon the assumption that there are atoms and molecules, or on the atomic theory as it is usually called, it is important from the very outset, to have a clear conception of what atoms and molecules are.

Let us then take a lump of the white substance commonly called chalk, and which chemists designate as calcium carbonate. Let us divide it into two pieces, and putting aside one, again divide in two the remaining half, and repeat this process till the piece left becomes too small for the naked eye conveniently to see it. Then calling in the aid of powerful microscopes, suppose we continue this operation. Assuredly a limit will finally be reached, beyond which further division by any physical means as yet known, will become impossible. The exceedingly small particle

of matter which is thus obtained is still chalk, with exactly the same properties as the original lump, just as a drop of water, however small, has still the same properties as the entire mass of water from which it has been separated.

We have not, however, yet done with our piece of chalk. We have merely exhausted all known physical means of further division. By the aid of the imagination, we can still go on dividing it till we come to the smallest conceivable particle of chalk, which accordingly can no longer be divided into *similar* particles even by the imagination. This last particle is what is called a *molecule** of chalk.

Now this molecule of chalk is in its turn made up of various elementary particles of matter, calcium, carbon and oxygen, so that we can break up the molecule into these elementary constituents; but then it ceases to be a molecule—it ceases to be chalk. It has been found impossible, on the other hand, to split up still further these elementary particles of calcium, carbon and oxygen, and because they cannot be thus *cut* up into anything simpler, they are called *atoms*.†

Knowing, therefore, the molecule of calcium carbonate to consist of one atom of calcium, one of carbon and three of oxygen, we represent it by the formula CaCO_3 ; so that the name of molecule is applied to the the entire group CaCO_3 . This molecule CaCO_3 consists

* (Lat. *molecula* = a small mass.)

† (Gr. *a* = not, *temno* = I cut.)

of atoms of calcium, of carbon and oxygen, but it would be wrong to speak of an atom of chalk; such a thing cannot exist.

Hence a **molecule** may be defined to be *the smallest particle of any substance which can exist as such in nature.*

Applying this definition to a particular substance, for example chalk, we say: a molecule of chalk is the smallest possible particle of chalk which can exist as chalk, and which we cannot further subdivide without destroying it as chalk, or, in other words, without changing its nature.

An **atom**, on the other hand, *is the smallest particle of matter which can take part in a chemical reaction, or the smallest particle of matter which can help to make up a molecule, but which cannot be split up into anything simpler than itself.*

9. Elementary and compound molecules.—We have seen that atoms, as a general rule, unite with atoms to form molecules. Now the atoms which combine with one another are either of the same kind or of different kinds. In the one case we have *elementary molecules*, in the other *compound molecules*.

For example, hydrogen atoms may combine with other hydrogen atoms, and oxygen atoms with other atoms of oxygen; the molecules thus formed will be respectively molecules of hydrogen and molecules of oxygen, and these are elementary molecules. On the

other hand, when hydrogen atoms combine with oxygen atoms, we get compound molecules of hydrogen oxide, or water. The molecule of calcium carbonate likewise is a compound molecule made up of the different atoms calcium, carbon, and oxygen.

N.B.—The word *atom* conveys the notion of an elementary body, and it is evident that there are as many different kinds of atoms as there are elements. The term *elementary body*, however, does not necessarily mean one atom only. An atom may, for our present purpose, be considered as the smallest possible particle of an elementary substance.

10. Chemical affinity—Cohesion—Adhesion.—

1°. So far we have stated that atoms unite with one another. There must consequently be some force that causes them thus to unite into molecules; to this force the name of **chemical affinity** has been given. Chemical affinity may then be defined as *that force which causes atoms to unite with one another into molecules, and which keeps them together when once united.*

2°. Molecules, in their turn, are kept together by some force, chiefly in a large class of substances called solids; and were it not for such a force, the existence of solid substances would be an utter impossibility. This force is known under the name of **cohesion**; it may be defined as *that force which keeps together the several molecules of the same substance.*

3°. To these two forces a third one may be added, viz., **adhesion**, or *that force which keeps together molecules of different substances when placed in contact with one another.*

Examples of these forces at work can easily be pointed out :

1°. It is chemical affinity that binds together atoms of calcium, carbon, and oxygen to form molecules of calcium carbonate ; it is chemical affinity that unites atoms of hydrogen and oxygen to form molecules of water. Similarly, it is owing to chemical affinity that the atoms of carbon and oxygen keep together in molecules of carbon dioxide gas.

2°. (a) On the other hand, it is cohesion that keeps together the many hundreds of thousands of little molecules of calcium carbonate which make up a bit of chalk ; and in this case, the force of cohesion between the molecules is so great, that these are unable to move freely one about the other: chalk exhibits what is called the *solid state*.

(b) Although the force of cohesion exists in water, still it is so weak, that its molecules can slide about each other with great freedom: water is an example of a body in the *liquid state*.

(c) Finally there is no cohesion at all between the molecules of oxygen and hydrogen, so that they move about independently of one another, and even tend to go as far away from one another as possible: such substances exhibit what is called the *gaseous state*.

3°. Lastly, when a piece of glass, for example, is dipped into water, some of the water adheres to the glass when lifted out; this is due to the force of adhesion.

With our present object in view, we may broadly assert that the work of the chemist consists chiefly in the study of atoms and of the way they act on one another in virtue of the force of chemical affinity; whereas the work of the physicist chiefly consists in the study of the forces which any way modify cohesion, as well as of some other forces which act on molecules at a greater or less distance.

11. Symbols.—It has been said that there are as many different kinds of atoms as there are different elements, since the atoms may be considered as the smallest, undividable particle of every elementary substance. Now, as each atom or element is different from the others, each must naturally be designated by a different name. But instead of repeatedly writing those names in full, chemists have agreed upon using conventional signs or *symbols* to represent the various elements. The symbols chosen are the first letter of their English or Latin name. When several names begin by the same letter, a second small letter, taken from the word itself, is added to distinguish them. Thus the letter O is used to designate the element oxygen, H stands for hydrogen, Cl for chlorine to distinguish it from C which represents carbon. Iron is called in Latin *ferrum*, and silver *argentum*; hence their symbols are Fe and Ag respectively. (See table of symbols, page 18.)

12. Atomic and combining weights.—Moreover, the atoms of the various elements differ from one

another in another respect. Each one not only has its own specific properties, but also its own weight. Now suppose we have here before us some 70 different stones and we are asked to compare their weights; we should naturally find out the lightest among them all and compare the weights of all the others with the weight of this. This is precisely what chemists have done with regard to the weight of the various atoms of the elements. Hydrogen was discovered to be the lightest of all the elements; consequently, the weight of one atom of hydrogen has been chosen as the standard of comparison, and it is represented by the number 1. By direct experiment the atom of oxygen has been found to be 16 times, the atom of carbon 12 times, that of calcium 40 times, and that of mercury 200 times, heavier than the atom of hydrogen. The weights of all the other elements have likewise been determined by direct or indirect methods. *This relative weight of any atom, compared with that of hydrogen taken as unit, is called its **atomic weight**, and it is the smallest quantity by weight of any simple substance, which can combine with other substances.* Hence the atomic weight of an element is sometimes called its **combining weight**.* Thus the combining weight of hydrogen is 1 and that of chlorine is 35.5, because when these two elements combine with one

* As it would be above the present standard to introduce the more precise meaning of combining or equivalent weights, we may for the moment consider the combining weights as *practically* the same as the atomic weights.

another to form hydrochloric acid, they always do so in the ratio of 1 part by weight of hydrogen to 35.5 of chlorine. The combining weight of oxygen is 16, because it is in this proportion that oxygen combines with hydrogen to form water, and in the same proportion it combines with 56 of iron to form iron oxide, and so with other elements.

The symbols of the elements not only stand for their names, but also for their atomic or combining weights. Thus O not only represents the element oxygen, but it also signifies one atom of oxygen and 16 parts by weight of oxygen.

TABLE

of the more important Elements with their
Symbols and Atomic Weights.

NAMES OF THE ELEMENTS.	SYMBOLS.	ATOMIC WEIGHTS.
ALUMINIUM	Al	27
Antimony (Stibium)	Sb	120
Arsenic...	As	75
Barium	Ba	137
Bismuth	Bi	208
Boron	B	11
Bromine	Br	80
Cadmium	Cd	112
CALCIUM...	Ca	40
CARBON	C	12
CHLORINE	Cl	35.5

NAMES OF THE ELEMENTS.	SYMBOLS.	ATOMIC WEIGHTS.
Chromium	Cr	52·1
Cobalt	Co	59
COPPER (Cuprum)	Cu	63
Fluorine	F	19
GOLD (Aurum)	Au	197
HYDROGEN	H	1
IODINE	I	127
IRON (Ferrum)	Fe	56
LEAD (Plumbum)	Pb	207
MAGNESIUM	Mg	24
Manganese	Mn	54
MERCURY (Hydrargyrum)	Hg	200
Nickel... ..	Ni	58·6
NITROGEN	N	14
OXYGEN	O	16
PHOSPHORUS	P	31
PLATINUM	Pt	194
POTASSIUM (Kalium)	K	39
SILICON	Si	28
SILVER (Argentum)	Ag	108
SODIUM (Natrium)	Na	23
Strontium	Sr	87·5
SULPHUR	S	32
TIN (Stannum)	Sn	118
ZINC	Zn	65

N.B.—The names of those elements with which we are more or less directly concerned here, are printed in capitals.

13. Molecular formulæ.—Since a molecule is made up of atoms, and each atom is represented by a symbol, to represent a molecule we have only to group together the various symbols of the atoms. Such a group of symbols is designated by the term *formula*. Thus the molecule of hydrochloric acid contains one atom of hydrogen and one atom of chlorine: we represent this by HCl, which is the molecular formula of hydrochloric acid. Water contains two atoms of hydrogen and one of oxygen; hence the formula of water is HHO, or more simply H_2O . A number placed immediately after the symbol or a group of symbols tells us how many such atoms or groups of atoms are present in the molecule; whilst a number placed before the formula in the form of a coefficient, indicates the number of molecules. Thus in $PbCl_2$ and $Pb(NO_3)_2$, the number 2 shows that the atom Cl and the group of atoms or *radical* NO_3 are to be taken twice. Again $2HCl$ stands for two molecules of hydrochloric acid.

14. Molecular weights.—Since each atom has its own atomic weight, the sum of all the weights of the various atoms in a molecule makes the *molecular weight*. Thus in the formula of water H_2O , H stands for hydrogen and represents the number 1, and O stands for oxygen and represents the number 16. Therefore H_2O represents $1+1+16$ or 18 parts by weight of water; the molecular weight of water is said to be 18. Similarly the molecular weight of carbon dioxide CO_2 is $12+2 \times 16$, *i.e.*, 44.

What then does a chemical formula teach us? Take for example CaCO_3 .

1° This formula represents calcium carbonate, *i.e.*, a particular compound of the elements calcium, carbon, and oxygen.

2° It represents *one* molecule of calcium carbonate or chalk.

3° It tells us that for every one atom of calcium in chalk there is one of carbon and three of oxygen.

4° That in every hundred parts by weight of chalk (molecular weight = $\text{CaCO}_3 = 100$) there are for

$$40 + 12 + 48$$

every 40 parts by weight of calcium, represented by the symbol Ca, 12 parts of carbon and 3×16 or 48 parts by weight of oxygen, represented by C and O_3 respectively.

CHAPTER III.

Chemical Action—its Cause, Conditions, Effects, Classification.

15. Chemical action or chemical changes.—It has already been said that it is the province of chemistry to investigate 1° the manner in which atoms are pulled asunder—*i.e.*, the manner in which the molecular edifice is destroyed, which happens whenever a simple decomposition takes place; 2° the manner in which atoms unite and build up a new molecule, as in a chemical combination; 3° the way in which atoms exchange their places and re-arrange

themselves into new and different molecules, as when two compounds are put in presence of one another, *i.e.*, whenever a double decomposition takes place. Any one of these three cases is a chemical action, or reaction, or a chemical change.

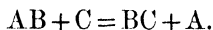
Now in a chemical reaction we may consider: 1° the *cause* that produces it, 2° the *conditions* under which it takes place, 3° the *effects* that accompany it, and 4° the form it assumes or its *classification*.

16. (1) Cause.—As for the cause of chemical changes, it has already been spoken of as that mysterious force which binds atoms together into molecules; it is a force which we know to exist from its effects, but the true nature of which is unknown to us.

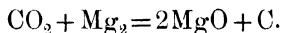
Because formerly chemists thought that chemical action was due to a certain *liking* of the elements for one another, or a certain similarity or *relationship* between them, they gave to this force the name of chemical affinity (Lat. *affinis*, related to). But it has been observed, on the contrary, that union takes place between different kinds of matter with different degrees of energy, and that, as a rule, the greater the difference between the substances, the more intense is this chemical action on each other; so that chemists prefer to the word *affinity* the word *attraction*, which, of course, does not tell us much more about the nature of this force. And because it acts only on the minutest particles of matter and at

extremely small distances, they call it *chemical attraction*, to distinguish it from *physical attraction* of gravitation which can act at very great distances.

Suppose we have a body AB, consisting of atoms A bound with atoms B, and another body C. If we put them in presence of one another, B will leave A and unite with C, to form a new and different compound BC. We explain this by saying that B has a greater affinity for C than for A :—



Experiment. Fill a test-tube with carbon dioxide gas, CO_2 , and introduce into it a piece of ignited magnesium wire. The magnesium continues to burn in the atmosphere of CO_2 combining with the oxygen and leaving free carbon behind, in the form of a black powder :—



We account for this reaction by saying that there is a greater affinity or attraction between magnesium and oxygen than between oxygen and carbon,—or that the cause of the new chemical action is the affinity of magnesium for oxygen, or reciprocally.

17. (2) Conditions.—Now, whatever this force of affinity may be in itself, it cannot act unless some favourable conditions be given.

(a) The first and most important is close **contact** ; and all the other conditions merely tend to produce this one. This fact again shows the characteristic difference between chemical and physical

forces, as these act at greater distances,* even from one end of the universe to the other, while the former can manifest their action at exceedingly small distances only *i.e.*, within the sphere of action of the atoms.

Exp. Pour into a glass some hydrochloric acid and

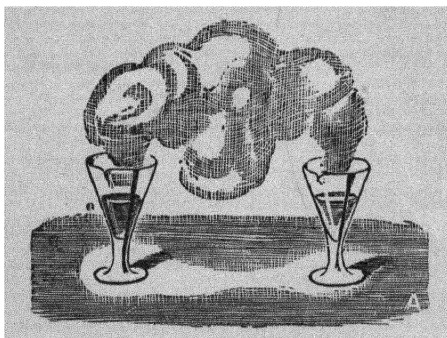


Fig. 3.—Ammonia and hydrochloric acid forming white fumes of ammonium chloride.

into another some ammonia, keeping the two glasses at some distance apart.—No action is perceived. Bring them now slowly near each other, and a dense, white cloud of ammonium chloride is formed (Fig. 3), as the two gases hydrochloric acid and ammonia come into contact.

* Cohesion which is a physical force can act, however, at very small distances only.

(b) One at least of the combining substances must be in the **liquid** or in the **gaseous** state in order to facilitate *contact*.

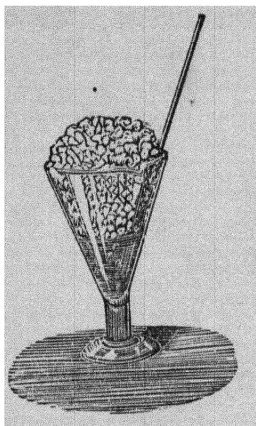


Fig. 4.—Effervescence produced by pouring tartaric acid upon sodium bicarbonate.

Exp. 1. Mix some powdered tartaric acid with dry sodium bicarbonate: no action is observed. Pour some water upon the mixture: an effervescence will indicate that the chemical action is taking place. If one of the powders be previously dissolved and the solution be poured upon the other, the effervescence will be greater still, and it will be greatest (Fig. 4) when both substances are separately dissolved before they are mixed, as then *contact* is closest and extends

throughout the mass.

Exp. 2. Shake some finely-powdered antimony into a glass-bottle filled with chlorine gas; chemical action manifests itself with incandescence (Fig. 5).

But here let us again remark that the antimony must be finely powdered in order to increase the area of *contact*.

(c) Many substances do not combine at the ordinary temperature, a certain degree of **heat**, called the *temperature of ignition*, must first be reached.

Thus a mixture of hydrogen and oxygen may be kept indefinitely, and no combination will take place between them; but they will combine instantaneously and with a loud report at the approach of a lighted taper.—Similarly, a candle, wood, magnesium, etc. will burn only when we heat a portion of their substance.—In most cases the heat produced by the chemical action itself is sufficient to cause the combustion to continue. The effect of the heat is to

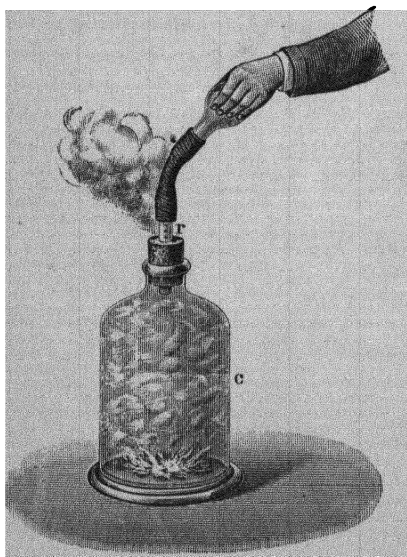


Fig. 5.—Antimony burning in contact with chlorine gas.

cause collision among the particles of bodies, and thus to bring them into closer contact.

(d) **Light and electricity** also occasion many chemical changes. Direct sunlight, for example, causes a mixture of hydrogen and chlorine to combine with explosion.—A current of electricity decomposes water; while an electric

spark causes a mixture of hydrogen and oxygen to combine and form water.

In all these cases there is an action in some respect similar to that of heat, in as much as it also favours closer *contact*.

• (e) To the foregoing cases we may add **mechanical agents**, such as pressure, grinding, a sudden shock, etc:

Thus by grinding somewhat smartly a mixture of potassium chlorate and sulphur (try with small quantities), we hear a series of decrepitations, and under a hard blow or when thrown against a stone, it explodes with violence. The explosion of gunpowder is explained in this way.

18. (3) Effects.—Chemical reactions are always accompanied by physical effects or phenomena: phenomena of heat, light, sound or electricity. When a chemical action is going on, heat is always evolved or absorbed, and there is sometimes produced along with it light, sound or electricity.

Heat always accompanies a direct chemical combination; for example, when water is poured upon quicklime, or when sulphuric acid is poured upon water, a match will take fire if made to touch the test-tube containing the mixture.

N.B.—In some cases of slow combustion (see art. combustion), the chemical action goes on so slowly, that the heat actually produced at any instant of time, is too weak to be directly perceived.

When carbon, sulphur, a thin iron, magnesium or zinc wire burns in oxygen, the phenomenon of heat and light invariably accompanies the chemical

combination.—When hydrogen and oxygen are made to combine, the chemical action is attended by heat, light and sound.—When a sheet of zinc is dipped into dilute sulphuric acid, the chemical action which takes place produces electricity. This may be rendered evident by connecting the extremity of the zinc outside the liquid with the end of a sheet of copper, dipping in the same acid. The current of electricity thus generated is sufficiently strong to ring an electric bell.

Phenomena of other descriptions, besides those just mentioned, accompany chemical changes, such as change of state, production of new colours, acquisition of toxic properties etc.

Thus two coloured solids, black charcoal and yellow sulphur, both of them non-poisonous substances, on combining produce a colourless, poisonous liquid, called carbon disulphide.

When we grind in a mortar solid sulphur with liquid metallic mercury, a black powder of mercuric sulphide is obtained, which on being heated becomes red and is then called *vermilion*.

Again a colourless solution of mercuric chloride produces with another colourless solution of potassium iodide a beautiful scarlet-red precipitate of mercuric iodide.

Two colourless gases, ammonia and hydrochloric acid, combine forming a white solid, ammonium chloride.

19. (4) Classification.—As has been said at the beginning of this chapter, a chemical action may be a case of simple decomposition, or of direct combination or of double decomposition.

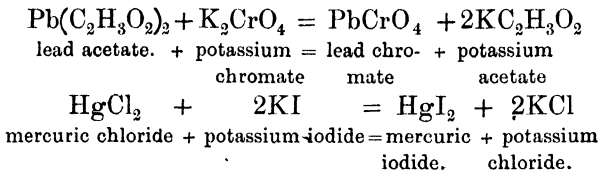
(a) **Simple decomposition** takes place when a compound is split up into its components or into some other simpler substances. Thus heat decomposes potassium chlorate into a simpler compound, potassium chloride, and into oxygen: also mercuric oxide into mercury and oxygen; whilst a current of electricity decomposes water into hydrogen and oxygen.

(b) There is **combination** when two bodies, simple or compound, unite according to certain fixed laws (see following chapter) to form a compound body possessing different properties.

For example, when iron combines with sulphur to form iron sulphide; when sulphur, carbon, hydrogen, magnesium etc., combine with oxygen to form their respective oxides; or when carbon dioxide combines with calcium hydroxide to form calcium carbonate.

(c) There is **double decomposition** when the two previous cases occur at the same time, i.e., when two compound substances, capable of reacting upon each other, are put in presence of one another. Each compound is first decomposed, and then the atoms, by a new combination, re-arrange themselves in a different way, giving rise to two other new

compounds, as when lead acetate reacts upon potassium chromate, or mercuric chloride upon potassium iodide.



CHAPTER IV.

Chemical Laws—Laws of Definite and Multiple Proportions—Chemical action takes place readily between unlike substances—Principle of Conservation of Matter—Chemical Equations.

20. Atomic theory.—A few facts have now been stated about the constitution of matter, *i.e.*, how matter ultimately consists of molecules, and these of atoms. We have seen what atoms are, what are their relative weights, what force binds them together into molecules. It now remains for us to add, that when atoms enter into combination, they do so not at haphazard, but in obedience to some well-determined chemical laws. All these notions together constitute what is called the **atomic theory**.

It is not within our scope to study here all the known laws that regulate chemical combinations. We must be satisfied with the statement of the most important among them; *the Law of Definite Proportions*, and *the Law of Multiple Proportions*. To these will be added some facts which are generally considered as having the importance of laws.

21. Law of Definite or Fixed proportions.—*In any given compound the weights of the elements which compose it, are invariably fixed, in whatever manner the combination be brought about.*

Ex. 1. Thus in whatever part of the world we make hydrogen combine with oxygen, we find that two parts by weight of hydrogen invariably unite with 16 parts of oxygen, to form 18 parts of water; and that any excess of either the one or the other gas will be left free or uncombined.

Similarly, wherever and in whatsoever manner we analyse water, we invariably obtain the two constituents of water in the ratio of 1 part by weight of hydrogen to 8 parts of oxygen, or in the ratio of 2 parts of hydrogen to 16 of oxygen in every 18 parts by weight of water.

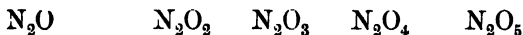
Ex. 2. In order to prepare hydrochloric acid, we must necessarily take for 1 part by weight of hydrogen, 35·5 parts of chlorine; and when we decompose hydrochloric acid, for every 36·5 parts of it by weight, we invariably obtain 1 of hydrogen and 35·5 of chlorine.

This is evident in the light of the atomic theory, which tells us that the molecule of water H_2O contains 1 atom of oxygen with its unchanging atomic weight 16, and two atoms of hydrogen each with its atomic weight 1. Also that the molecule of hydrochloric acid HCl , consists of 1 atom of chlorine with its atomic or combining weight 35·5, and one atom of hydrogen weighing 1.

22. Law of Multiple Proportions.—*When the same two elements can form several compounds, the different higher proportions of one element which combine with the same constant weight of the other element, are simple multiples of the lower.*

Ex. Nitrogen forms with oxygen five different compounds :

Nitrogen monoxide, -dioxide, -trioxide, -tetroxide, -pentoxide.



Here the amount of Nitrogen does not vary, while that of oxygen does. Therefore, for 28 parts of nitrogen there are:

in the first,	N_2O ,	1×16	or 16	parts of oxygen,
„	second, N_2O_2 ,	2×16	„ 32	„ „ „
„	third, N_2O_3 ,	3×16	„ 48	„ „ „
„	fourth, N_2O_4 ,	4×16	„ 64	„ „ „
„	fifth, N_2O_5 ,	5×16	„ 80	„ „ „

From which we see that the different higher proportions of oxygen, *viz.*, 32, 48, 64, 80, which combine with the same weight of nitrogen, *viz.*, 28, are simple multiples of the first, 16, that is, the weights bear to one another the simple ratio 1 : 2 : 3 : 4 : 5.

The fact is a natural consequence of the atomic theory. For, considering the nature itself or the indivisibility of atoms, and remembering that the atomic or combining weight of oxygen is invariably 16, when the proportion of oxygen increases, it must do so by the addition of *entire* atoms of oxygen, and

as each of these weighs 16, the higher proportions must necessarily be some simple multiple of 16.

The same remark applies to any other set of compounds of this kind, as CO and CO₂, SO₂ and SO₃, HgO and Hg₂O, etc.

23. Chemical action takes place more readily between unlike substances.—It has already been said that, as a rule, the greater the difference between the various elements or groups of elements that combine, the more readily the combination takes place.

It is well known that positive electricity repels positive, but attracts negative electricity. There is something analogous in the case of the chemical elements. Metals are considered as electro-positive, non-metals, on the other hand, as electro-negative, for reasons which will be seen later on. No affinity or attraction is known to exist between the metals, whilst these on the contrary, the electro-positive elements, manifest a great attraction for the non-metals or electro-negative elements.

Sodium and potassium, for example, which are so similar in properties, do not combine with one another to form a compound, whilst both these metals, very dissimilar to chlorine for example, or to oxygen, combine with them most readily. Similarly, bases or *metallic* oxides or hydroxides do not combine with one another, but they combine very readily with acids or *non-metallic* compounds to form salts.

24. Principle of Conservation of Matter, or *when a chemical reaction goes on, nothing new is created, nothing is really lost.*

There is such a continual wear and tear going on in the world, such a vast consumption of fuel, so many substances continually perishing or decaying, that there seems to be, at first sight, a gradual though slow diminution of the total amount of matter in the universe.

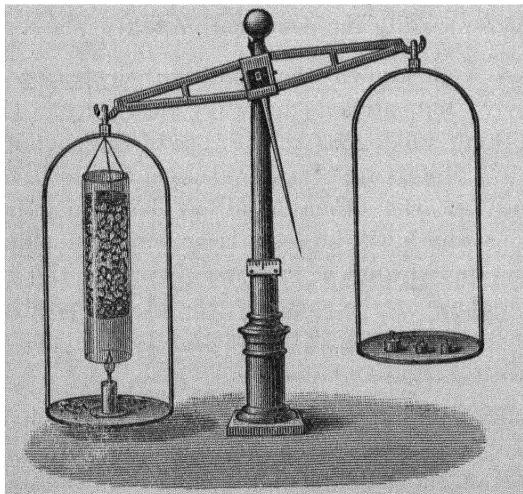


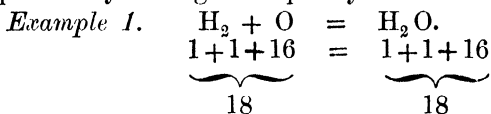
Fig. 6.—Experiment illustrating the principle of conservation of matter.

To use a well-known, familiar example, if we look at a burning candle, we see it growing smaller and smaller until at last it totally disappears. The matter

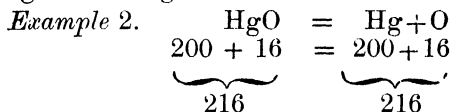
of the candle seems to be lost. An experiment will show us that this is not so. Place a candle upon the scale-pan of a balance, and over it suspend a wide glass-tube partly filled with fragments of caustic potash. Counterpoise the whole by placing weights in the opposite scale-pan (Fig. 6). Now light the candle; after some time we remark that the scale-pan carrying the burning candle goes down as if becoming heavier, and still the candle diminishes in size. On the one hand, therefore, there seems to be loss of matter so far as the candle is concerned, and on the other, there seems to be an increase in the total weight of the scale-pan. The fact is this:—The candle itself diminishes and finally disappears altogether; the candle as such is said to be lost. But it is not so with the matter or the constituents of the candle; these have been transformed into invisible gases or vapours, and for this reason the candle has gone out of sight, its constituents have combined with something else, *viz.*, with the oxygen of the air which also has weight, to form carbon dioxide and vapour of water, these being absorbed by the caustic potash suspended above it.

25. Chemical Equations.—The notion of chemical equations now flows as a natural consequence from the law of conservation of matter. Since in a chemical reaction nothing new is created and nothing is really lost, when the reaction is over there must necessarily be as much matter remaining in the

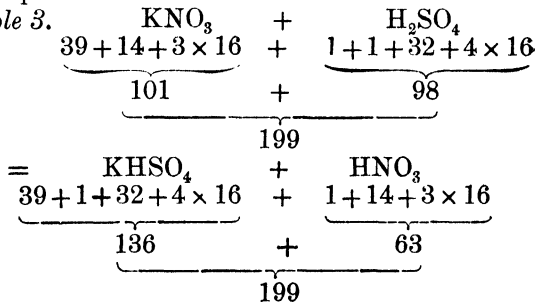
products formed, as there was in the original substances, though, of course, in a different form. This we represent by an equation. On the left-hand side of the equation we write down the substances which undergo changes or which react upon one another, and on the right-hand side the products formed. In virtue of the principle of conservation of matter, the total weight of the latter must be equal to the former, and therefore we separate the two members of the equation by the sign of equality.



i.e., the weight of the water formed by the combination of hydrogen and oxygen, must be equal to the weights of the gases that have combined.



i.e., when mercuric oxide is decomposed by heat, the weights of mercury and of oxygen set free, must be together equal to the amount of oxide taken.



i.e., when hydrogen sulphate reacts upon potassium nitrate to form potassium sulphate and hydrogen nitrate, the total weight of the latter two must be equal to that of the former two.—To sum up, a chemical equation is a short way of expressing, by means of chemical formulæ or symbols, 1° what happens in a chemical reaction, 2° how much of each of the various materials must be taken, 3° how much of each of the several products is formed. It is based on the *principle of conservation of matter*, or on the fact that in a chemical change there is neither destruction nor creation of matter but merely a re-distribution of the reacting materials.

N.B.—1. With the help of such equations, chemical problems can be worked out. If we consider, for instance, Ex. 3, we see that with 98 parts by weight (pounds, ounces, grains or grammes) of sulphuric acid, or with 101 parts of potassium nitrate we can obtain 63 parts of nitric acid. Consequently by a very simple calculation we may know what amount of nitric acid can be obtained from any weight of potassium nitrate or of sulphuric acid, and *vice versa*.

2. The manner of writing molecular formulæ in a chemical equation will be explained in the next chapter in connection with atomicity.

CHAPTER V.

Atomicity or Valency.

Remark.—It is not the purpose of the present chapter to discuss the still unsettled question of atomicity. The only reason for introducing the subject here, is its practical usefulness in writing chemical formulæ and chemical equations. Consequently the question

is treated only so far as it explains the formulæ of the more ordinary chemical compounds with which we have to deal.

Moreover, it must be borne in mind that the valency of elements is not so absolutely fixed as not to vary in some cases. On the contrary, its variability, which may be due to temperature, pressure etc., is in perfect agreement with the modern views of chemical dynamics.

26. Atomicity of elements.—When treating of chemical affinity we simply considered it as that force which causes atoms to combine with other atoms in general, but did not say what number of atoms of one kind unite with an atom of another kind. Here we shall have to consider *how many* atoms of one kind of matter, can combine with an atom of another kind. This combining capacity of elements with one another, so far as *number* of atoms is concerned, is usually called *atomicity* or *valency*.

To represent this notion of atomicity more vividly to our minds, let us imagine five different sets of boys, the boys of each set possessing, *by a strange supposition*, a different number of hands, ever ready to catch hold of something.

Each single-handed boy can catch hold of the only hand of another single-handed boy; whilst a two-handed boy can grasp either the two hands of another like himself, or two single-handed boys. A three-handed boy catches hold of three single-handed boys or of another like himself. A four-handed boy gets hold of four single-handed, or of two two-handed ones, or of one four-handed like himself; whilst a five-

handed monster grasps five one-handed fellows or one similar to himself, or five hands of any number of boys.

Now suppose the boys, in the above example, represent the atoms, the different number of hands will represent the different combining capacities or points of attraction, possessed by the various atoms for other atoms. (We might perhaps also imagine those different points of attraction to act like the poles of a magnet or like bodies charged with equal amounts of opposite electricity).

Now it has been found that

one atom of Cl can combine with one atom of H in HCl;

„ „ O „ „ two atoms „ „ H₂O;

„ „ N „ „ three „ „ „ NH₃;

„ „ C „ „ four „ „ „ CH₄;

moreover one atom of N combines with five atoms of similar atomicity in NH₄Cl.

Therefore :

the atomicity of Cl and H is one, Cl and H are mon-

„ „ O is two, O is diatomic ; [atomic ;

„ „ N is three, N is triatomic ;

„ „ C is four, C is tetratomic ;

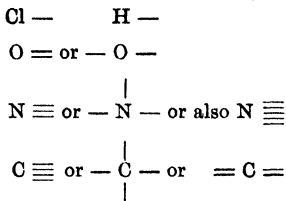
„ „ N is five, N is also pentatomic.

Thus we see that Nitrogen is both pentatomic and triatomic or trivalent.

It is to be remarked that, when the same element exhibits several combining capacities, the highest of them is often called its *atomicity*, for clearness' sake only ; while for the lower ones the name of

valency is preferred. Thus in ammonia NH_3 , Nitrogen is trivalent, or its actual valency is three; and it is pentatomic in NH_4Cl and in NH_4OH , or its atomicity is five.

The atomicity of the elements is expressed graphically by a corresponding number of lines attached to a symbol, thus :



TABLE

of the more important elements, classified according to their atomicity.

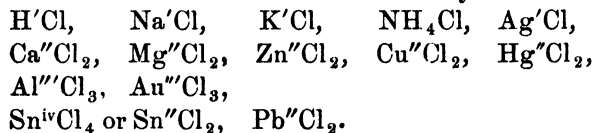
N.B.—The Roman numbers in front of each set denote the atomicity, and the marks affixed to some symbols denote their more usual valency. Thus Pb is in the group of tetratomic, and N among the pentatomic elements; but N''' and Pb'' signify that N is also trivalent as in NH_3 , and Pb more commonly divalent as in PbCl_2 , $\text{Pb}(\text{OH})_2$.

I. Monatomic	{ H, F, Cl, Br, I, Na, K, (NH_4), Ag. ✓ }
II. Diatomic	{ O, S; Ca, Sr, Ba, Mg, Zn, Cd, 'Cu', 'Hg' }
III. Triatomic	{ B, Al, As. ✓ }
IV. Tetratomic	{ C, Si, Sn'', Pb'', Cr'', Mn'', Fe'', Co'', Ni'', Pt'' }
V. Pentatomic	{ N''', P''', As''', Sb''', Bi''' }

27. Atomicity of compound radicals.—Not only single atoms but also groups of atoms have their atomicity. Thus in water H_2O (hydrogen monoxide), or HOH (hydrogen hydroxide), it is evident that the group OH , taken separately, is monatomic, since it combines with one monatomic element H , thus: $H-OH$. For similar reasons, in hydrogen nitrate HNO_3 (nitric acid), the radical NO_3 is monatomic $(NO_3)'$. In hydrogen sulphate H_2SO_4 (sulphuric acid), the radical SO_4 is diatomic $(SO_4)''$. In hydrogen carbonate H_2CO_3 (carbonic acid), the radical CO_3 is diatomic $(CO_3)''$.

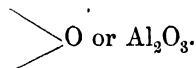
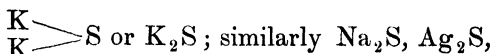
Therefore, with the help of the above table and the formulæ for the chief acids HCl , HNO_3 , H_2SO_4 , H_2CO_3 ...we are now able to write the formula for any common compound, even without having ever seen their formula written down, remembering also that the hydrogen atoms of acids can be replaced by metals, to form corresponding salts. (See definition of acids and salts).

Thus to write the formula of any chloride, remember that chlorine is monatomic and see what the atomicity of the other combining element is. The chief chlorides will therefore naturally be:



Similarly for the hydroxides ending in $(OH)'$:





28. Remarks.—

(1) It is more generally admitted that, although sometimes two valencies or *affinities* may disappear or remain *latent* (as in CO, C=O) by neutralizing one another, yet one alone cannot thus disappear or remain without uniting with something else.

In some cases Hg and Cu have one only of their two valencies free to act on an atom of a different kind, as in the groups —Hg—Cl and —Cu—Cl. In this case one group—Hg—Cl unites with another

similar group forming the molecule $\begin{array}{c} \text{Hg—Cl} \\ | \quad \cdot \\ \text{Hg—Cl} \end{array}$ or Hg_2Cl_2 , in which

the two mercury atoms satisfy one affinity between themselves.

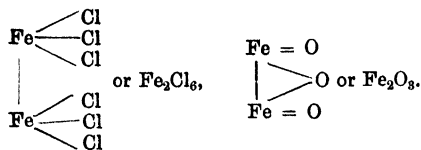
The same applies to Cu in $\begin{array}{c} \text{Cu—Cl} \\ | \\ \text{Cu—Cl} \end{array}$ or Cu_2Cl_2 . In these examples

mercury and copper are apparently monovalent, but from other considerations they are in reality diatomic. This is what is meant by 'Cu' and 'Hg' in the above table.

(2) Fe is divalent, Fe'', or tetratomic, Fe^{iv}. But in the latter case it is to be observed that in such compounds (ferric compounds, see below) two iron atoms generally are present, satisfying one affinity between themselves, so that six only out of the total

eight, remain free, thus $\begin{array}{c} \text{Fe}=\ \\ | \quad \diagdown \\ \text{Fe}=\end{array}$, and the group Fe₂ may be

considered as hexatomic :



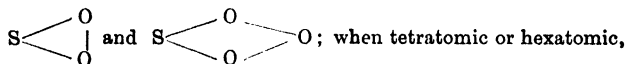
(3) Compounds in which all the atomicities are satisfied, are distinguished by the conventional termination "ic"; those in which all the atomicities are not satisfied, by the termination "ous".

Chief "ous" compounds :

Chief "ic" compounds :

Cuprous	chloride	Cu_2Cl_2	Cupric	chloride	CuCl_2
Mercurous	,,	Hg_2Cl_2	Mercuric	,,	HgCl_2
Cuprous	oxide	Cu_2O	Cupric	oxide	CuO
Mercurous	,,	Hg_2O	Mercuric	,,	HgO
Stannous	chloride	SnCl_2	Stannic	chloride	SnCl_4
Ferrous	,,	FeCl_2	Ferric	,,	Fe_2Cl_6
Ferrous	hydroxide	$\text{Fe}(\text{OH})_2$	Ferric	hydroxide	$\text{Fe}_2(\text{OH})_6$
Ferrous	oxide	FeO	Ferric	oxide	Fe_2O_3
Ferrous	sulphate	FeSO_4	Ferric	sulphate	$\text{Fe}_2(\text{SO}_4)_3$

(4) When sulphur is diatomic, we explain the compounds SO_2 and SO_3 by the following representation :



as it is believed to be by some authors, then we have a different way of representing them, viz. : $\text{O}=\text{S}=\text{O}$ and $\text{O}=\text{S}=\text{O}$. ✕



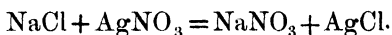
29. Manner of writing Chemical Equations.—

Here is the place to add a few hints as to the manner of writing chemical equations. We suppose the more common case in which two substances react upon one another by double decomposition. Two such substances being therefore given, we must consider what

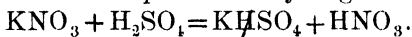
new substances are formed, and then write down correctly their formulæ. Now as to the correct formulæ, the table of atomicities and the formulæ of the ordinary acids will guide us, as seen above; and in order to know what substances are formed, it will be enough to remember in general that the new substances are of the same denomination as those which react upon one another, except that an exchange takes place between the metallic radicals.

For example :

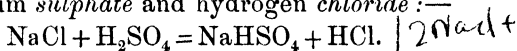
A *chloride* and a *nitrate* produce a *nitrate* and a *chloride* :—



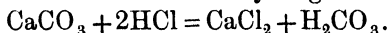
Similarly, potassium *nitrate* and hydrogen *sulphate* produce potassium *sulphate* and hydrogen *nitrate* :—



Also, sodium *chloride* and hydrogen *sulphate* produce sodium *sulphate* and hydrogen *chloride* :—



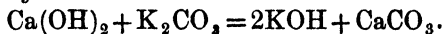
Again calcium *carbonate* and hydrogen *chloride* produce calcium *chloride* and hydrogen *carbonate* :—

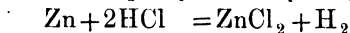
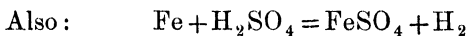


N. B.— H_2CO_3 being very unstable is decomposed into H_2O and CO_2 , so that practically the reaction is to be represented as follows :—

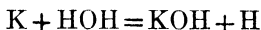


Calcium *hydroxide* and potassium *carbonate* produce potassium *hydroxide* and calcium *carbonate* :





that is, an element (Fe or Zn) and a *sulphate* or a *chloride* produce a *sulphate* or a *chloride* and an element (H), and



that is, potassium and hydrogen *hydroxide* (another manner of representing water), produce potassium *hydroxide* and hydrogen.

CHAPTER VI.

Classification of Elements and Compounds.

30. Classification of the Elements.—For convenience' sake the elements are generally divided into metals and non-metals. Under a physical point of view the metals are solid at the ordinary temperature, except mercury which is liquid; they have a characteristic shining appearance, usually spoken of as their *metallic lustre*; they are generally hard bodies, more or less malleable and ductile; they are good conductors of heat and electricity; their oxides and hydroxides form bases; when isolated from a combination by the electric current, they go to the negative pole; hence they are often called electro-positive elements.

The non-metals (or metalloids, as some chemists like to call them) are gaseous or solid, bromine being the only liquid non-metal. When solid, they have generally a dull, non-metallic aspect, and are rather

soft or brittle; they are bad conductors of heat and electricity; their hydroxides form acids, *i.e.*, bodies of an entirely different character from bases; when isolated from their combinations by the electric current, they go to the positive pole; hence they are called electro-negative elements.

Under the point of view of chemistry, this classification of elements into metals and non-metals is not a strictly scientific one, as it rests merely upon physical appearances; but it affords a certain convenience in an elementary course. More generally the elements are classified according to their atomicity, by grouping together into families such as exhibit similar chemical properties, as may be seen in the above table of atomicities.

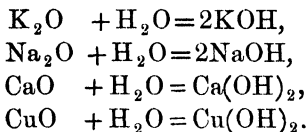
31. Classification of the compounds.—All the compounds which we have to study, may be classified under the following heads: **oxides, hydroxides** (including *bases* and *acids*) and **salts**.

1. **Oxides.**—*An oxide is a compound resulting from the combination of oxygen with any other element.*

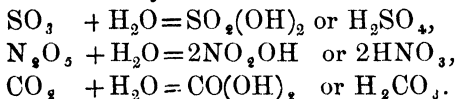
All the elements (except fluorine) given in the table just referred to, can form oxides. Corresponding to the metals and non-metals, there are naturally *metallic oxides* and *non-metallic oxides*. The former are, for the greater part, base-forming oxides or *bases*; the latter are acid-forming oxides, and are called *anhydrides*.

2. **Hydroxides**, sometimes but less correctly called *hydrates*, are compounds resulting from the direct or indirect combination of oxides with one or more molecules of water (usually in the form of OH, called hydroxyl).

Etc. gr. — Metallic hydroxides or *bases* :—



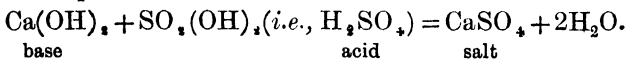
Non-metallic hydroxides or *acids* :—



The metallic hydroxides, as well as the greater number of metallic oxides, are bases; whilst the non-metallic hydroxides only are acids. When the former are put in presence of the latter, a new class of compounds, called *salts*, is produced.

(a) **Bases** are oxides or hydroxides that contain a metal, which, when the oxide or hydroxide is put in presence of an acid, exchanges place with the hydrogen of the acid to form a salt.

Example :—



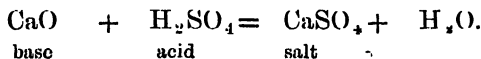
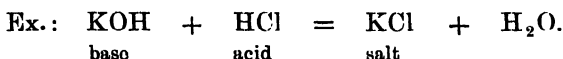
1. (*N.B.*)—Bases have properties the opposite of those of the acids. They also turn red litmus paper or solution blue.

(2) The *alkalis* are a class of strong bases very soluble in water. They are KOH, NaOH, NH₄OH. Thus we see that all alkalis are bases, but all bases are not alkalis.

(b) **Acids** are compounds containing one or more atoms of hydrogen, which can be replaced either partly or entirely by a metal or a radical playing the part of a metal (NH₄, ammonium) to form a salt.

Acids are readily distinguished from bases by their property of turning blue litmus red or by their easily decomposing carbonates and causing effervescence.

· **3. Salts.**—The definition of a *salt* is easily understood from the definition of acids and bases, viz., a *salt is that body which results from the combination of an acid with a base, in which reaction the hydrogen of the acid is replaced by the metal of the base and water is formed.*



Acid, normal and basic salts.—When only a portion of the atoms of hydrogen in an acid is displaced by a metal, we have an *acid salt*, as KHSO_4 (potassium acid sulphate), NaHCO_3 (sodium acid carbonate, usually called sodium bicarbonate), Na_2HPO_4 (sodium hydrogen phosphate, from H_3PO_4 , phosphoric acid).

When all the atoms of hydrogen are replaced, we have a *normal salt*, sometimes, but less correctly, called *neutral salt*, as K_2SO_4 , Na_2CO_3 , Na_3PO_4 .

On the other hand, when the proportion of the base is greater than is required to replace the hydrogen of the acid to form a normal salt, we have a *basic salt*, consisting of a normal salt and one or more additional molecules of the base. Ex.: $\text{PbCO}_3 \cdot \text{Pb(OH)}_2$, basic lead carbonate, commonly called white lead; $\text{HgSO}_4 \cdot 2\text{HgO}$, basic mercury sulphate.

Caustic sodium

CHAPTER VII.

Combustion—Oxidation—Flame.

32. Nature and definition of combustion.—The nature of combustion was unknown till a rather recent epoch, and very strange opinions have at times been put forward to explain it. Even at the present time, the popular meaning of combustion is destruction by fire; but this is merely one of the effects of combustion.

The scientific explanation of combustion was first given about a hundred years ago by the French chemist Lavoisier. According to him, *to burn* means *to combine with oxygen*. Now, when this combination with oxygen takes place *slowly*, the phenomenon is usually called **oxidation**,* whilst the name of **combustion** is given to that kind of *combination with oxygen which takes place rapidly and is usually accompanied by the production of heat and light*. This is the definition of combustion in its *ordinary* meaning.

* The contrary phenomenon of removing oxygen from a compound is called **reduction**. Thus, when lead oxide is heated by means of a blow-pipe in a cavity on charcoal, it is reduced to metallic lead, and when carbon dioxide gas (CO_2) is made to pass through red-hot charcoal, it is reduced to carbon monoxide (CO) by the carbon of the charcoal.

But in a wider and more general sense, combustion implies every chemical combination. Thus we see



Fig. 7. Antimony burning in chlorine.

that oxygen is not always, strictly speaking, required for combustion, which may take place with any other element that enters into combination.

The commonest examples of combustion in which oxygen does not take part, are: antimony (Fig. 7), or phosphorus, or finely divided copper burning in chlorine; copper or iron combining with sulphur, the phenomenon being

accompanied by incandescence; phosphorus burning in contact with iodine.

The more familiar cases of combustion in which oxygen takes part are the burning of wood, oil, paper, a candle, etc.

33. Combustible bodies and supporters of combustion.

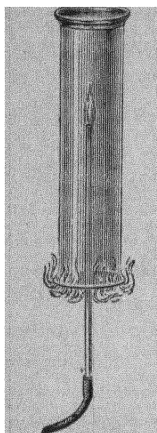


Fig. 8.—Oxygen burning in hydrogen.

—A substance which burns is usually called a *combustible body*, and the one in which or with which it burns, is said to be a *supporter of combustion*. Ex.: charcoal, oil, paper, hydrogen, are combustible bodies; air or oxygen, supporters of combustion.

But these terms are simply relative: for a jet of oxygen gas burns in an atmosphere of hydrogen exactly in the same way as a jet of hydrogen burns in an atmosphere of oxygen. To show this by an experiment, fill a cylindrical glass-vessel with hydrogen (Fig. 8). Light the gas at the mouth and introduce into the vessel a jet of oxygen gas. The oxygen takes fire at the burning hydrogen and continues to burn quietly inside the cylinder in the atmosphere of hydrogen.

34.—Conditions of combustion.—(1) In order to start combustion, the temperature of the combustible substance must usually be raised to a certain degree at some point of the mass. When once begun, combustion generally continues by itself, owing to the heat produced by the chemical action.

Conversely, to put out fire, for example a burning candle, the temperature must again be lowered; this may be done either by holding a cool object over it, or by blowing upon it.

(2) There must be a free supply of oxygen, and this oxygen is usually taken from the air. Hence

the use of bellows, or of mouth blow-pipes for minor laboratory operations, and the special arrangements in the construction of chimneys to produce draughts ; this is to increase the amount of air, that is, of oxygen, in a given time, and consequently to activate combustion.

On the contrary, to diminish or altogether to prevent combustion, it is sufficient to lessen the free access of air or to stop it completely. Hence a practical method of putting out a lamp which takes fire, is to cover it with sand, or to wrap a cloth around it, the latter method being still more effective if the cloth is wet.

(3) In order to burn, the body must be capable of combining with oxygen ; consequently, all saturated compounds can no longer burn. Thus, when hydrogen has once been oxidized or burnt, and hydrogen monoxide, commonly called water, is formed, this cannot burn any longer. Hence water is most effective in putting out fire, the more so as it easily penetrates into the pores of the burning substance and thus prevents access of air. It is for this reason that green wood does not easily burn.

Carbon monoxide can burn, *i.e.*, combine with more oxygen to form carbon dioxide. But carbon dioxide is unable to combine with more oxygen ; it is called an incombustible body. Sand, glass, stones and rocks are likewise substances saturated with oxygen ; they also are incombustible.

NOTE.—Some substances may burn freely and even very energetically in vacuo, or in a closed vessel, or when blocked up in the bore of a mine without any access of air, as gunpowder, tinder, etc. In this case, the burning substances are in presence of other bodies well supplied with oxygen, which they give up easily and abundantly, as potassium nitrate or chlorate.

35. Quick and slow combustion.—All combustion or chemical combination goes on more or less energetically, and evolves more or less heat. Sometimes it takes place so slowly that the heat evolved at any given instant of time is too small to be perceived; as for example, when metals become rusty or oxidize in damp air, when wood decays, and so on. This is called slow combustion. On the other hand, when combustion goes on rapidly and energetically, the heat evolved is great enough to be perceived, and is often sufficient to raise the combining substances, or the products formed, to *incandescence*, that is, they become so hot as to emit light, e.g., when iron or carbon burn in oxygen; it is a case of quick combustion. When the incandescent substance is a gas or vapour, then we have what is called a *flame*.

It is to be remarked that when a piece of iron, for example, gets *slowly oxidized* or *rusty* in the air throughout its entire mass, the same amount of heat is, on the whole, evolved as when the *same weight* of iron is rapidly burnt with incandescence in oxygen gas.

36. Respiration a case of slow combustion.—The act of respiration is a phenomenon of slow combustion. The oxygen taken up by the blood in the lungs and carried all over the body, combines with

(i.e., burns) the waste products in the living tissues, forming chiefly carbon dioxide CO_2 and vapour of

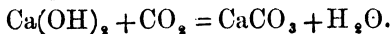


Fig. 9.—Breathing through lime-water.

water H_2O . The heat evolved by the combination all over the body is the cause of its temperature, and is known as *animal heat*.

Therefore, so far as the products of combustion are concerned, respiration is somewhat similar to the burning of a candle.

Thus when we breathe through clear lime-water (Fig. 9), or when we make the products of a burning candle pass through it (by an arrangement similar to that in Fig. 10), a kind of milkiness is produced in the water due to the formation of insoluble CaCO_3 , and this indicates the presence of carbon dioxide in each case:—



Again, if we breathe over a cool, polished surface, such as a piece of glass or a slate, or if we hold over

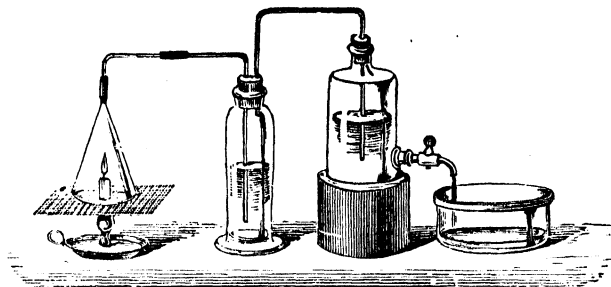


Fig. 10—The products of a burning candle made to pass through lime-water.

the flame of a candle a glass tumbler kept cool, little drops of water will make their appearance. These are due to the condensation of the vapour of water resulting from the combustion.

37. Luminosity of flame.—It has been said that *flame* is nothing else but burning gases or vapours, raised to incandescence, *i.e.*, emitting light. The luminosity of flames is due to two chief causes: to the *presence of solid particles* in the flame, according to Davy, and to *pressure*, according to Frankland. Thus the flame of pure hydrogen is scarcely visible, its yellowish appearance being mainly due to the sodium compounds present in the glass used as a jet tube; but if we hold in it some solid substance, such as a platinum wire, or if we drop upon it finely divided carbon dust, the flame becomes luminous.

The luminosity of oil lamps, of burning coal gas, or of candles, is generally due to the presence of solid particles of unburnt carbon. (See structure of flame, Fig. 11.)

As for pressure, it has been observed that a candle, for example, burns less brilliantly at the top of a high mountain, where the atmospheric pressure is less than at its foot, where the pressure is greater.

38. Structure of the candle flame.—The flame of a burning compound body is not uniform throughout; thus if we observe, for instance, the flame of a candle (Fig. 11), we easily distinguish in it three different parts:

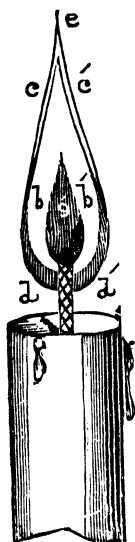
(1) An *obscure elongated inner or central cone, a*, consisting of unburnt gases surrounding the wick, where the temperature is not great.

The heat of the flame causes the solid matter of the candle to melt; the liquid, ascending through the wick by capillarity, is decomposed into combustible gases which constitute the dark part *a* of the flame; those gases, for want of oxygen, do not burn.

The existence of unburnt gases in this part of the flame may be shown by bringing into it one end of a bent glass tube; after a while the unburnt gases will pass up the tube, and may be ignited at the other end, by applying a lighted taper to them.

(2) A *luminous middle zone bb'*, surrounding the central dark cone; here combustion is incomplete and the temperature comparatively high.

In this part of the flame, the gases, compounds of carbon and hydrogen, coming from the inner cone, receive some amount of oxygen from the air; but



this supply is not yet sufficient to burn away all the carbon, so that part of the latter is set free in the form of soot; the solid carbon particles, heated to incandescence, impart luminosity to this portion of the flame.

The existence of free carbon particles in the middle zone is proved by the fact that any cool object, as for example a glass rod, placed within it is quickly covered with soot.

(3) A *non-luminous outer zone or envelope cc'*, where combustion is complete and the temperature highest.

Fig. 11.—Candle flame.

The white-hot particles of carbon coming from the middle zone, are now sufficiently supplied with oxygen and burn away to CO_2 ; and as there are no more solid particles present here, this portion of the flame is non-luminous.

The principle of luminous and non-luminous flames is also well illustrated by the help of a Bunsen burner (Fig. 12). The combustible gas is made to enter the brass tube at the bottom. Air is at the same time

admitted through openings in the lower part of the tube. The gas thus mixes well with the air and burns at the top of the tube. Now, according as we regulate the admission of air, the flame is more or less bright. If there is air, *i.e.*, oxygen enough to combine with all the

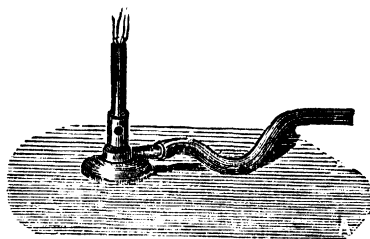


Fig. 12.—Bunsen burner.

carbon of the gas, the flame will be pale and scarcely visible. If we partly close the air-holes at the bottom of the burner, the flame will become more visible but also more smoky on account of the unburnt particles of carbon, for the complete combustion of which the supply of oxygen is now insufficient.

39. Use of glass chimneys.—A petroleum lamp having no glass chimney burns, as may easily be observed, with a ruddy, smoky flame; but no sooner is a properly fitting chimney added than the light becomes brilliant and smokeless. The action of the chimney is to cause a draught of air *round* the flame, which by giving a greater supply of oxygen causes a more perfect combustion of the carbon particles in the *outer zone*, and at the same time raises the temperature of the luminous part (see structure of the flame), which consequently emits more light.

PART II.

NON-METALS.

NOTE.—The number of non-metals treated of here, is too small to admit of the more scientific classification which is based upon atomicity. We shall therefore adopt a more practical division: first we shall study the four principal *gaseous*, and next the four *solid* non-metals, noticing, as we go, their more important compounds.

CHAPTER I.

† OXYGEN.

Symb., O. At. wt. = 16. *Diatomic.*

40. Occurrence.—Oxygen exists in both the *free* and *combined* state, and is the most abundant and most widely distributed element in nature; for—

(a) about half the weight of the earth's *solid* crust consists of oxygen;

(b) $\frac{8}{9}$ of the weight of the *liquid* water of the earth consists of the same;

(c) $\frac{1}{5}$ of the volume of the *gaseous* atmosphere consists of oxygen in the free state.

41. History.—Oxygen was discovered by Priestley, an English chemist (1774). Its chemical character and the part it plays in nature, were studied by Lavoisier a French chemist, who gave it its name *oxygen* (Gr. acid producer), because he believed it to be a necessary constituent of acids.

42. Preparation.—Oxygen is most commonly prepared by heating some substance which contains it and which easily gives it up. Pure oxygen is got by

heating mercuric oxide or potassium chlorate, or by the electrolysis of water; whilst, for practical purposes and on a large scale, oxygen is obtained by heating potassium chlorate mixed with manganese dioxide.

(1) **Oxygen obtained from mercuric oxide**—Some red or yellow mercuric oxide, HgO , is put in a glass

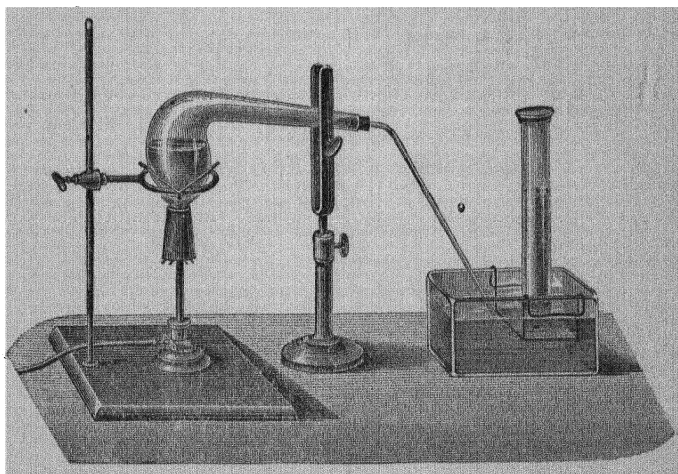
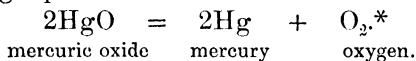


Fig. 13.—Preparation of oxygen from mercuric oxide.

flask or retort (Fig. 13), provided with a delivery tube, the end of which dips under water, and over it a glass tube, filled with water, is inverted. Heat being applied to the retort, gas is seen to bubble through the water at the extremity of the delivery tube. At first this gas is nothing but heated air and is allowed to escape, but after a short time pure

oxygen is liberated, and is collected in the inverted tube. At the same time we begin to see in the upper and cooler parts of the retort a shining deposit of metallic mercury, which condenses there, having been separated from the oxide of mercury by the action of heat. This decomposition may be represented by the following equation:—



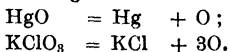
Note.—Considering that the atomic weight of mercury is 200, and that out of 216 parts by weight of mercuric oxide, which is very costly, we get but 16 of oxygen, this method is not economical where oxygen is required in large quantities.

(2) **Oxygen obtained from KClO_3 and MnO_2 .**—The same arrangement is made as in the previous case (Fig. 13), but instead of mercuric oxide we use potassium chlorate KClO_3 , mixed with manganese dioxide MnO_2 , or black oxide of manganese. The action may be represented by the equation—



* The reason why we write 2HgO and 2KClO_3 , is to obtain in the second part of the equation entire molecules of oxygen. As in the greater number of elements, chiefly of gases, the molecule contains two atoms, the molecules of oxygen, hydrogen, nitrogen, chlorine etc., are represented by O_2 , H_2 , N_2 , Cl_2 , etc. By exception the molecule of mercury contains only one atom; hence we write 2Hg which signifies two molecules, and not Hg_2 .

If we merely wish to represent the substances that result from the decomposition of HgO and KClO_3 , without taking into consideration the correct molecular representation, it is enough to write the equations in the following manner:—



This remark applies to many other similar equations. The relation between the weights remains exactly the same, and this method affords greater facility in working chemical problems.

Manganese dioxide is not represented in the equation, because it does not seem chemically to take part in the reaction, as the substance remains the same at the end of the experiment as it was at the beginning. It is a fact that the presence of manganese dioxide helps to obtain the oxygen more easily and regularly, but how it does it is not well known.

The more probable opinion seems to be, that manganese dioxide, being a good conductor of heat, allows the heat to be rapidly and uniformly distributed throughout the mixture. The potassium chlorate is thus decomposed at once, and is not, as happens when it is heated alone, first transformed into potassium perchlorate, which gives up oxygen at a much higher temperature only.

This is moreover confirmed by the fact that other metallic oxides, as those of iron, copper or lead, produce the same effect as manganese dioxide.

To show the influence of manganese dioxide in this case, heat some potassium chlorate alone in a test-tube, and in another test-tube heat potassium chlorate mixed with manganese dioxide. Bring a glowing splinter of wood to the mouth of each tube and remark how rapidly the splinter bursts into flame in the latter case.

43. Properties of Oxygen.—Oxygen is a colourless, tasteless, non-inflammable gas, a little heavier than air, *i.e.*, when the weight of a certain volume of hydrogen is represented by 1, the same volume of air will weigh 14.44 times more, the same volume of oxygen 16 times more.

It is to a small extent soluble in water. It is this oxygen, dissolved in water, which fishes breathe, and not the oxygen chemically combined with hydrogen.

Oxygen combines with all the elements, except fluorine, to form *oxides*. This combination sometimes takes place with great energy giving rise to the phenomenon of heat and light; at other times the combination takes place slowly, no heat and no light being perceived (Cf. no. 35).

To burn, it has been said, generally means to combine with oxygen. Bodies therefore, which burn in air, *i.e.*, in oxygen diluted with four times its volume of nitrogen, burn with increased rapidity and brilliancy in pure and undiluted oxygen, which, on this account, is said to be a good supporter of combustion.

44. Tests for Oxygen.—If a taper with only a red-hot tip at the end, or a glowing chip of wood, be introduced into a jar containing oxygen, it will burst into flame and burn brilliantly, carbon dioxide being formed, as can be seen by applying the lime-water test.

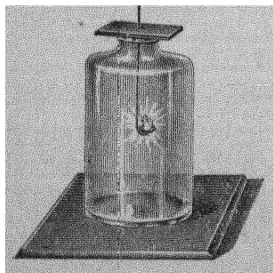


Fig. 14.—Sulphur burning in oxygen.

In air sulphur burns with a pale blue flame; but in oxygen the flame is bright violet, sulphur dioxide being formed, which may be recognised by its peculiar smell, and by changing the colour of litmus paper, which has been moistened previously, from blue to red.

Phosphorus also burns with a very dazzling bright flame in oxygen.

Metals too burn in oxygen. Take a piece of iron or steel wire bent in the form of a spiral, and dip one end, which has been previously heated, into flowers of sulphur. Next light the sulphur which has adhered to the wire, and introduce it into a jar filled with oxygen. The iron will be seen to burn with a very brilliant light, but without flame, and an oxide of iron will be left behind.

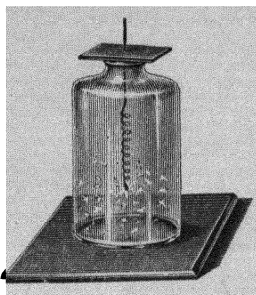


Fig. 15.—Iron burning in oxygen.

A bit of tinder may also be used, instead of sulphur, to ignite the wire.

45. Use.—The chief use of oxygen in nature is to help combustion and the respiration of animals.

46. Ozone.—Oxygen exists in an allotropic form, *ozone* (Gr. smell), so called because it is the presence of ozone which causes that peculiar smell perceived when an electric machine is worked.

Oxygen contains two atoms in its molecule, O_2 , whilst ozone contains three, O_3 .

Ozone seems to be a very valuable bleaching agent and disinfectant.

XHYDROGEN.

Symb. H. At. wt. = 1. *Monatomic*

47. Occurrence.—Hydrogen is rarely found *free* in nature, but it exists abundantly and widely distributed in the *combined* state, for—

(a) it is one of the chief constituents of water : $\frac{1}{9}$ by weight, $\frac{2}{3}$ of the gaseous volume ;

(b) it is one of the chief constituents of all vegetable and animal tissues. Besides, hydrogen is one of the necessary constituents of all acids.

48. History.—Hydrogen has been long known under the name of “inflammable air.” Lavoisier gave it the name of *hydrogen* (Gr. water-producer), because by combining with oxygen it forms water. Pure hydrogen was first obtained by Cavendish (1766).✓

49. Preparation.—The chief compounds from which hydrogen is usually prepared, are *water* and *acids*, especially hydrochloric and sulphuric acids.

A. Preparation of hydrogen by decomposing water.—Water may be decomposed in various ways : (1) by **Electrolysis**, (2) by the action of **metals**. This again is done (a) at the *ordinary temperature*, (b) at *high temperatures* only.

(1) Water is decomposed into hydrogen and oxygen by the process of **electrolysis**, *i.e.*, by passing an electric current through it (see next chapter). In this way we obtain pure hydrogen, but in small quantities only. As electricity is a physical agent, this method of obtaining hydrogen is called a *physical method*.

(2) Water may be decomposed by the action of **metals**.

(a) Water is decomposed at the ordinary temperature by the metals *sodium* and *potassium* ; and as these metals act chemically upon water, setting the

hydrogen free, this method of obtaining hydrogen is called a *chemical method*.

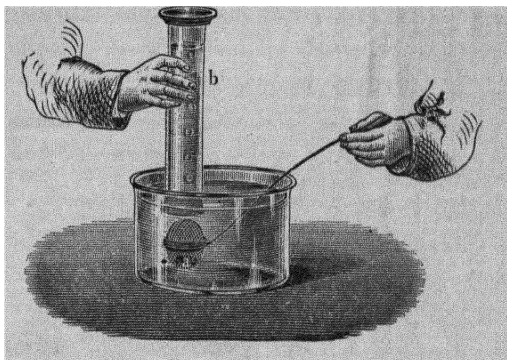


Fig. 16.—Collection of hydrogen.

When a fragment of potassium or sodium is thrown upon water, it floats about as a molten globule upon the surface, decomposing the water and liberating hydrogen. With potassium so much heat is developed, that the hydrogen evolved at once takes fire and burns (*Cf.* no. 118), so that this decomposition of water is of no *practical* use for the preparation of hydrogen. With sodium the action is less violent, and the hydrogen evolved does not inflame under ordinary circumstances. It may however be made to ignite by placing the metal on a piece of blotting paper which prevents it from moving about and from being so rapidly cooled by the water.

In order to collect the hydrogen liberated by the action of sodium upon water, the metal is enclosed in

some fine wire gauze and held below the surface of water, underneath the mouth of an inverted jar filled with water (Fig. 16); the gas ascends in the tube and displaces the water.

Another method of collecting the gas consists in mixing sodium with mercury, to get an amalgam of the two metals. This heavy amalgam will sink to the

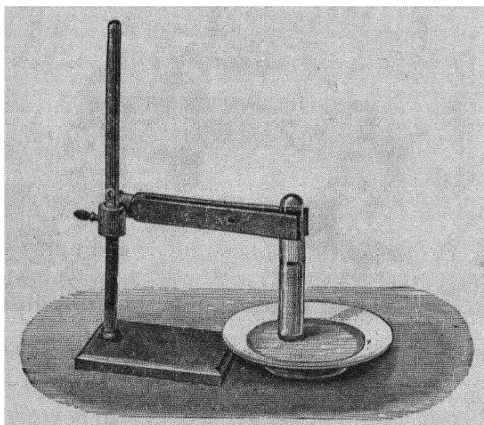
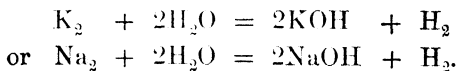


Fig. 17.—Preparation of hydrogen with sodium amalgam.

bottom of the vessel or dish of water into which it is dropped. Over it is inverted a test-tube filled with water, to collect the liberated hydrogen (Fig. 17).

The chemical reaction which takes place, may be expressed by the equation:—



(b) *Iron* is one of the metals which decompose water at a high temperature. For this a current of steam is made to pass through a porcelain tube, filled

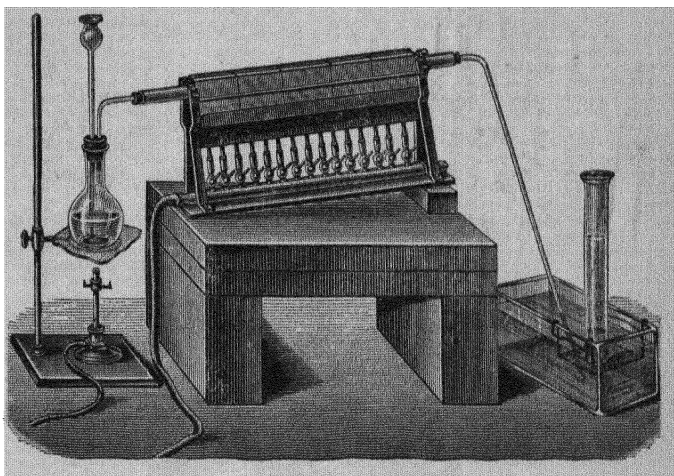
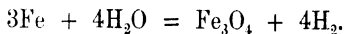
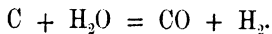


Fig. 18.—Hydrogen prepared by passing steam over heated iron.

with iron filings heated to redness by a set of gas-burners (Fig. 18). The iron combines with the oxygen of the water, whilst the hydrogen is set free and is collected in the usual way:—



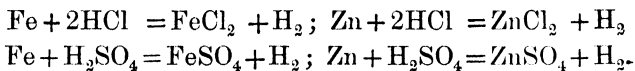
NOTE. When steam is passed over red-hot charcoal, hydrogen is also formed but mixed with carbon monoxide, CO:—



This mixture of hydrogen and carbon monoxide is called *water gas*, and is used in some countries for illuminating purposes; but the poisonous nature of carbon monoxide renders its use dangerous.

B. *Preparation of hydrogen by decomposing other compounds of it* :—

This is usually done, at the ordinary temperature, by the action of hydrochloric or sulphuric acid on iron or zinc :—



It is the most convenient method of preparing hydrogen, and the one usually employed.

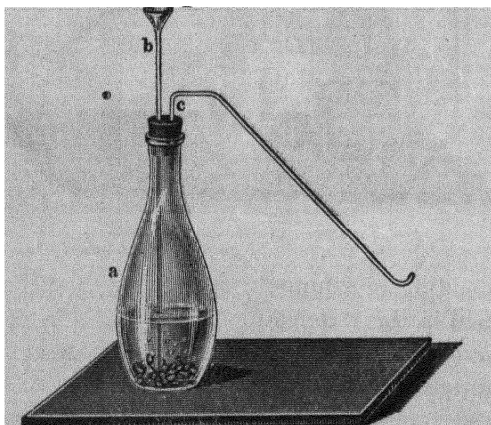


Fig. 19.—Hydrogen prepared by pouring HCl upon Zn.

Into a flask *a* (Fig. 19) a sufficient quantity of zinc, cut into small pieces,* is introduced, and some water

* Zinc is preferred to iron on account of its greater purity; and it is usually "granulated" *i.e.*, reduced to small fragments by being melted and poured into cold water.

is poured over it ; the water will serve the purpose of diluting the acid we shall use. Shut the flask with a cork provided with a bent delivery tube *c* and a safety

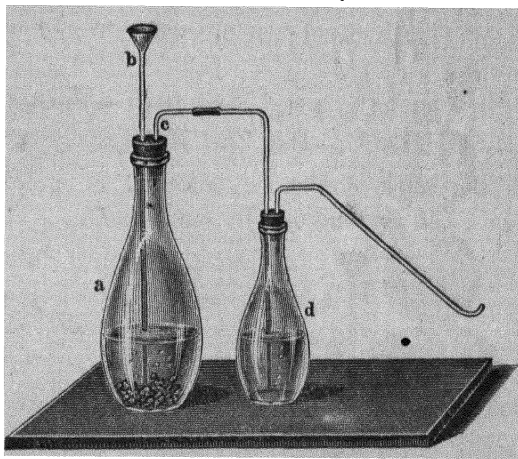
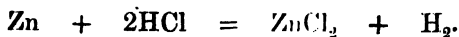


Fig. 20.—Hydrogen purified.

tube *b* ending in a funnel ; the safety-tube should be immersed in the water and should nearly reach to the bottom of the flask. If hydrochloric acid be now poured into the flask, through the funnel-tube, an effervescence will take place, due to the escape of bubbles of hydrogen through the liquid. The hydrochloric acid is decomposed by the zinc, zinc chloride being formed and hydrogen liberated :—



The hydrogen may be collected in jars or bottles filled with water and inverted in a water-trough ;

but before collecting the gas, the air which fills the flask should first be entirely displaced.

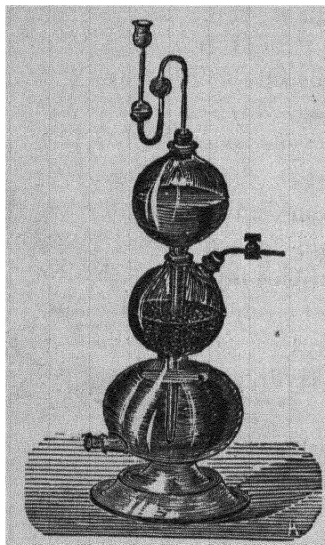


Fig. 21.—Kipp's apparatus.

When the action stops, a new quantity of acid may be supplied through the funnel-tube, so long as there remains zinc in the flask.

If it be desirable to obtain hydrogen free from traces of water, it is made to pass, before use, through a solution of caustic potash, or through sulphuric acid contained in a wash-bottle *d* (Fig. 20).

The above arrangement for preparing hydrogen is now generally replaced by Kipp's apparatus (Fig. 21), which gives a constant supply of hydrogen, and which allows of its action being stopped at will.

50. Properties.—Hydrogen is a colourless gas; it is also odourless and tasteless, when pure.

It is a very light gas, the lightest of all known bodies, being about $14\frac{1}{2}$ times lighter than air.

There are several ways to show that hydrogen is a light gas:—it can be collected by upward displacement,

i.e., by introducing the delivery tube into a test-tube

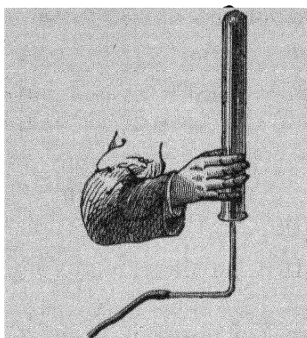


Fig. 22.—Hydrogen collected upwards.

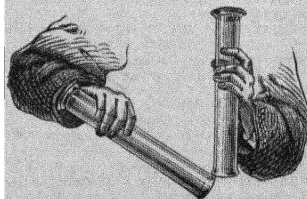


Fig. 23.—Hydrogen poured upwards.

held mouth downwards over it (Fig. 22);—it can be poured from one test-tube into another upwards (Fig. 23);—a bottle full of hydrogen burns with a larger flame when held mouth upwards than when held downwards;—soap bubbles filled with hydrogen easily rise in the air.

Hydrogen is an inflammable gas which burns with a pale-blue flame. It gives out very much heat, but does not itself support combustion. If a lighted taper be plunged into a jar of hydrogen, the gas

at the mouth of the jar will be lighted, but the taper will be extinguished within the jar (Fig. 24). On taking out the taper it will be rekindled by the burning gas. This is the ordinary test for hydrogen.

Before lighting a jet of hydrogen, care must be taken to allow the air which fills the generating flask to escape entirely, as otherwise an explosion might occur.

A mixture of hydrogen and oxygen is violently explosive. To show this without the least danger,

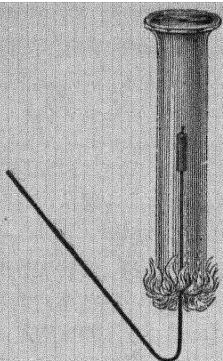


Fig. 24.—Hydrogen burns, but does not support combustion.

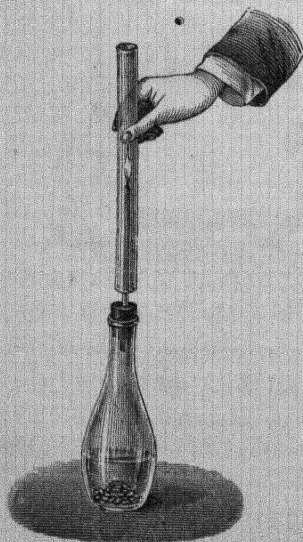


Fig. 25.—Harmonica chimica.

introduce a certain quantity of the mixture into a stout test-tube or glass bottle with a wide mouth, and apply a lighted taper. A loud explosion will be the result, but the tube or bottle will be found intact.

Another interesting way of performing the same experiment consists in setting fire to soap-bubbles that have been filled with the mixture, *but only after the soap-bubbles have been shaken off the delivery tube.*

To perform this last experiment on a larger scale, pour some soap-solution in an iron mortar, and allow a mixture of hydrogen and oxygen to pass through. Very large bubbles or agglomerations of bubbles are thus obtained, which will explode with a very loud report when a lighted taper is applied to them.

The explosions may take place with the production of a harmonious sound, if they are made to succeed each other rapidly and at regular intervals. These conditions are realized by burning a jet of hydrogen

in a glass tube (Fig. 25). This series of regular explosions receives the name of *Harmonica chimica*.

The product of the combustion of hydrogen with oxygen is water.

This may be observed in the previous experiment, as the tube, in which the hydrogen burns, becomes covered inside with fine droplets of water. The same may be shown on a larger scale by an arrangement similar to that of Fig. 26. The hydrogen generated in the Wolff's bottle A, is dried by passing over

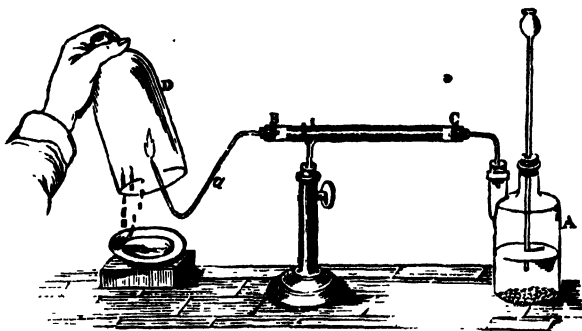


Fig. 26.—Water is formed when hydrogen is burning.

fragments of caustic potash, or calcium chloride, in the tube BC, and is burnt at the extremity of the bent tube *a*. Over the flame is inverted a glass bell-jar, in which the vapour of water condenses in small drops, that trickle down the sides of the glass vessel.

51. Uses.—Hydrogen is used :—to fill balloons, on account of its great lightness; to fuse platinum with

the oxy-hydrogen blowpipe, on account of the great heat produced when hydrogen burns in oxygen; to reduce some metallic oxides, on account of the great affinity of hydrogen for oxygen; for example, to reduce copper oxide CuO , in the quantitative analysis of water.

CHAPTER II.

+ WATER

or Hydrogen monoxide, H_2O .

52. Occurrence.—Water is abundant and widely distributed in both organic and inorganic nature.

(a) In **inorganic** nature water is found *upon the earth*, in the sea, in rivers, lakes, etc.; *in the sky*, in the form of vapour, mist, clouds, rain.

(b) Water is found in **organic** nature inasmuch as a great portion by weight of the substance of plants, vegetables, fruits and animal bodies consists of water.

53. History.—Until about the end of last century water was thought to be an element. Towards 1780 Cavendish in England, and Lavoisier in France, discovered that when hydrogen is burning water is formed, and thus found its composition.

54. Composition.—The composition of water has been determined both by *volume* and by *weight*.

A. Composition of water by volume.—The composition of water by volume has been ascertained

both by analysis or *decomposition* and by synthesis or *direct combination* of its components.

(1) **By analysis.**—The analysis or decomposition of water is effected by *electrolysis*. By electrolysis

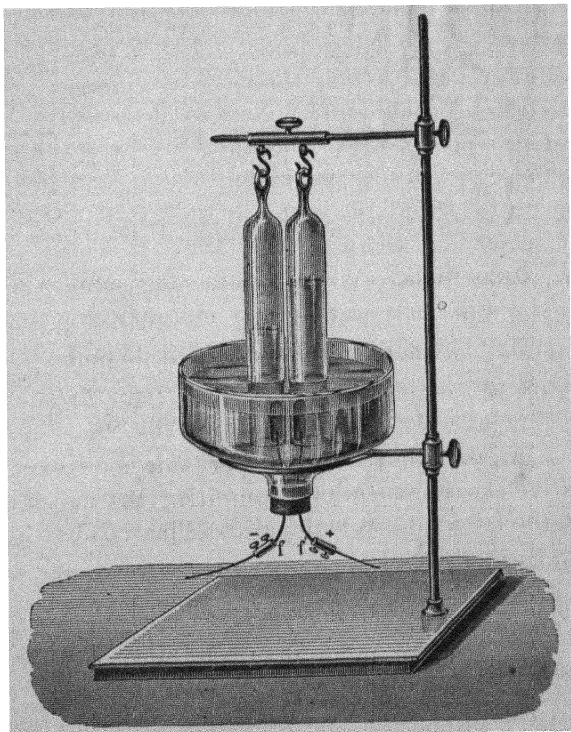


Fig. 27.—Electrolysis of water.

of water is meant the decomposition of water into its elementary constituents, oxygen and hydrogen, by means of a current of electricity.

To perform the experiment take a funnel-shaped glass vessel (Fig. 27) closed at the bottom with a cork, or some other substance which does not conduct electricity, and through which two platinum wires are passed. Pour some water into the vessel and acidulate it with a few drops of either hydrochloric or sulphuric acid to help the electricity to pass more easily through the liquid. Now connect the free outward extremities of the platinum wires with the poles of an electric battery, and immediately gas bubbles are seen to ascend through the water, all along the two free ends of the platinum wires.

To collect these gases, invert two test-tubes of equal size and filled with water, over the platinum wires; the gas bubbles rise in the tubes and displace the water.

After some time it is observed, that in the tube which is inverted over the negative electrode, twice as much gas is evolved as in the other placed over the positive electrode, so that, when the former is full of gas, the latter is only half full.

In order to test the nature of these two gases, take the first tube out of the water, after closing it with the finger. Apply now a glowing splinter of wood to the mouth of the tube: it is not rekindled; but if a burning taper be applied to it, the gas will take fire and burn with a very slightly luminous, bluish flame, while the taper will be put out if it be introduced into the tube. Hence we conclude that this gas is hydrogen.

If we hold a red-hot splinter of wood to the mouth of the second tube, it will burst into flame; this is a sign of the presence of oxygen.

Chemists have never yet been able to get from water anything different from hydrogen and oxygen, nor have they obtained the two gases in any other proportion by volume. Hence they conclude

(a) that water is composed of two different gases, hydrogen and oxygen, and

(b) that these two gases are always obtained from water in the proportion of two volumes of hydrogen for one of oxygen.

(2) **By synthesis.**—Another method of finding the volumetric composition of water is by synthesis (Gr. the putting together), that is, by exploding a mixture of

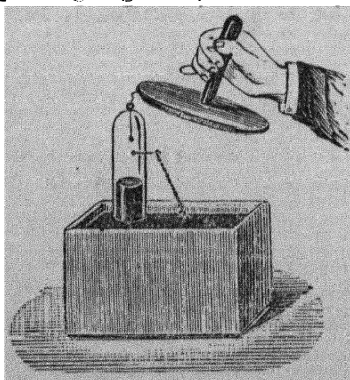


Fig. 28.—Synthesis of water.

exactly two volumes of hydrogen and one of oxygen over mercury, by means of an electric spark (Fig. 28). The entire gaseous mixture is changed into water without any residue, and if we take more of either hydrogen or oxygen than the stated proportion, the excess will remain uncombined.

B. Composition of water by weight.—The composition of water by weight is determined by making a current of purified and dry hydrogen pass over heated black copper oxide.

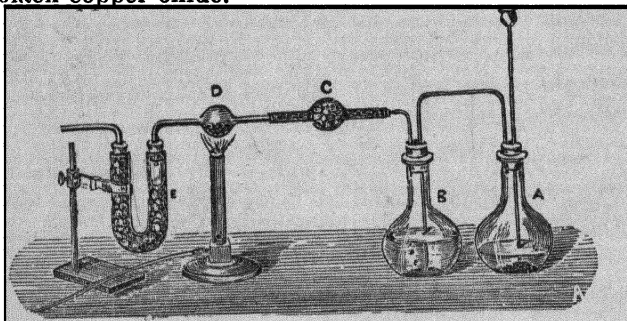
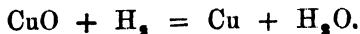


Fig. 29.—To show the manner in which the composition of water by weight is ascertained.

A current of hydrogen, generated in the flask A (Fig. 29) and dried by bubbling through concentrated sulphuric acid in the wash-bottle B, and over lumps of caustic potash in C, is slowly passed over dry black oxide of copper in the bulb D of a difficultly fusible glass tube. This glass tube is connected with a U-shaped tube E, filled with fragments of caustic potash or of calcium chloride, to absorb all the water which will be formed in the reaction. After the air has been expelled from the apparatus by the current of hydrogen, heat is applied to the bulb D. The gas reduces the heated copper oxide, combining with the oxygen to form water, and leaving metallic copper behind :—



It is evident that, after the experiment, the copper oxide tube has become lighter, as it has lost oxygen, while the U-tube has become heavier, as it has gained water. Now the diminution in the weight of the copper oxide gives us the weight of oxygen in the water thus formed, and the difference between the weight of the water and that of the oxygen, gives the weight of the hydrogen present in the water.

Suppose that the copper oxide tube before the experiment was found to weigh 80 grammes; after the experiment it weighs but 72 grammes.

Therefore the difference in weight is due to the loss of oxygen, *viz.* 8 grammes.

On the other hand, the U-tube before the experiment weighed 60 grammes; after the experiment it is found to weigh 69 grammes.

Therefore the increase in weight is due to the water formed, *viz.*, 9 grammes, and consequently the weight of hydrogen must be 9—8 or 1 gramme, that is to say, for every 9 parts by weight of water, there is 1 part of hydrogen and 8 parts of oxygen; or water contains 2 parts by weight of hydrogen to 16 of oxygen; and in this way the composition of water by weight is ascertained.

55. Properties of water.—Water exists in nature in the solid state as ice, snow, hail, etc., in the liquid state, and in the state of gas or vapour in the atmosphere. Its freezing point is 0°C., it reaches its greatest density at 4°, its boiling point is 100°C.

Water has a very remarkable solvent power, there being comparatively few substances which are not to a greater or less extent soluble in water. Thus water dissolves—

(a) **solids**: some are easily soluble, as sugar, alum, copper sulphate; others are sparingly soluble, as gypsum (CaSO_4); a few are insoluble, as chalk, sand, etc.;

(b) **liquids**: water also dissolves other liquids or is dissolved by them; it is said to be *miscible* with them. With some liquids—alcohol, sulphuric acid, etc.—water is miscible in any proportion; with others—oil, ether—in very small quantities;

(c) **gases**: when gases are dissolved in water they are said to be *absorbed*; hence we speak of absorption of gases. The chief gases absorbed by natural water are oxygen, nitrogen, and carbon dioxide.

56. Solution and crystallisation.—Many solid substances are capable of intimately mixing with water; so that, when they are plunged into water, they gradually disappear in it and communicate to it their own properties, as colour, taste.—*Ex.*: sugar, copper sulphate, potassium chromate, alum. The water which thus holds another substance in this peculiar state of liquidity is called a **solution**.

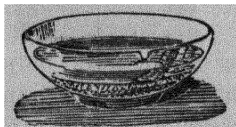


Fig. 30.—Slow evaporation of solutions and formation of crystals.

Remark. For simplicity's sake we speak here of water alone as a solvent; but the same is true of all liquids in general.

When solutions are allowed to evaporate slowly in an open vessel (Fig. 30), the liquid disappears, while the substance dissolved reappears, generally in the

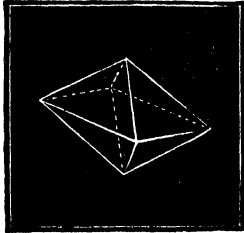


Fig. 31.—Crystal of sodium carbonate.

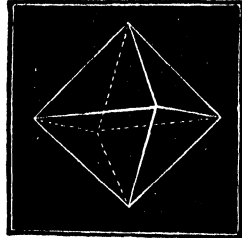


Fig. 32.—Crystal of alum.

form of regular solids limited by plane surfaces and exhibiting geometrical figures. These are called

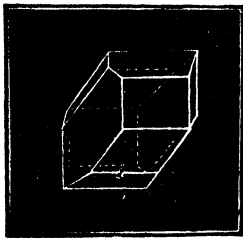


Fig. 33.—Crystal of copper sulphate.

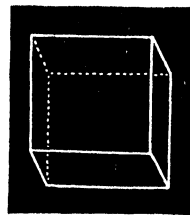


Fig. 34.—Crystal of common salt.

crystals, and the process is called **crystallisation**.

Moreover, the crystalline shape, always the same

for the same substance, is different in different substances (see Figs. 31, 32, 33, 34), so that if, for example, a mixture of alum and copper sulphate solution be allowed to evaporate, the alum will crystallise into regular, white octahedra (Fig. 32), quite different from the blue crystals of copper sulphate (Fig. 33). This fact is sometimes taken advantage of to separate mixed substances from one another, or to purify some salts.

Crystals of salts are usually obtained by this method: it is known as *crystallisation by solution*.

There are however other methods: Thus when metals are liquefied by heat, and are allowed to cool *slowly*, they assume a crystalline form: *crystallisation by fusion*.

Other substances, like iodine, camphor, when made to *sublime*, i.e., pass into vapour directly from the solid state, crystallise on condensing: *crystallisation by sublimation or vapourization*.

Some substances, such as sulphur, can be crystallised in all the three ways.

57. Water of crystallisation.—Efflorescence and deliquescence.—Many salts, when they crystallise out of their solution, combine with a certain fixed quantity of water, which is necessary for the form they assume, and this water is called *water of crystallisation*.

Ex. $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$; $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$; $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$; $\text{K}_2\text{Al}_2(\text{SO}_4)_4 \cdot 24\text{H}_2\text{O}$ (ordinary or white alum).

Some among these salts, when exposed to the atmosphere, part with their *water of crystallisation*, the bright, transparent, glassy crystal becoming opaque and white, and finally crumbling to powder; such crystals are said to *effloresce*, and the phenomenon is called *efflorescence*.

Some others, on the contrary, as potassium carbonate, sodium sulphite, when exposed to the atmosphere, absorb moisture,

becoming damp and even liquid : they are said to *deliquesce*, and the phenomenon is called *deliquescence*.

That the water of crystallisation is essential to the form and sometimes to the colour of the salt is shown in the case of copper sulphate $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$, which becomes white and loses its crystalline form when heated, and becomes blue again when water is poured upon it. Also in the case of cobalt chloride, $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$, which becomes blue on being heated, *i.e.*, when it is deprived of its water of crystallisation. Its original rose colour is restored by the addition of some water. On the other hand, crystals of alum and gypsum fall to powder when they are heated.

58. Impure waters, and how they are purified.--

None of the various forms of water met with in nature, are altogether free from impurities. These may be of two kinds :

- (1) mechanically suspended, or insoluble impurities,
- (2) soluble or dissolved impurities.

The former can be removed either by *decantation*, *i.e.*, by allowing the solid suspended particles to subside and then pouring off the clear liquid, or by mechanical *filtration*, *i.e.*, by making the impure water pass through some porous substance, as charcoal, fine sand or filtering paper, which allows the water only to pass through, but not the solid suspended particles.

The latter kind of impurities, that is, those dissolved in water, cannot be removed by either decantation or filtration, but must be separated by *distillation* or by some chemical reaction.

59. Distillation.—Distillation is the process by which a liquid, when subjected to heat, is transformed

into vapour, the vapour being collected and again condensed into the liquid state. The chief purpose of this operation is to separate liquids from non-volatile substances which they keep in solution, and one of the practical advantages derived from it is to obtain pure water, *i.e.*, water free from dissolved impurities, which cannot be removed by filtration.

Dissolve some common salt in a small quantity of water, which may moreover be made turbid by shaking up with it some chalk or other such substances. Put this salt-water in a retort (Fig. 35), the end of which is introduced into a long-necked glass flask kept cool by a constant flow of fresh water,

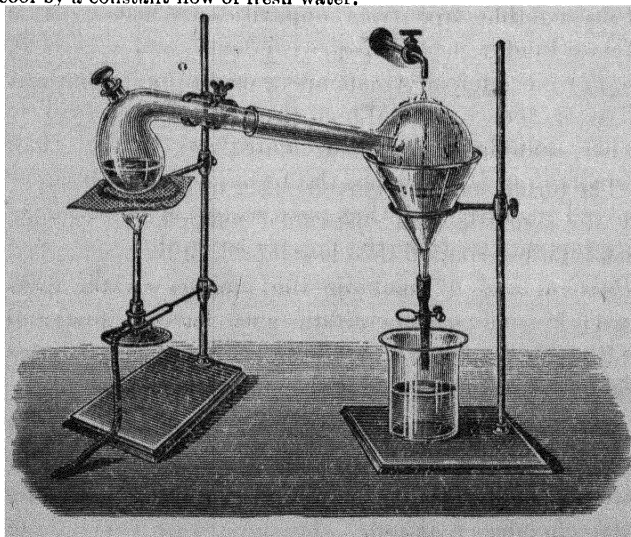


Fig. 35.—Distillation of water.

On heating the retort, the water is changed into steam, and the steam alone can pass over into the condensing flask, while the

dissolved salt, or any other impurity which may be present either dissolved or suspended, will remain behind in the retort.

The water obtained in this way is chemically pure water, and does no longer taste salt.

Something similar to what is done on a small scale in distillation, occurs on a very large scale in nature, that is, in the formation of rain water.

60. Rain water.—Rain or snow water is the purest form of natural waters. Under the influence of the sun's heat the water of rivers, lakes, but chiefly of the sea, is evaporated and ascends into the sky in the form of an invisible vapour of chemically pure water. At a certain height these vapours condense, and appear in the form of clouds, which are carried* by the winds all over the earth. When these clouds become so dense, that the particles of water are heavier than the air, they fall down in the form of rain (or snow in cold regions), and thus pure water is obtained in large quantities from the impure sea-water.

Such water—1° leaves no solid residue when evaporated; 2° gives no precipitate with silver nitrate, as ordinary sea or salt water does; 3° produces no curd when soap-solution is added to it. (Cf. no. 61).

As soon as rain-water falls on to the ground, it dissolves the soluble constituents of that portion of the soil through which it passes, and thus becomes more and more impure, the impurities varying, of course, according to the nature of the soil. In this way are produced the various kinds of hard and

mineral waters. These waters unite to form streams and rivers, and finally reach the ocean from which they were first taken, and which they are thus slowly and gradually filling with impurities.

It is evident that, in the vicinity of towns, water gets contaminated with organic impurities, and is thus rendered unfit for drinking purposes. ✓

61. Hard and soft water.—Water that contains in solution salts of calcium or magnesium or of both, does not readily form a lather when a solution of soap is added to it, but only after a certain quantity of soap has been consumed; such water is said to be **hard**.

If, on the contrary, water lathers as soon as a solution of soap is added to it, it is said to be **soft**, and this water does not consequently contain in solution the above mentioned salts. (Cf. no. 136.)

N.B.—Pure water is necessarily soft, but soft water does not necessarily mean pure water; for it may contain sodium, potassium and ammonium salts which cannot form a curd, i.e., a precipitate with soap, as the compound obtained by the union of soap with any of these salts is itself soluble in water.

62. Temporary and permanent hardness.—Hardness of water is said to be **temporary** or **permanent**, according as it can be removed or not by mere boiling, or by adding lime-water to it. Now, as dissolved carbonates are again precipitated by heat, whilst sulphates or chlorides cannot be thus removed, it is evident that temporary hardness is due to the presence of calcium or magnesium carbonates, whilst

permanent hardness indicates the presence of calcium or magnesium sulphates or chlorides.

Temporary hardness can be removed (1) by boiling the water, (2) by adding lime-water to it.

In order to understand the reason of the process, we must know the following facts :

(a) limestone or chalk (CaCO_3) is *insoluble* in *pure water*,

(b) but is soluble in water containing dissolved carbon dioxide ;

(c) heat easily expels this carbon dioxide, which keeps calcium carbonate in solution, and the calcium carbonate is thus again precipitated,* and can be separated either by decantation (by using a siphon in the case of a large reservoir), or by filtration. It follows that temporary hardness can be removed by merely boiling the water, which consequently becomes soft.

Again (d) lime-water forms an insoluble compound (CaCO_3) with carbon dioxide gas.

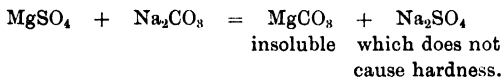
Therefore, if enough lime-water be added to combine exactly with all the carbon dioxide gas which keeps calcium carbonate in solution, the dissolved calcium carbonate reappears in the form of an insoluble precipitate, and the carbon dioxide itself combines with lime-water to form a new quantity of

* The "furring" of kettles and boilers, or the formation of deposits consisting mainly of calcium carbonate, seems to be due to this cause.

calcium carbonate. In this way all the insoluble calcium carbonate is precipitated, and after it has subsided, the clear liquid, free from chalk, may be poured off.

This is the method usually employed when water hard with chalk, is to be softened on a *large scale*.

N.B.—Permanently hard water can be improved by adding to it a solution of sodium carbonate or washing soda, Na_2CO_3 , which precipitates the calcium or magnesium salts in the form of carbonates :—



63. Sea-water.—From what has been said above, it is evident that sea-water is hard water.

For every 100 parts by weight of sea-water, there are about $3\frac{1}{2}$ parts of solid matter, of which more than $2\frac{1}{2}$ consist of sodium chloride only. After sodium chloride or common salt, the substances which are most abundant in sea-water are magnesium chloride and sulphate.

Bromine and iodine compounds are also found in sea-water. Certain sea-plants store them up in their tissues, and from the ashes of these plants, bromine and iodine are usually obtained.

CHAPTER III.

NITROGEN.

Symb. N. At. wt. = 14. *pentatonic r t*

64. Occurrence—Nitrogen exists abundantly and in the *free* state in the atmosphere, of which it forms four-fifths by volume. It also exists in the *combined* state, as in ammonia and in ammoniacal salts, and as nitrates, the chief of these being saltpetre or potassium nitrate KNO_3 , found in India and generally in warm countries; Chili saltpetre or sodium nitrate, NaNO_3 , found in Chili and Peru. Nitrogen is more-over one of the constituents of animal bodies while its compounds are also found in plants.

65. History.—The discovery of oxygen in the air naturally led to the discovery of nitrogen. Lavoisier called this element *azote* (Gr. *a* = not, *zoe* = life), because it is unable to support life. The name of nitrogen was subsequently given to it, when it was found to be a constituent of *nitre* (nitrogen = nitre producer).

66. Preparation.—Nitrogen can be obtained
1° From the air by merely removing the oxygen; or
2° from some compound of nitrogen.

1° *Nitrogen obtained from the air by removing the oxygen* :—

(a) A flat cork, carrying a small dish with a bit of dry phosphorus upon it, is made to float in a basin partly filled with water (Fig. 36), and a bell-jar, open

at the top, is inverted over it. The phosphorus is then lighted by being merely touched with a heated

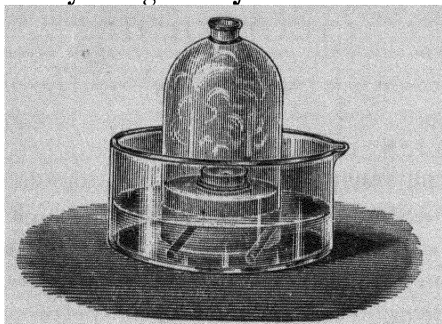


Fig. 36.—Preparation of nitrogen.

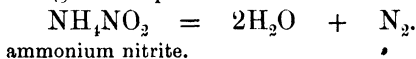
iron wire through the open top of the bell-jar, which is then rapidly closed. The phosphorus burns vividly for a short time so long as there is a sufficient supply of oxygen, while dense white fumes of phosphoric pentoxide, the result of the combustion, fill the bell-jar. These, after some time, become less and less dense, and the water within the bell-jar gradually rises above its outside level. This is because phosphoric pentoxide dissolves in the water, forming phosphoric acid (blue litmus paper is now turned red by the water), and the water rises to fill up as much of the space as was before occupied by the oxygen present in the air. The colourless gas which remains in the bell-jar, is nitrogen, in a more or less pure state.

It is observed that the water occupies about $\frac{1}{5}$ of the original volume of air, and this corresponds to

the quantity of oxygen which was present. Hence it is concluded, that the proportions of oxygen and nitrogen in the air are approximately as 1 : 4.

(b) Another way of getting nitrogen from the air, is to pass slowly a current of air over heated copper. The oxygen combines with the copper to form copper oxide, and the nitrogen is left free.

At 2° Pure nitrogen can also be easily obtained by merely heating *ammonium nitrite*, or more conveniently a mixture of ammonium chloride and either sodium or potassium nitrite, in a glass retort, when the following decompositions occur:—



67. Properties or tests.—Nitrogen is a colourless, odourless, tasteless gas, a little lighter than air and very little soluble in water.

It is neither combustible nor a supporter of combustion.

Animals are quickly suffocated in an atmosphere of pure nitrogen, but only for want of oxygen, nitrogen itself not being directly poisonous.

In these last two properties nitrogen resembles carbon dioxide, but it is easily distinguished from the latter by the fact that it produces no milkiness in lime-water and is neutral to litmus paper.

Nitrogen is remarkable for its inertia, *i.e.*, its weak affinities for other substances, as it unites directly with no other element at the ordinary temperature.

- (1). One of its chief functions in nature is to dilute the oxygen of the atmosphere, as oxygen cannot be breathed pure. ✓

THE AIR.

68. History.—The air was thought to be an element by the Ancients. Its true nature and composition are known since the discovery of oxygen.

69. Composition.—The air is a mixture of several gases, the two most abundant being nitrogen and oxygen:—

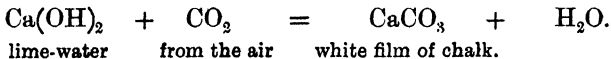
	by volume:		by weight:	
oxygen	21	nitrogen	23	}
nitrogen	79	oxygen	77	

out of every 100 parts.

In addition, a small quantity of carbon dioxide CO_2 , and of vapour of water, is invariably present.

Enough has already been said about oxygen and nitrogen.

Carbon dioxide.—(a) *Its presence proved.*—If some lime-water be exposed in a flat dish in the open air, within about an hour or two its surface will be covered with a white film of chalk, or calcium carbonate, due to the action of the atmospheric carbon dioxide upon lime-water:—



(b) *Its sources.*—These are various; but the greater proportion result from *combustion*, which is constantly going on all over the earth, and from *respiration*, which is a kind of slow combustion.

(c) *Its use in nature.*—Carbon dioxide is necessary for the growth of plants. The green parts of living plants have the power of decomposing this gas under the influence of sunlight. They take up the carbon which serves for the nutrition of the plant, and set free the oxygen which, thus restituted to the atmosphere, goes to support combustion, or the respiration of animals, and is once more transformed into carbon dioxide. In this way a constant equilibrium is kept between the amount of oxygen and carbon dioxide present in the atmosphere.

The decomposition of carbon dioxide by plants and the consequent liberation of oxygen may be illustrated in the following simple manner: A water plant or a bunch of young growing leaves is exposed

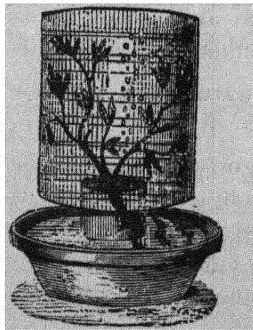


Fig. 36-a Plants decompose carbon dioxide in presence of sunlight.

to direct sunlight (Fig. 36-a) under water containing carbon dioxide. Bubbles of gas will soon form

which collect at the top of the inverted jar, and which, on examination, are found to be oxygen.

Vapour of water.—(a) *Its presence proved.*—If a glass vessel containing a mixture of ice and salt be exposed to the air, small drops of water will be seen to gather on the surface of the glass outside. The moisture of the atmosphere, in the form of steam or vapour, condenses and is deposited, owing to contact with the cold surface of the vessel.—Or, if a measured quantity of sulphuric acid be exposed to the air, after a few days the volume of the liquid will be found to have increased; this is because sulphuric acid has the power of absorbing the vapour of water present in the air.

(b) *Its use.*—Although the quantity of vapour of water in the atmosphere be very small, yet it is necessary for animal and plant life.

70. Air a mixture, not a chemical compound.—This can be proved in the following ways (see no. 7):

(1) If oxygen and nitrogen be mixed in the proportion in which they exist in the air, no heat is evolved, no contraction or diminution in volume takes place, as commonly happens when gases combine chemically in unequal proportions by volume.

(2) Such a mixture does not acquire new properties altogether different from both oxygen and nitrogen.

(3) If air were a chemical compound, it would dissolve in water as such, and consequently after

expelling the dissolved air from water by applying heat, we should find it to have the same percentage composition as before; but, as a matter of fact, we find that about twice as much oxygen dissolves in water as nitrogen.

(4) The relative amounts of oxygen and nitrogen present in the air, are not their combining weights, nor any multiple thereof, as is required in all compounds.

Remark. Although oxygen and nitrogen are present in the atmosphere as a mixture, it is remarkable that the proportion remains almost as constant as if air were a chemical compound.

AMMONIA, NH_3 .

71. Preparation — Ammonia, a gaseous compound of nitrogen and hydrogen, was formerly obtained by distilling horn (hence its former name of *spirit of hartshorn*), bones or any other kind of animal matter.

It is now chiefly obtained as a by-product in the manufacture of coal gas.

The most convenient way of preparing it in the laboratory, is to heat ammonium chloride or sulphate in presence of a stronger base, as soda, or quicklime which is cheaper.

By simply rubbing in the hand a small quantity of sal ammoniac or ammonium chloride with quicklime, the pungent odour of ammonia is perceived.

A mixture of sal ammoniac and excess of quicklime or slaked lime is introduced into a flask (a copper

or iron flask is preferable as glass is liable to crack on heating the mixture). The flask is next provided with a cork, into which is inserted a delivery tube,

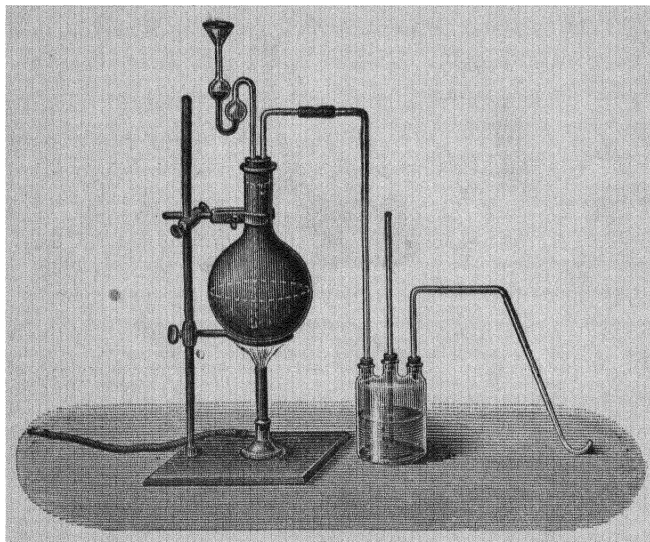


Fig. 37.—Preparation of *liquor ammonia*.

and also a safety tube charged with a very short column of mercury (Fig. 37); gentle heat is applied and the following reaction occurs:—



The ammonia obtained is collected in a Wolff's flask partly filled with distilled water, and thus we have a solution of ammonia, commonly called *liquor ammonia*.

If the gas only is desired, it is collected over mercury, or by upward displacement in a dry flask or test-tube (Fig. 38).*

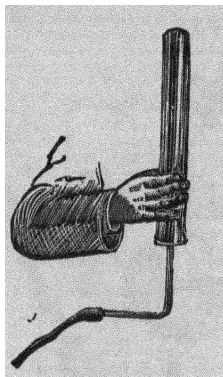


Fig. 38.—Collection of ammonia.

72. Properties.—Ammonia is a colourless gas, with a very pungent odour, bringing tears to the eyes.

* Gases may be collected either over water, or over mercury, or by displacement of air.

(a) Gases, insoluble or only slightly soluble in water, are usually collected over water.

(b) Gases, very soluble in water, are collected over mercury, unless they attack mercury.

(c) Gases, that both dissolve in water and attack mercury, are collected by displacement of air in *dry* jars or test-tubes.

In general, any gas that is *considerably* heavier or lighter than air, can be collected by either downward or upward displacement, according as it is much heavier (Cl, HCl, CO₂, SO₂, H₂S) or lighter (H, NH₃) than air.

It is very soluble in water, which can dissolve about 750 times its volume.

To show this property experimentally, fill a glass bottle A (Fig. 39), with ammonia gas. Close the neck

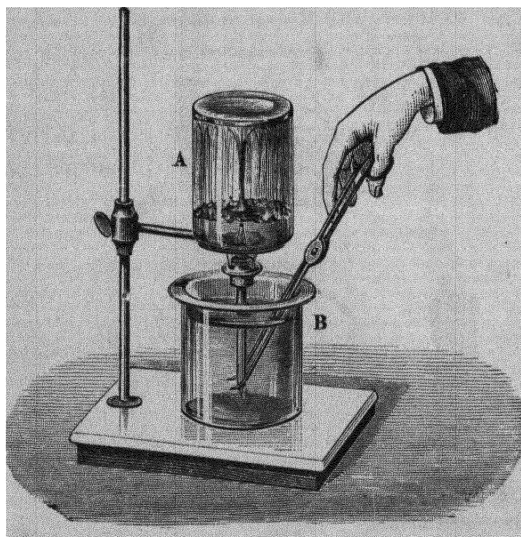


Fig. 39.—To show the great solubility of ammonia in water.

with a cork, and through the cork, pass a glass tube with its outer extremity sealed up. Dip now this closed end of the glass tube into another glass vessel B, filled with water, and break it open under water. The ammonia coming in contact with water dissolves in it, and thus a vacuum tends to be produced. Under the influence of the atmospheric pressure, the water

ascends in the glass tube, slowly at first, but as soon as it reaches the extremity of the tube, it bursts forth into the upper glass vessel in a beautiful jet, the gas being now exposed to a great surface of water in which it can dissolve rapidly. If the water has previously been coloured red with an acidulated litmus solution, the jet of water will turn blue.

Ammonia is readily absorbed by freshly-burnt

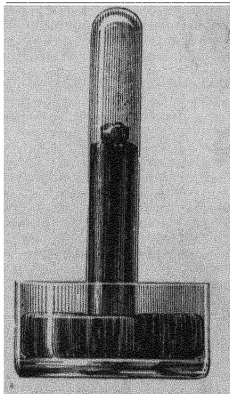


Fig. 40.—Ammonia absorbed by charcoal.

charcoal. Into a large test-tube filled with mercury and inverted over a trough of mercury, introduce dry ammonia gas. Heat a piece of charcoal to redness, and rapidly plunge it into the mercury.* When it is extinguished, allow it to ascend into the test-tube; it will be seen to rest on the surface of the mercury in the tube (Fig. 40). Presently the column of mercury is seen to rise in the test-tube, as the ammonia gas is rapidly absorbed by the charcoal.

Ammonia has a strongly alkaline reaction. It combines with water forming ammonium hydroxide NH_4OH , which is a strong base like the alkalis KOH .

* This is done in order to expel from the charcoal the gas or air which it may already have absorbed, and to prevent it from absorbing new gases.

and NaOH. As it very easily evolves the gas, it is called the volatile alkali, whilst the other two are called the fixed alkalis.

73. Uses.—Ammonia is used as a chemical reagent. In the liquid form, *i.e.*, condensed to liquefaction, it is largely used in the manufacture of artificial ice.

74. Tests.—A characteristic odour; white fumes with HCl; a deep blue colouration with copper salts.

NITRIC ACID

or Hydrogen nitrate, HNO_3 .

75. Occurrence.—Nitric acid, formerly called *aqua*

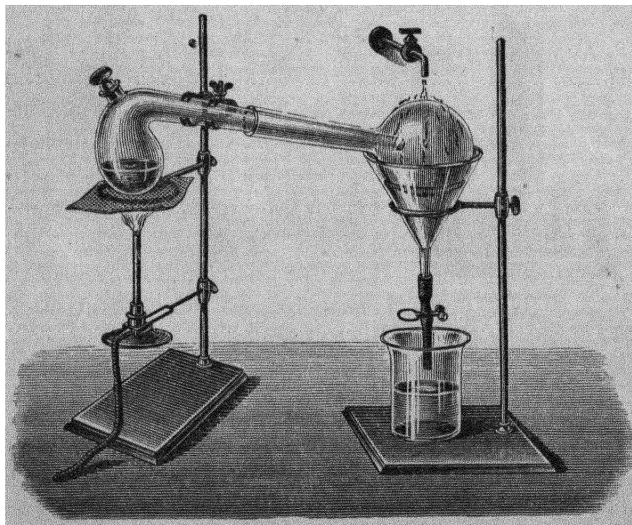


Fig. 41.—Preparation of nitric acid.

fortis, *i.e.*, *strong water*, from its power of dissolving

nearly all the metals, exists in very small quantities in the atmosphere.

76. Preparation.—Nitric acid is prepared in the laboratory by gently heating about equal weights of sulphuric acid and potassium or sodium nitrate in a glass retort. The end of the retort is introduced into the neck of a flask kept cool by immersion in a trough of water, or by allowing cold water to flow upon it from a tap (as in Fig. 41). Nitric acid is formed by the reaction; it distils over and condenses to a liquid in the cool receiver. The reaction is represented by the following equation:—

$$\begin{array}{l} \text{KNO}_3 + \text{H}_2\text{SO}_4 = \text{KHSO}_4 + \text{HNO}_3 \\ \text{or } \text{NaNO}_3 + \text{H}_2\text{SO}_4 = \text{NaHSO}_4 + \text{HNO}_3 \end{array}$$

(an acid salt) potassium bi-
nitrate salt part
or nitrate salt part

On the large scale nitric acid is manufactured by heating the above substances in cast-iron vessels. Then under the influence of a greater heat, the following reaction takes place:—



i.e., for the same quantity of H_2SO_4 , we get twice as much of nitric acid.

77. Properties.—Nitric acid, when pure, is a colourless fuming liquid. It is slowly decomposed by light—hence the yellowish tint it assumes, owing to the presence of some oxides of nitrogen resulting from the decomposition.

It is one of the most powerful acids; it dissolves all the metals, except platinum and gold, and neutralizes bases forming *nitrates* in the greater number of cases.

3. It is a very corrosive liquid, *i.e.*, it attacks and eats away substances with which it comes into contact.

4. It destroys vegetable colours, and causes a bright yellow stain on such substances as silk, wool, the skin, etc.

5. It is a powerful oxidizing* agent, as it easily gives up part of its oxygen. Thus, when strong nitric acid is poured upon red-hot charcoal, the latter bursts into flame.

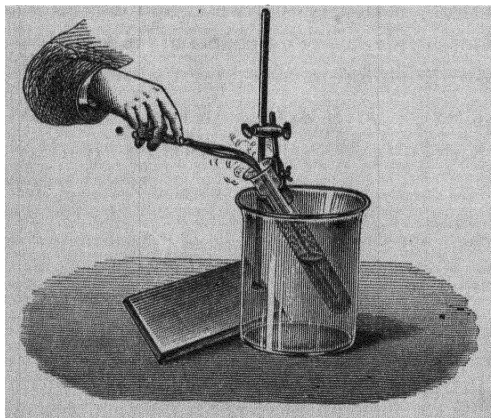
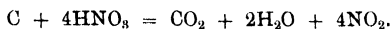


Fig. 42.—Red-hot charcoal burns in strong nitric acid.

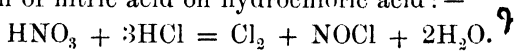
Plunge the end of a stick of charcoal, heated to redness, into strong fuming nitric acid in a wide test-tube (Fig. 42); the charcoal

* Substances that readily part with their oxygen—such as nitric acid, potassium chlorate, black copper oxide etc.—are termed *oxidizing agents*, whereas substances such as hydrogen, carbon (cf. no. 32), carbon monoxide, that easily take up oxygen at the expense of the compounds containing it, are termed *reducing agents*.

will be seen to burn, the acid providing the oxygen necessary for the combustion. The change that takes place is the following :—

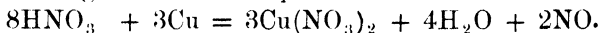


A mixture of HNO_3 and HCl is called *aqua regia*, (i.e., *royal water*), because this mixture is able to dissolve gold, the *king of metals*. Its activity depends upon the fact that free chlorine is evolved by the action of nitric acid on hydrochloric acid :—



The chlorine disengaged combines with gold forming gold chloride. The action of *aqua regia* on platinum is explained in the same manner.

When ordinary HNO_3 is poured upon copper, the following reaction takes place :—



NOTE.—The fumes of NO are colourless, but as soon as they are brought into the air, they are oxidized to brown fumes of NO_2 , of a characteristic odour.

With silver the reaction may in the simplest manner be thus represented :—



The reaction of nitric acid upon metals is, however, generally more complicated.

78 Uses.—Nitric acid is used in the manufacture of sulphuric acid, and to dissolve metals in the preparation of nitrates, and also to engrave upon copper.

In this latter case, the desired figure is drawn with the point of a needle upon a plate of metallic copper covered with a thin layer of varnish or wax. The nitric acid acts upon the copper in those

places only where its surface has been freely exposed, and corrodes it, while it has no action on the wax. The wax is then washed away and the figure appears engraved on the copper.

CHAPTER IV.

CHLORINE.

Symb. Cl. At. wt. = 35.5. *mo*

79. Occurrence.—Chlorine is *never* found *free* in nature. *Combined* it occurs as chloride of sodium NaCl, of potassium KCl, and of magnesium $MgCl_2$, both in the solid state and dissolved, chiefly in sea-water.

80. History.—Chlorine was discovered in 1774, by Scheele, a Swedish chemist. The name *chlorine* (Gr. *chloros* - green) has been given to it on account of its greenish appearance.

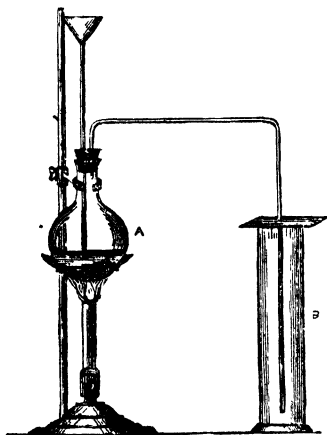


Fig. 43.—Collection of chlorine.

81. Preparation.—Chlorine may conveniently be prepared in two ways:

(1) By gently heating a mixture of *manganese dioxide* and *hydrochloric acid* in a flask, provided with a safety and a delivery tube. The heavy gas may be collected directly by simply introducing the delivery tube into a dry jar

(Fig. 43) till it nearly reaches the bottom; or, if it is to be freed from traces of acid, it is first made to pass through warm water (Fig. 44):—

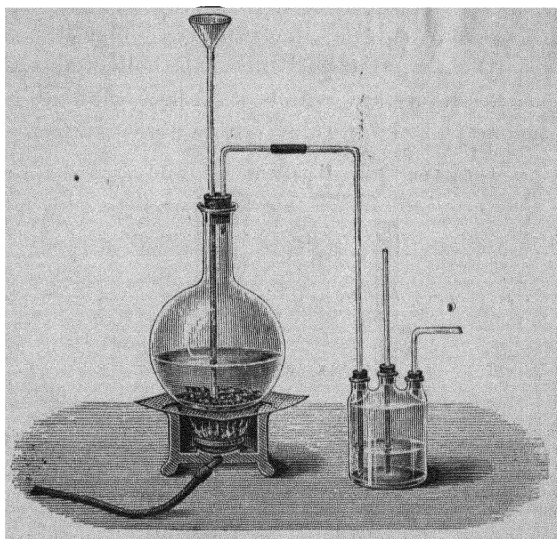
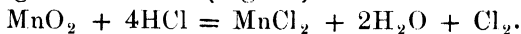
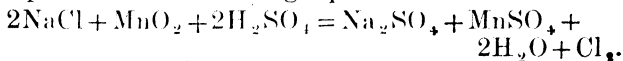


Fig. 44.—Chlorine purified.

(2) By heating a mixture of *common salt*, *manganese dioxide* and *sulphuric acid*. Here two reactions take place: in the first stage hydrochloric acid is formed, from NaCl and H_2SO_4 (see preparation of HCl), and then the HCl thus produced, reacts upon MnO_2 as in the first case; these two reactions are comprised in the following equation:—



82 Properties.—Chlorine is a yellowish-green gas, of a very pungent and irritating odour. 1) When breathed in considerable quantities, it acts as a poison, causing spitting of blood and even death.

2) It is greatly soluble in ordinary cold water, and readily attacks mercury; hence it is collected either over *warm* water, in which it is less soluble, or by displacement of air in a dry jar, as above stated.

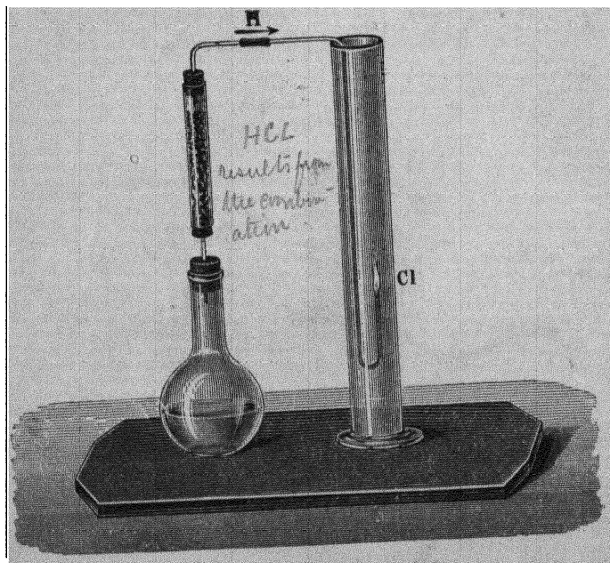


Fig. 45.—Hydrogen burns in chlorine.

It is a heavy gas, *viz.*, $\frac{35.5}{14.44}$ times heavier than air, and 35.5 times heavier than hydrogen.

Chlorine has a great affinity for hydrogen, and very readily combines with it, but either heat or light is necessary to start the combination.

Thus a mixture of the two gases may be kept in *the dark* without combination taking place; in diffused daylight, combination takes place slowly; while the gases combine rapidly and with explosion if a lighted taper be applied to the mixture, or if the bottle which contains them, be merely exposed to the direct light of the sun or of burning magnesium. The product of the combination is HCl .

When a jet of hydrogen is lit and introduced into a jar of chlorine, the jet continues to burn (Fig. 45), HCl being formed.

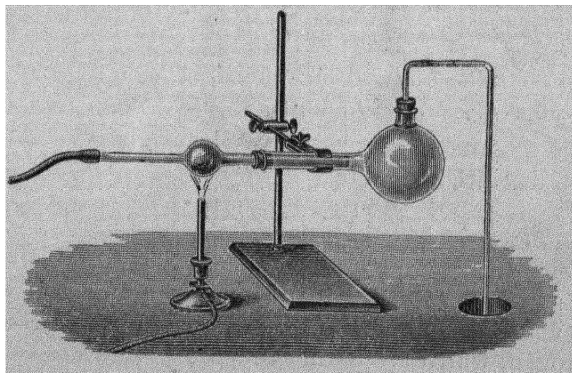


Fig. 46. —Formation of aluminium chloride by the action of chlorine on heated aluminium.

Owing to its affinity for hydrogen, chlorine supports the combustion of a taper though it is not itself

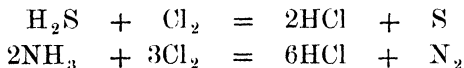
inflammable. The taper burns with a smoky flame, the hydrogen of the wax forming HCl with the chlorine, and carbon being set free.

When a stream of chlorine gas is made to pass over finely divided antimony (Cf. no. 17) or arsenic, they take fire and burn with incandescence forming their respective chlorides.

Likewise on passing dry chlorine over strongly heated aluminium powder, the aluminium begins to glow forming dense, white fumes of aluminium chloride (Fig. 46), which may be condensed in an appropriate receiver as a crystalline solid. Under the same conditions copper also forms its chloride.

83. Uses.—Chlorine is used (1) as a *disinfectant* and (2) as a *bleaching agent*.

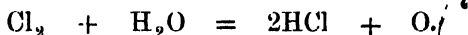
(1) On account of its chemical activity and on account of its great affinity for hydrogen, chlorine destroys noxious germs, as well as offensive gases, such as H_2S , NH_3 , etc.



i. e., instead of the above offensive gas H_2S , and pungent-smelling NH_3 , there remain sulphur and nitrogen, which are harmless.

(2) The *bleaching* action of chlorine (*i. e.*, its power of destroying vegetable colours) is also due to the affinity of chlorine for hydrogen. It is known that the substance to be bleached must be moist.

Now chlorine combines with the hydrogen of the water, and the oxygen which is thus set free, forms with the colouring matter other colourless compounds, and in this way the substance is decolourised or bleached:—



But free chlorine is somewhat troublesome to deal with; and as its action is rather violent and might destroy the tissues, it is more commonly used in the form of a compound, calcium chloro-hypochlorite, which gives up chlorine but slowly. To this compound, which exists in the form of a white powder, the name of *bleaching powder* is given (see bleaching powder, no. 148). *vide pp 163, no 148*

HYDROCHLORIC ACID

or hydrogen chloride, HCl.

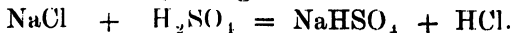
84. Occurrence.—Hydrochloric acid (formerly called *muratic acid*) is a gas, and exists among the gaseous products evolved from volcanoes.

It is the only compound known of hydrogen and chlorine.

85. Preparation.—Hydrochloric acid gas is easily obtained from almost any chloride, by the action of sulphuric acid. Usually sodium chloride is employed on account of its cheapness.

Lumps of previously fused common salt—(with powdered salt the action would be too violent on

account of the great surface exposed to the acid)—are introduced into a flask, fitted with a funnel or safety tube and a delivery tube. Sulphuric acid is then poured down upon the NaCl through the funnel tube. Hydrochloric acid gas is given off:—



But as it is very soluble in water, it must be collected by displacement of air in the same way as chlorine. It may also be collected over mercury.

A solution of the gas is prepared, by making it pass through a series of Wolff's bottles partly filled with water.

86. Properties. 1) Hydrochloric acid is a colourless gas with a pungent odour; 2) it is neither inflammable

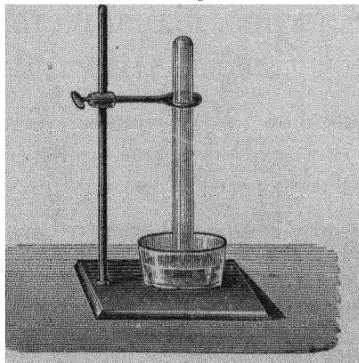


Fig. 47.—To show that HCl is soluble in water.

3) It is very soluble in water, one volume of water being able to dissolve about 500 volumes of HCl at 0°C.

To illustrate this solubility by an easy experiment, put some common salt in a flask with a long neck, or in a stout test-tube. Upon it pour sulphuric acid and, if necessary, heat gently. When the flask or tube is full of HCl gas, invert it over a basin filled with water, (as shown in Fig. 47), when the water will be seen to ascend rapidly in the tube. The reason of the phenomenon is the same as that given for ammonia (72). Moreover the same arrangement may be made here as in Fig. 39.

4) Hydrochloric acid is a very strong acid, and intensely reddens litmus paper. 5) It attacks and dissolves many metals, as K, Na, Zn, Fe, Al, Sn, etc., forming corresponding chlorides.

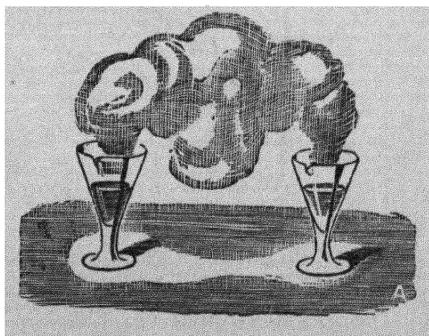
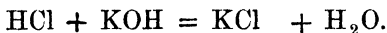
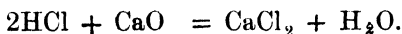


Fig. 48.—Combination of HCl with NH_3 .

6) It also decomposes metallic oxides and hydroxides, and it neutralizes bases, with formation of chlorides

and water, as :



On coming in contact with ammonia gas, the acid forms dense, white fumes of ammonium chloride (Fig. 48).

A solution of pure hydrochloric acid gas is colourless ; often, however, it is coloured yellow, this yellow tint being due to impurities, chiefly iron ; such acid is known as *commercial hydrochloric acid*.

87. Uses.—The chief use of HCl is in the preparation of chlorine and bleaching powder.

CHAPTER V.

SULPHUR.

Symb. S. At. wt. = 32. *Diatomic*

88. Occurrence.—Sulphur exists abundantly in nature, both *free* and *combined*. Free, it is chiefly found in volcanic countries, among the products emitted from volcanoes. Combined, it exists in the form of sulphides and sulphates. The chief *natural sulphides* are those of iron and copper, FeS., CuS, called pyrites,—of lead, PbS, known as galena,—of zinc, ZnS, called blende,—of mercury, HgS, whose common name is cinnabar.

The chief *natural sulphates* are those of calcium, CaSO_4 or gypsum,—of magnesium, MgSO_4 or Epsom salt,—and of sodium, Na_2SO_4 or Glauber's salt.

Art. 89. Extraction.—The greater part of sulphur comes from volcanic countries as Sicily and Iceland. As the

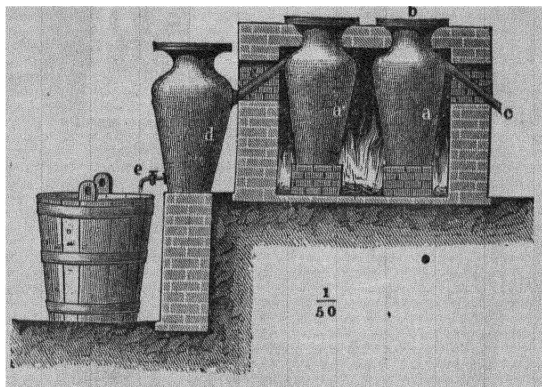


Fig. 49.—Extraction of sulphur.

sulphur is mixed with earthy impurities, it is roughly distilled in earthenware retorts (Fig. 49), the vapour of sulphur being condensed in receivers of the same material.

To purify it still more, it is again distilled in iron retorts, and the vapour is condensed in large brick chambers (Fig. 50). At first the vapour condenses on the cool walls in the form of a fine powder, which is known as *flowers of sulphur*. But when the walls of the chambers become heated above the melting point of sulphur, the latter condenses into a liquid, and

is made to flow in cylindrical moulds in which it

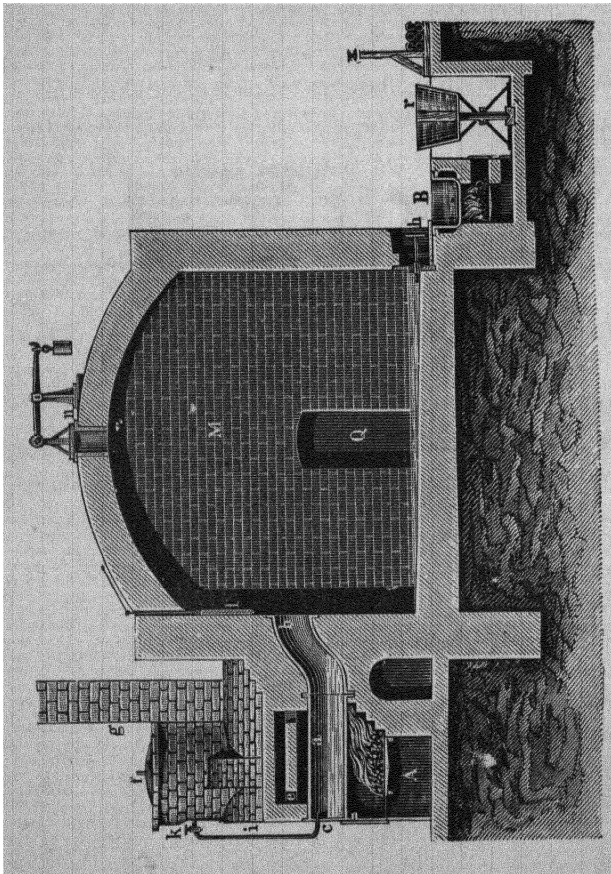


Fig. 50. — Refining of sulphur.

solidifies in the form of sticks; this is *roll sulphur* or *brimstone*.

90. Properties. 1) Sulphur is a brittle solid, of a lemon-yellow colour, 2) a bad conductor of heat and electricity, 3) insoluble in water (therefore tasteless), but soluble in carbon disulphide.

When sulphur is heated in the air, it first melts, its colour becoming brown and dark; then it boils, and the vapour given off burns in contact with air or oxygen. Sulphur burns with a pale-blue flame, producing sulphur dioxide, a gas of a suffocating odour.

4) Sulphur possesses a great affinity for many other elements, chiefly for metals such as Ag, Cu, Fe, etc., forming sulphides with them.

5) Sulphur exists in two *allotropic* forms: ^{• allotropy} in ^{in 3 forms} ~~in 2 forms~~ ^{of m.}

One form consists of octohedral crystals. These crystals are either found free in nature, or may be obtained from ordinary sulphur by dissolving it in carbon disulphide, and allowing the solution to evaporate.

The other variety consists of long needle-shaped crystals, obtained by melting sulphur in a crucible, and allowing it to cool. When the surface has solidified, it is broken, and the remaining liquid sulphur is poured out; the walls of the crucible will then be seen to be covered with the above-

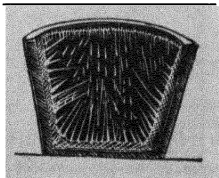


Fig. 51.—Needle-shaped crystals of sulphur.

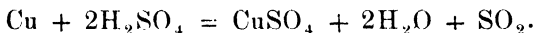
mentioned crystals (Fig. 51).

Allotropy is the property which certain elements, as O, S, C, P and Si, possess of existing in two or more distinct forms. The difference is probably due to the variation in the number of atoms in the molecule, or to a different arrangement of them.

91. Uses.—Sulphur is used in the manufacture of sulphuric acid, in the manufacture of matches (to help the wood to take fire), and of gunpowder, which is a mixture of sulphur, charcoal and nitre.

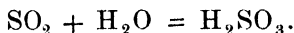
SULPHUR DIOXIDE, SO_2 .

92. Properties and uses.—Sulphur dioxide is formed when sulphur is burning. To prepare pure sulphur dioxide, mercury or copper is heated with sulphuric acid :—



It is a colourless gas with a pungent, suffocating odour like the smell of burning matches. It is neither combustible nor a supporter of combustion.

It is soluble in water, forming with it sulphurous acid :—



This solution, which possesses powerful bleaching properties, is used instead of chlorine to bleach vegetable substances, silk, wool, straw, etc., which would be destroyed by chlorine.*

* Some authors are of opinion that the bleaching action of sulphur dioxide in presence of water is due to its absorption of

Its bleaching action may easily be shown in the following way:—Place a small dish with sulphur near a bunch of moistened coloured flowers; light the sulphur and cover the whole with a bell-jar to collect the sulphur dioxide (Fig. 52). After a short time the flowers begin to lose their colour.

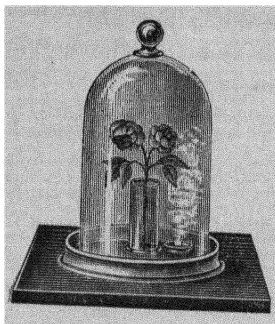
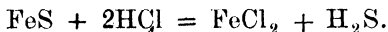


Fig. 52.—Flowers bleached by sulphur dioxide.

It is also used to disinfect sick-rooms, and to destroy noxious insects as well as offensive odours.

SULPHURETTED HYDROGEN or hydrogen sulphide, H_2S . *Amel*

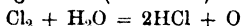
93. Preparation, Properties, and Uses.—Hydrogen sulphide is obtained by pouring dilute hydrochloric acid upon iron sulphide:—



oxygen from the water and consequent liberation of *hydrogen* which *reduces* the colouring matter:—



It will be remembered that chlorine also decomposes water, but, unlike sulphur dioxide, it seizes the hydrogen, leaving *oxygen* free to *oxidize* the colouring matter (cf. No. 83):—



It is probable that in some cases the bleaching action of sulphur dioxide is due to its direct combination with the colouring matter.

It is a colourless gas with a strong, disagreeable smell of rotten eggs. It is combustible in a free supply of air, but is a non-supporter of combustion.

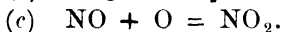
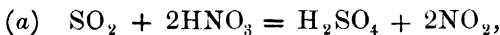
It is easily soluble in water; it is much used in the laboratory to detect metals in solution, as this acid forms a characteristic insoluble precipitate, a sulphide, with many metals.

SULPHURIC ACID

or hydrogen sulphate, H_2SO_4 .

94. Manufacture.—This acid, of very great importance both in chemistry and the arts, is manufactured on a very large scale by passing sulphur dioxide gas, vapours of nitric acid, steam and air into large leaden chambers, and allowing them to mix freely.

The reaction is somewhat complex, but can be expressed in the simplest manner by the following equations:—



Again the NO_2 thus formed in equation (c), in presence of a new quantity of SO_2 and steam continually supplied, reacts as in equation (b), producing sulphuric acid and NO . This gas NO (for $N_2O_2 =$ nitrogen dioxide), in presence of air, is again oxidized to NO_2 , and thus the reactions (b) and (c) may go on indefinitely. *(Realted)*

The sulphuric acid formed, dissolves in the water at the bottom of the chambers, and very dilute sulphuric acid is obtained. This is next evaporated and subsequently concentrated in leaden pans till a certain point is reached when it begins to attack lead. After this it is still further concentrated by being heated in glass or platinum vessels, which are not attacked by the acid.

To prepare sulphuric acid as a class experiment, the following simple apparatus (Fig. 53) may be fitted up. A is a large flask with a cork bored for five tubes. One of the four long tubes is connected with a smaller flask a, in which sulphur dioxide is being generated

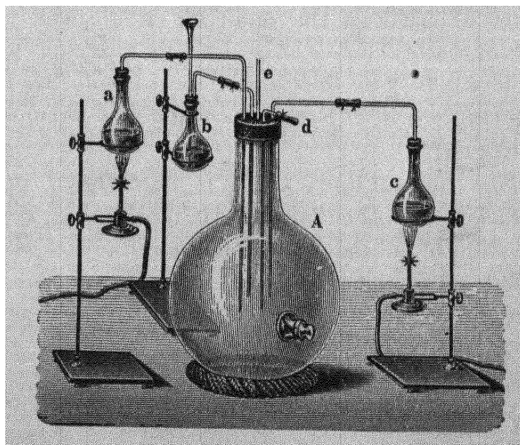


Fig. 53.—Preparation of sulphuric acid.

by the action of concentrated sulphuric acid on copper turnings, and the application of heat. Another tube receives the nitric oxide formed in the flask b, by the action of nitric acid on copper turnings. (For convenience sake the above-mentioned gases may be prepared beforehand and stored in india-rubber bags.) A third tube serves

to convey steam from another small flask c, where water is kept boiling. Air or oxygen is supplied through the tube d, and lastly, the short tube e provides an exit for waste gases. The chemical reaction of these various gases is as described above.

95. Properties and uses.—Sulphuric acid, when pure, is a colourless, heavy (density = 1.84), oily-looking liquid: hence its name of *oil of vitriol*.

1. It has a very great attraction for water: hence it chars organic substances, as wood, sugar, etc., by removing from them the elements oxygen and hydrogen, in the proportion to yield water, and setting free the carbon.

2. When exposed in the open air it absorbs much vapour of water: hence it is used as a drying agent by chemists, *i.e.*, to dry gases by causing them to bubble through it. (See preparation and purification of hydrogen).

3. Besides this, sulphuric acid is much used in the preparation of other acids—HCl, HNO₃; in the preparation of sulphates; in the manufacture of the so-called soda Na₂CO₃, and of soap; it is also used for dyeing, calico-printing and in setting up voltaic batteries, etc.

96. Some common sulphates—Hydrogen sulphate, H₂SO₄, acts chemically upon many metals, metallic oxides and hydroxides (or bases), the hydrogen of the acid being replaced by the metallic radical, and a salt (sulphate of the metal) being formed. Thus,

with Zn or ZnO it forms ZnSO₄=zinc sulphate or white vitriol;
 ,, Fe ,, FeO ,, ,, FeSO₄=iron sulphate or green vitriol;

with Cu or CuO it forms CuSO_4 = copper sulphate or blue vitriol ;
 (heated)
 ,, Mg ,, MgO ,, ,, MgSO_4 = magnesium sulphate or Epsom
 [salt ;
 ,, Na ,, NaOH ,, ,, Na_2SO_4 = sodium sulphate or Glauber's
 [salt ;
 ,, Ca ,, Ca(OH)_2 ,, ,, CaSO_4 = calcium sulphate or gypsum ;
 ,, Ba ,, Ba(OH)_2 ,, ,, BaSO_4 = barium sulphate or heavy spar.

CHAPTER VI.

CARBON.

Symb. C. At. wt. = 12. *Tetratomic*

97. Occurrence.—Carbon, *free*, and still more in the *combined* state, is very widely distributed in nature.

1° *Free*, it exists (a) as *diamond*, its purest and rarest variety, (b) as *graphite*, and (c) as very impure *carbon* in some varieties of coal.

2° *Combined*, it is found both in the mineral or inorganic, and in the living or organic kingdom.

In the inorganic kingdom
 carbon is found { as carbon dioxide CO_2 , in
 the atmosphere, and
 dissolved in water ;
 as carbonates in rocks
 (limestone, marble, etc.)

In the organic kingdom
 carbon is found { as one of the necessary
 constituents of vegetable
 and animal tissues ;
 as one of the components
 in most of the products
 of plant and animal life.

98. Varieties of carbon.—Carbon is found crystallised and amorphous. There are two kinds of crystallised carbon: diamond and graphite. The amorphous variety, *i.e.*, that which does not exhibit a crystalline structure, comprises wood-charcoal, lamp-black or soot, and the various kinds of coal.

(1) **Diamond** (*a*) is the hardest substance known; it scratches glass or any other substance, but is scratched by none. It can be polished only by being rubbed with its own dust.

(*b*) It possesses the property of refracting light in a high degree.

(*c*) It is, very rare and has not yet been reproduced artificially.*

Its *uses* depend upon the above properties:

(*a*) On account of its remarkable hardness diamond is used by glaziers to cut glass; it is also used to perforate rocks in constructing tunnels.

(*b*) On account of its great brilliancy, and

(*c*) on account of its rarity, to which may be added the difficulty of reproducing it artificially, it

* In his remarkable monograph of the allotropic forms of carbon, Dr. Moissan of Paris describes the manner in which he has lately obtained microscopic samples of the several varieties of diamond found in nature, under the combined action of a very high temperature and pressure upon carbon. He also shows that the supposed microscopic diamond crystals of Despretz (1849), and of Hannay (1880), were not diamond at all, and that Marsden (1880-1881) may have obtained at most black diamond or *carbonado*, and that but accidentally. (*Annales de Chimie et de Physique*, Août, 1896.)

is very highly estimated, and used as a gem or precious stone.

When diamond is highly heated in oxygen, it burns away, nothing but carbon dioxide being formed.

(2) **Graphite** is also called *plumbago* or *black lead* (*plumbum* = lead), because it was formerly believed to contain lead; but it contains no trace of it.

1) It has a peculiar, metallic-gray, lead-like lustre. It has a great power of conducting heat and electricity. It is very soft, and leaves a mark on paper when rubbed over it.

2) It is used to diminish friction in machinery, instead of oil, over which it has the distinct advantage of not gathering dust so easily, and of not being decomposed by heat.

3) It is used in the manufacture of the so-called lead-pencils.

(3) **Charcoal.**—A. *Preparation.*—Charcoal (or carbon in an impure state) is prepared by heating wood in a closed vessel, allowing only a small opening for the escape of gases or vapours formed by the decomposition of the wood. It is from these volatile products that alcohol of wood, or *methyated spirit*, is prepared.

Exp. Heat a splinter of white wood in a test-tube; after some time, the wood is seen to become black or charred, while dense fumes or gases are given off,

which will burn at the extremity of the tube, if a lighted match be applied to them.

Charcoal is prepared on a large scale by stacking up piles of wood and covering them with earth or peat, care however

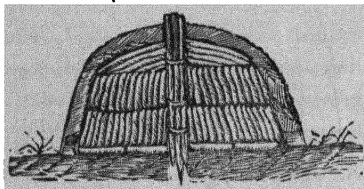


Fig. 54.—Section of a mound.

being taken to leave a vertical opening or chimney in the centre of the stack, which may serve to introduce inflammable matter and to regulate the admission of air.

Fig. 54 represents the section of such a stack or mound, and Fig. 55 the complete mound ready for

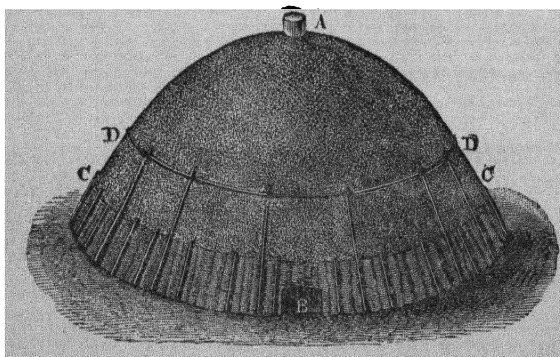


Fig. 55.—Preparation of wood charcoal.

lighting. On setting fire to the inflammable matter in the centre of the mound, a portion only of the wood

burns owing to the limited supply of air, but the heat thus produced converts the remainder into charcoal. This is known as *wood charcoal*.

B. *Properties and uses*.—The uses of charcoal are based upon the following *properties* :

- (a) Charcoal absorbs and condenses gases within its pores, and consequently destroys dangerous and fetid odours.

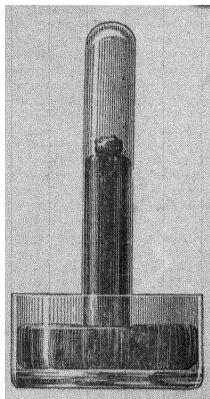


Fig. 56.—Charcoal absorbs gases.

Exp. 1. A test-tube, filled with dry ammonia and inverted over mercury, may be freed from the gas by introducing into the test-tube a piece of freshly burnt charcoal (Fig. 56).

Exp. 2. Pour a few drops of a solution of ammonia, or of sulphuretted hydrogen, in a glass flask, and shake it up with finely powdered charcoal; after some time, all odour has disappeared. On heating the flask, it will reappear once more.

- (b) Charcoal absorbs colouring matters from solutions.

Exp. Shake up in a flask with powdered charcoal some water previously coloured with litmus, aniline, logwood, or any other colouring substance. After some time pass the liquid through filter-paper, to

remove the charcoal, and the water will be as clear as before the experiment.

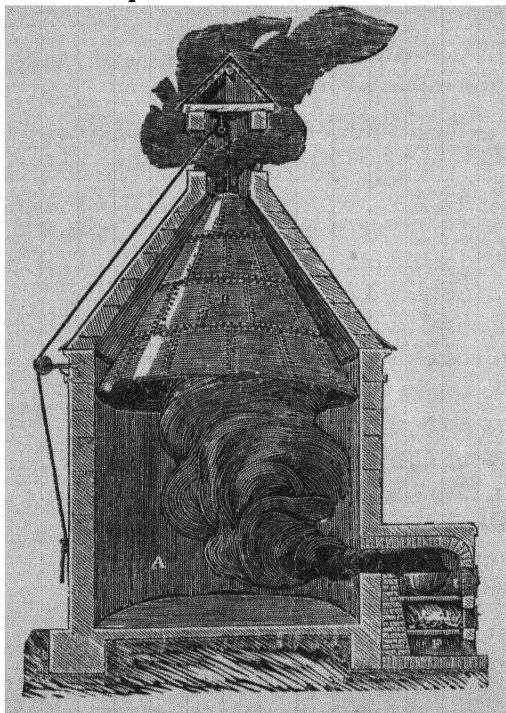


Fig. 57.—Preparation of lamp-black.

(c) Charcoal keeps back suspended solid particles.

Exp. Pour some muddy water over a charcoal filter; clear water is obtained.

Corresponding to the above properties of charcoal, are its *uses* :

(a) as a disinfectant and antiseptic (*i.e.*, to prevent putrefaction) in hospitals, and in the passage of vitiated air from sewers ;

(b) to decolourize certain substances, particularly to refine sugar ;

(c) for water filters, to separate suspended impurities and other objectionable organic matters from water.

‡Charcoal is also used in the manufacture of gunpowder.

Animal charcoal (*bone-black* or *ivory-black*) is made by heating bones and other animal substances in closed vessels.

Lamp-black or *soot*.—Lamp-black is produced by the incomplete combustion of organic substances rich in carbon. To prepare it, resin, tallow, turpentine or any such substance is burned in a limited supply of air, and the smoke is collected and condensed in special chambers (Fig. 57).

It is used for printing-ink and other black pigments.

The name of *soot* is more commonly given to that black substance which is deposited from smoke in chimneys or around fire-places.

Coal.—Coal is a form of carbon less pure than wood charcoal. It is found in vast layers in various countries of the world.

Its origin is commonly ascribed to the slow decay of plants, which at a remote period existed upon the

earth. These plants have in course of time been buried and slowly transformed into coal, under the influence of various agencies, such as moisture, heat and pressure.

That this has been the case, and that coal is but the remains of former vegetation, is proved by the fact that if we examine a thin slice of coal under a magnifying glass, we may observe in it distinct marks of plants. Impressions of ferns and other plants are sometimes visible to the naked eye. In coal pits we come across masses of sandstone and hardened clay (Figs. 57 *a* and *b*), where similar imprints are seen.

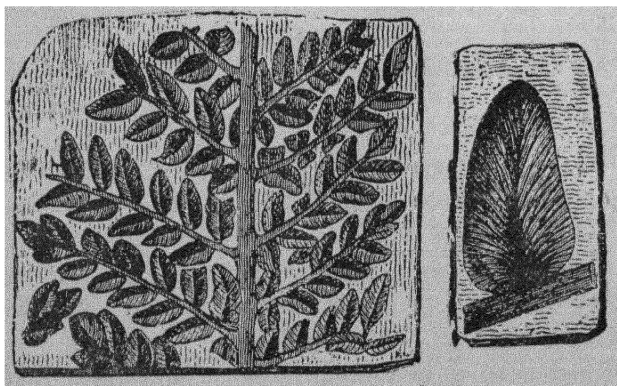


Fig. 57-*a*.—Imprint of plant in sandstone.

Fig. 57-*b*.—A leaf of the same plant magnified.

There are several kinds of coal, some better fitted for heating purposes, others for the manufacture of

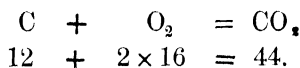
coal gas, according as they contain a greater or smaller percentage of carbon.

Remark.—Diamond, graphite and amorphous carbon are chemically the same substance: they are *allotropic* forms or different modifications of the same element. This can be proved in the following way:

(1) When either diamond, or graphite, or pure amorphous carbon is burnt in a vessel filled with pure oxygen, nothing but CO_2 is formed.

(2) Not only is CO_2 alone produced, but also for the same weight of each of the above three substances, the same weight of CO_2 is produced.

Ex. gr. If we heat 12 grammes of diamond, or 12 grammes of pure graphite, or of pure carbon in a closed vessel filled with pure oxygen, they will burn and leave each 44 grammes of CO_2 and nothing else:—



COAL GAS.

99. Distillation of Coal.—When coal is heated to a high degree of temperature in a closed vessel, provided with a delivery tube, a great number of products distil over, and there remains in the closed vessel a residue of carbon, which is more or less pure and receives the name of *coke*.

The process of thus heating a substance in the absence of a free access of air, is known as *destructive distillation*, which we may therefore define as *the breaking up of complex substances into simpler ones under the influence of heat, out of contact with air.*

The volatile products may, according to Roscoe, be divided into three classes:

(1) **Coal gas**, a mixture of many gaseous compounds (see below);

(2) **Coal tar**, a thick, oily, strong-smelling liquid;

(3) **Gas water**, or ammonia-water, an aqueous distillate containing ammonia, ammonium carbonate and sulphide, and other substances in solution.

(a) Ammonia-water is at present the chief source of all ammonium compounds.

(b) Tar also contains many substances, from some of which the well-known aniline colours are extracted.

(c) Coal gas consists of a mixture of various compounds of carbon and hydrogen, called *hydrocarbons*, some of which are useful for lighting, others for heating purposes, whilst some are noxious and must be got rid of. All these compounds are on the average, lighter than air; hence coal gas is used to fill balloons.

100. Chief constituents of coal gas:—

- | | |
|---|---|
| (a) olefiant gas, benzene vapour and various other hydrocarbons. | } useful gases; they burn with a luminous flame. |
| (b) hydrogen H_2 , carbonic oxide CO , marsh gas CH_4 . | |
| (c) carbon dioxide CO_2 , sulphuretted hydrogen H_2S , carbon disulphide CS_2 . | } useless or noxious gases; they are mostly removed in the process of purification of coal gas. |
| | |

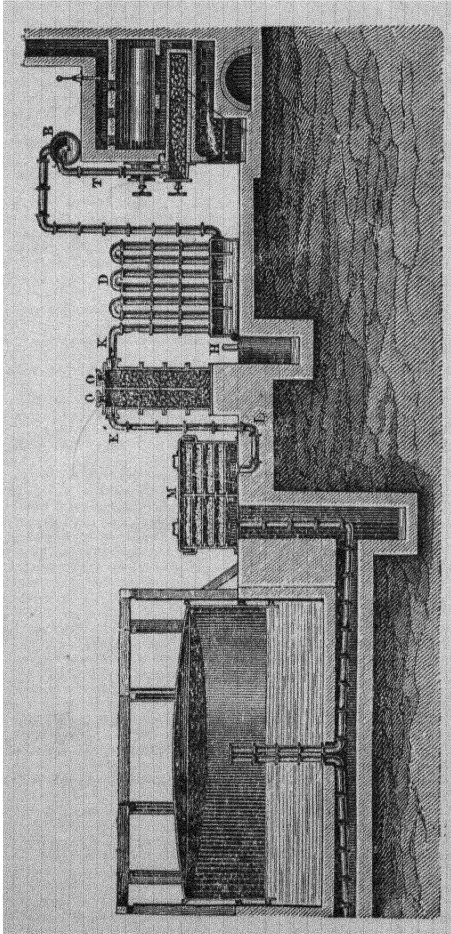


Fig. 58.—Manufacture of coal gas.

The chief parts in the apparatus for the manufacture of coal gas are (Fig. 58) :—

(1) *The retorts, in which coal is distilled. (2) The condensers D, in which the liquid products of the distillation are condensed and separated from the gaseous products. (3) The scrubber O, containing fragments of coke, through which water is allowed to trickle to absorb the remaining ammoniacal vapours. (4) The purifiers M, in which the gaseous impurities are removed. (5) The gas-holder, in which the gas is stored up and from which it is distributed to the places where it is burnt.*

NOTE.—Groups (a) and (b) properly constitute coal gas: group (c) comprises the impurities which accompany the more or less purified gas.

101. **Remarks.**—(1) To prepare coal gas in small quantities, powdered coal is heated in an iron or porcelain tube or vessel (a clay tobacco-pipe may serve for the purpose). The gas escaping from the tube (closed at one end) or from the vessel, is made to pass through a series of bottles, partly filled with cold water, where it is washed. The more or less impure gas is finally collected over water in test-tubes.

(2) If coal be burnt in an open grate, with free access of air, all the carbon is burnt away, and there remains nothing but the incombustible ashes of the coal.

(3) To prove that coal gas contains carbon and hydrogen, hold a cool object over the burning gas: It will soon be covered with a black layer of soot or solid carbon. Moreover, if the products of combustion are made to pass through lime-water, it will turn milky. Therefore CO_2 is one of the products of combustion, and carbon must consequently be present in coal gas. (See test for CO_2 , No. 108).

Again, hold over the flame of burning coal gas a glass vessel kept cool. Small drops of water will soon make their appearance on the inner sides; knowing that water is the product of burning hydrogen, we conclude that coal gas must contain hydrogen.

MARSH GAS, CH_4 .

102. **Occurrence and properties.**—Marsh gas is formed by the decomposition of dead vegetable substances out of contact with air. It occurs naturally in stagnant water or in *marshes*; hence its name. Thus the bubbles of gas that rise to the surface when the mud of a pool is gently disturbed, consist largely of marsh gas, which may be collected by allowing the bubbles to rise in an inverted glass jar

filled with water (Fig. 59).

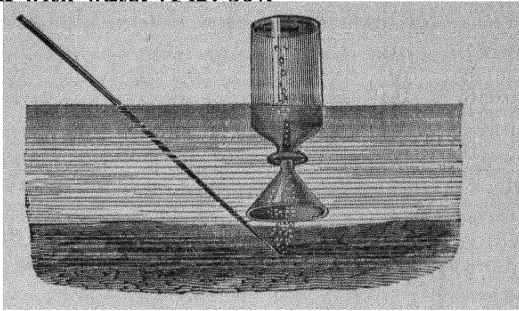


Fig. 59.—Collection of marsh gas.

It is also evolved in coal mines, where mixed with

air it forms a highly explosive mixture, known to miners as *fire-damp*, the cause of many fatal accidents. These accidents or explosions are avoided by using Davy's safety lamp, an improved model of which is represented in Fig. 60. It is an ordinary oil lamp, inclosed in a covering of fine wire gauze, the flame being partly surrounded by a cylinder of stout glass so as to give more light. The principle on which such a lamp is based, is easily understood from the following simple

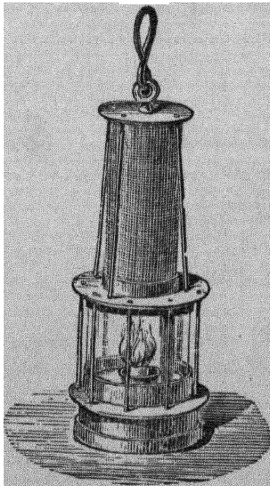


Fig. 60.—Davy-Lamp.

experiment. If a piece of wire gauze be lowered on

to the flame of a spirit lamp, the flame appears to be pressed down by the gauze and does not pass through it until the latter has become heated to whiteness.

Again, if we hold a piece of gauze close over a jet of gas and the gas be lit, it will burn above the gauze (Fig. 61) without the inflammable gas taking fire underneath. In both cases, the flame cannot pass through the gauze, because the heat from the flame being rapidly conducted away by the metallic wires and diffused over the whole

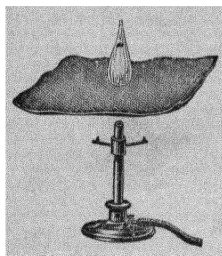


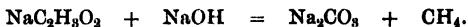
Fig. 61.

extent of the gauze, the temperature of the gas on the other side cannot rise to the point of ignition.

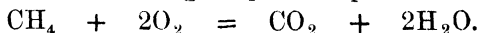
These principles have been applied in the construction of the safety lamp: the flame is surrounded by a cylinder of metallic wires which it cannot cross for the reason just mentioned; although explosive gases from outside still find access to the flame, the latter is prevented from traversing the gauze and all risk of explosion is thereby obviated.

The mixture of gases after the explosion has taken place, consisting chiefly of carbon dioxide and nitrogen, both altogether unfit for respiration, is on this account called *choke-damp* or *after-damp*.

Marsh gas may be prepared by heating in a copper flask a mixture of sodium acetate with soda. To prevent the mixture from fusing, some lime is added; the following reaction takes place:—



Marsh gas is a colourless, odourless and tasteless gas, slightly soluble in water; it does not support combustion, but is itself combustible, and burns with a bluish flame, forming CO_2 and vapour of water :—



CARBON MONOXIDE, CO.

103. Production and properties.—Carbon forms with oxygen two gaseous compounds or oxides, viz., carbon monoxide CO , and carbon dioxide CO_2 .

Carbon monoxide is formed while a combustible substance containing carbon—wood, coal, coal gas, etc.—is burning with an insufficient supply of oxygen.

It is readily prepared by heating oxalic acid with sulphuric acid, whereby water is removed and a mixture of CO_2 and CO is obtained :—



When the mixture is passed through potash, CO_2 alone is absorbed.

It may also be obtained from carbon dioxide. Thus, when carbon dioxide is made to pass through a red-hot iron tube packed with pieces of charcoal, the gas that is evolved does not turn lime-water milky, but burns with a blue flame. It is carbon monoxide :—



If oxygen be substituted for carbon dioxide and the carbon be strongly heated, the same gas is formed :—



It is a colourless, odourless and tasteless gas, insoluble in water; it burns with a pale-blue flame, with formation of CO_2 .

When breathed, it acts as a violent poison.

CARBON DIOXIDE, CO_2 .

104. Occurrence.—Carbon dioxide forms a small but constant and essential constituent of the atmosphere, and is the product of a great number of reactions which take place on the earth's surface, such as the combustion of carbon and organic matters, respiration, the phenomena of putrefaction and fermentation. It exists dissolved in water, and is also evolved from volcanoes.

105. Preparation.—Carbon dioxide is most conveniently prepared by pouring dilute hydrochloric or

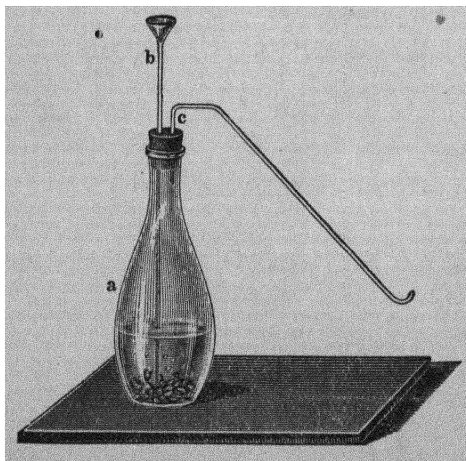


Fig. 62.—Preparation of CO_2 .

sulphuric acid upon calcium carbonate either in the form of chalk or of marble.

Lumps of chalk or fragments of marble are introduced into a glass flask, provided with a safety and a delivery tube (Fig. 62).

The glass vessel is partly filled with water, and hydrochloric acid is gradually added through the funnel tube. Carbon dioxide gas is at once liberated with effervescence :—



The gas, although to some extent soluble in water, can nevertheless be collected over the liquid, as the reaction goes on very rapidly. It is, however, more conveniently collected by mere downward displacement, like chlorine.

106. Properties.—Carbon dioxide, also called carbonic anhydride (*i.e.*, carbonic acid producer), is commonly called carbonic acid. But this name is not correct; for an acid contains hydrogen (see definition of acids, p. 42); the formula of true carbonic acid should be H_2CO_3 , from $\text{CO}_2 + \text{H}_2\text{O}$.

With metals it forms salts called carbonates; but the combination is so weak, that it is set free with effervescence by any other acid.

Under ordinary circumstances, carbon dioxide is a colourless gas; but with diminution of temperature and increase of pressure, it can be liquefied and even solidified to a snow-white substance.

) Carbon dioxide is non-inflammable and is commonly said to be a non-supporter of combustion. A burning

* It is incorrect to say that carbon dioxide is a non-supporter of combustion; for, if a burning magnesium wire, or burning sodium

taper or candle is put out when introduced into the gas or when the gas is poured upon it.

3) CO_2 is a heavy gas; it can be made to flow down from one cylindrical vessel A into another B, like water (Fig. 63).

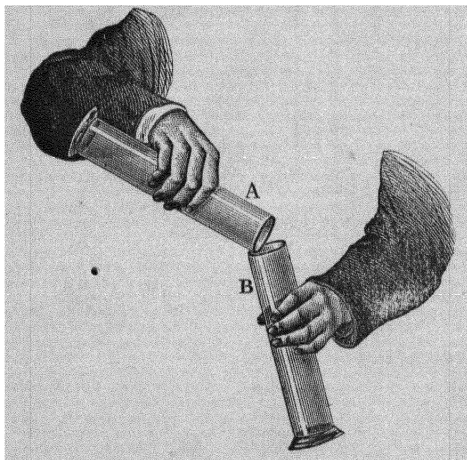


Fig. 63—Carbon dioxide gas poured downwards.

This is rendered more evident by lime-water turning milky when poured into the cylinder B; or by a lighted candle being put out when introduced into the vessel (Fig. 64).

Carbon dioxide is heavier than air and consequently tends to fall to the ground. This can be shown by

or potassium, be introduced into a vessel filled with CO_2 , they continue to burn. The reason is that the above-mentioned metals have a greater affinity for oxygen than carbon has. (Cf. p. 19.)

introducing into a tall, broad glass-cylinder a set of

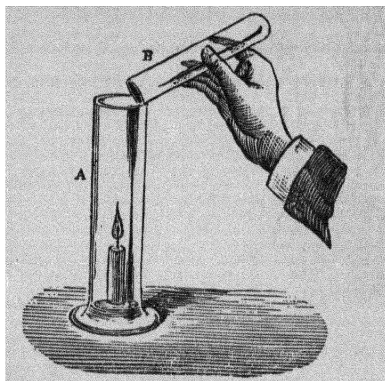


Fig. 64.—Carbon dioxide poured upon a burning candle.
lighted tapers, placed at different heights (Fig. 65).

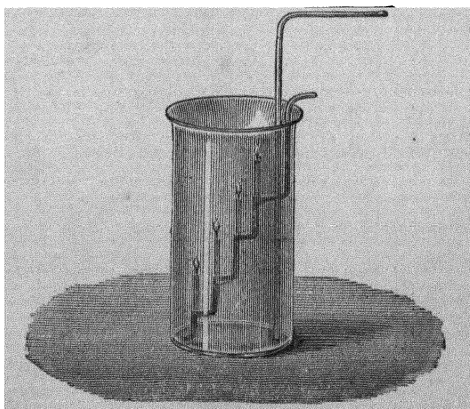


Fig. 65.—Carbon dioxide tends to fall to the ground.

If now a current of carbon dioxide gas is made to

enter the cylinder by a tube reaching nearly to the bottom of the vessel, the lowest taper goes out first, then the second, and so on.

That carbon dioxide is heavier than air, can also be shown by siphoning the gas from a higher to a lower level as in the arrangement of Fig. 66.

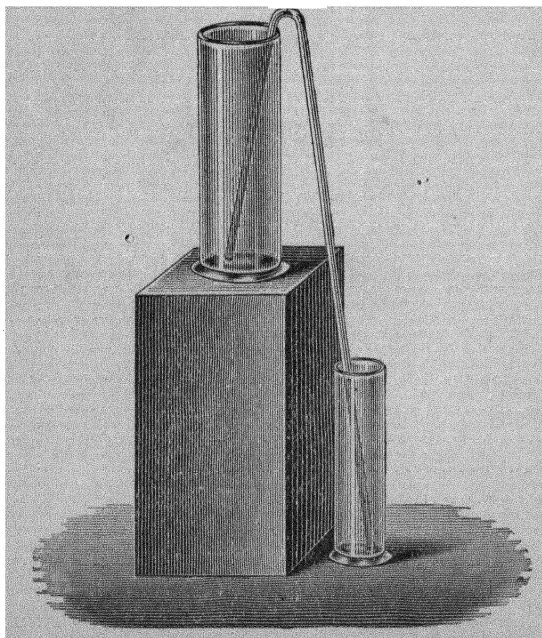


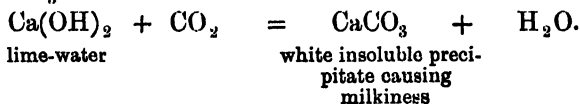
Fig. 66.—Carbon dioxide siphoned from a higher to a lower level.

107. Physiological action of carbon dioxide, or its action upon the human body.—Carbon dioxide is unfit for respiration just as it is unfit for combustion.

Thus, in a room crowded with people, and badly ventilated, respiration becomes more and more painful, and even death may ensue. The reason of it is twofold: 1° there is want of oxygen, 2° there is excess of carbon dioxide.

That want of oxygen prevents respiration, which has been said to be a case of slow combustion, is self-evident. But even with a sufficient supply of oxygen, if there be too much of carbon dioxide, life cannot subsist; not because carbon dioxide is a direct poison, as it is often believed to be, but because when it reaches a certain limit in the atmosphere, the carbon dioxide of the blood can no longer be discharged through the lungs,—(a little in the same way in which water ceases to flow when it has found its level,—the blood is no longer purified, death follows by *asphyxia*, as it is called.

108. Test.—The distinctive test for carbon dioxide is the fact that it turns lime-water milky. The milkiness is due to the formation of insoluble CaCO_3 :—



If, however, we continue to pass the gas through the liquid, the precipitate will partly be re-dissolved, as CaCO_3 is soluble in an excess of carbonic acid, H_2CO_3 , or CO_2 dissolved in water. (See hard and soft water).

CHAPTER VII.

PHOSPHORUS. *pent*

Symb. P. At. wt. = 31.

109. Occurrence.—Phosphorus is *never* found *free* in nature; *combined* it is one of the constituents of many rocks, in the form of *phosphates*. It also exists in the bones of animals as calcium phosphate $\text{Ca}_3(\text{PO}_4)_2$.

110. Preparation.—Phosphorus is now chiefly obtained from *bone-ash* (calcium phosphate), the white solid mass left behind when bones are calcined. Bone-ash, treated with H_2SO_4 , yields a substance which, when mixed with charcoal and strongly heated in a retort, gives off vapours of phosphorus; these are condensed in cold water, and constitute *ordinary* or *yellow* phosphorus.

Red phosphorus is obtained by heating for some time ordinary phosphorus in a closed vessel filled with an inert gas, such as nitrogen: if we heat it while exposed to the air, the phosphorus will merely take fire and burn away.

111. Properties.—Phosphorus exists in two allotropic modifications, known as *yellow* and *red phosphorus*. Both are chemically the same substance, but they differ widely in their general properties.

Yellow Phosphorus.—Ordinary or yellow phosphorus

- (a) is a soft, pale-yellow, waxy-looking solid,
- (b) is luminous in the dark, or phosphorescent,
- (c) is soluble in carbon disulphide,
- (d) melts easily, *i.e.*, at 44°C ., and then takes

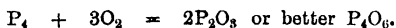
fire spontaneously—hence it must be kept under water for greater security,

- (e) is crystalline,
- (f) is violently poisonous.

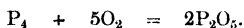
Red phosphorus, on the other hand,

- (a) has a dark-red appearance,
- (b) is not luminous in the dark,
- (c) is not soluble in carbon disulphide,
- (d) does not easily melt, nor does it take fire spontaneously—hence need not be kept under water ;
- (e) is amorphous,
- (f) is not poisonous.

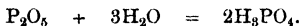
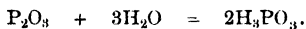
NOTE.—The phenomenon of **phosphorescence** is generally attributed to the slow oxidation which phosphorus undergoes in presence of air, *phosphorous* oxide being formed according to the equation :—



The reaction is distinct from that which takes place when phosphorus burns in oxygen or in air. In this latter case *phosphoric* oxide is formed as is indicated by the equation :—



Both these oxides react on water, forming *phosphorous* and *phosphoric* acid respectively :—



Thus it happens that the water in which phosphorus has been kept for some time, turns blue litmus paper red. This is due to *phosphorous* acid which has been formed by the slow oxidation of phosphorus in presence of the air contained in the water and the action of water on the oxide thus formed.

112. Uses.—Phosphorus is chiefly used in the manufacture of *ordinary* and *safety* matches.

Ordinary matches are made by first dipping the prepared wooden slips in melted sulphur or paraffin, by which they are rendered easily inflammable. These are then tipped with a mixture of phosphorus, glue, lead peroxide, and fine sand, and allowed to dry.

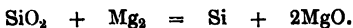
In *safety matches* the ordinary phosphorus is replaced by red phosphorus, which moreover is not applied to the match itself, but to the side of the match box, so that the match will take fire only when rubbed against this specially prepared surface, and nowhere else.

SILICON.

Symb. Si. At. wt. = 28.

113. Occurrence.—Silicon does not occur *free* in nature, but *combined* it is, next to oxygen, the most abundant of all the elements. In the form of an oxide, SiO_2 , or Silica, it exists as *flint* and *quartz* or *rock crystal*. Sand is also a more or less pure kind of silica. It is one of the chief constituents of most rocks, except those of coal and limestone.

As an element, silicon is of no great importance. It may be obtained by strongly heating a mixture of silica and magnesium powder. The product is afterwards treated with an acid to remove magnesia and excess of magnesium :—



The process is analogous to that by which carbon may be obtained from carbon dioxide gas (see no. 16).

114. Compounds of silicon.—Glass.—Silicon forms with oxygen and the metals various *silicates*; thus in combination with aluminium and oxygen it forms aluminium silicate or *clay*—the chief material of all pottery, bricks, porcelain and china-vessels.

Again silicates of sodium, potassium, calcium, and lead, are the basis of all the varieties of glass.

Window or *plate glass* is made by fusing together white sand (silica), slaked lime and sodium carbonate; this gives sodium calcium silicate.

Flint glass consists chiefly of potassium and lead silicate, and is obtained by fusing together sand, lead oxide and potassium carbonate.

115. Coloured glass.—Coloured glasses are made by adding to the white glasses small quantities of metallic oxides; thus cobalt oxide gives an intense blue glass; chromium gives green; copper, green or blue; iron, the characteristic bottle-green glass.

To show this experimentally, heat some borax or sodium borate in the loop of a platinum wire with the help of a blowpipe. After a short time, a transparent, glassy bead will be formed. Fuse into the borax bead so obtained a minute quantity of any of the above-mentioned metallic oxides. The oxide will dissolve in the bead, and impart to it its characteristic colour.

PART III.

METALS.

NOTE.—The order which is followed in the study of the metals, is that of their atomicity, given in the table page 40.

CHAPTER 1.

POTASSIUM.

Symb. K. (Kalium). At. wt. = 39. *Monatomic*

116. Occurrence.—Potassium is never found *free* in nature ; but *combined*, it is met with as a chloride KCl, as a nitrate KNO_3 or saltpetre, and as a silicate, chiefly in the form of mica and felspar, in granitic rocks.

Potassium salts are also contained in the substance of most plants, by which they are absorbed through the roots while in solution in the soil. When these plants are burnt down, the potassium remains behind in the form of potassium carbonate K_2CO_3 , one of the chief constituents of ashes.

NOTE.—The sap which ascends into the plants by the roots, contains mineral salts in solution. When the combustible parts of the plants are burnt away, these mineral, incombustible salts remain behind, and constitute what are called the *ashes* of the wood. The chief mineral substances found in the ashes, besides potassium carbonate, are salts of calcium and magnesium.

117. Preparation.—Potassium was for the first time obtained free by H. Davy (1807), by means of the electrolysis of fused caustic potash.

Potassium is usually prepared by heating in an iron retort a mixture of K_2CO_3 and carbon. The vapour of potassium which distils over is condensed in a

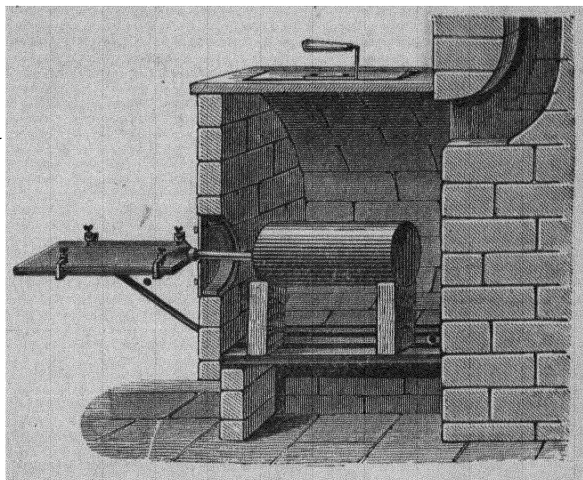
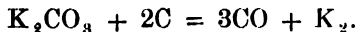


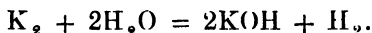
Fig. 67.—Preparation of potassium.

shallow receiver (Fig. 67) and made to flow into petroleum (rock-oil), or naphtha oil, which prevent it from oxidizing, as they contain no oxygen :



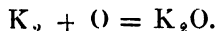
118. Properties.—Potassium is a light metal, capable of floating upon water, and soft like wax. Its freshly-cut surface is silver-white, but very soon tarnishes on exposure to the air owing to the formation of a dull-white layer of oxide, potassium having a remarkably great affinity for oxygen.

When a piece of potassium is thrown upon water, it rushes about on the surface of the liquid, and takes fire burning with a violet flame, and finally disappears with a slight explosion. What happens is this: potassium decomposes the water, setting free part of the hydrogen and combining with the oxygen and the other part of hydrogen to form caustic potash:—

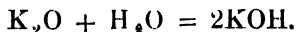


The heat evolved by the chemical reaction is so great as to set fire to the liberated hydrogen, while the melted potash, being raised to a high temperature, suddenly produces a small quantity of steam which causes the slight explosion. The violet colour of the burning hydrogen is due to the presence of vapour of potassium.

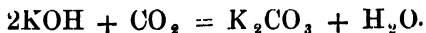
When metallic potassium is left exposed to the atmosphere, it first forms an oxide:—



This oxide next dissolves in the atmospheric vapour forming potash:—



In its turn the potash thus formed absorbs CO_2 from the air, forming K_2CO_3 :—



Potassium has also a great affinity for chlorine and forms with it a chloride, KCl .

119. Uses.—Potassium may be used for the same purposes as metallic sodium, but the latter is usually

preferred as it is cheaper and more manageable; moreover a smaller quantity is sufficient to produce the same effect, as the atomic weight of sodium is smaller than that of potassium.

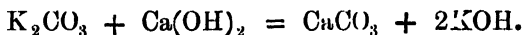
CHIEF POTASSIUM COMPOUNDS.

120. K_2CO_3 , potassium carbonate, commercially known as *potashes*, was originally obtained solely from wood-ashes boiled in pots—hence its name. At present it is also obtained in various other ways, particularly from KCl, which is found in vast deposits at Stassfurt in Germany.

It is used in the manufacture of glass, soft soap, and in the preparation of many other potassium compounds.

121. KOH, potassium hydroxide — *potash* or *caustic potash*.

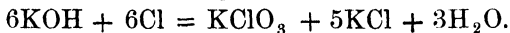
Preparation.—By adding calcium hydroxide to a solution of K_2CO_3 :—



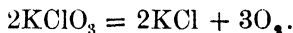
After the insoluble $CaCO_3$ has subsided, the clear liquid is poured off and evaporated in iron or silver pans, any other metal being attacked by boiling potash.

Properties and uses.—KOH has a great attraction for water and carbon dioxide. Hence its use to dry gases and to absorb CO_2 . It is also used in the manufacture of soap.

122. KClO_3 , potassium chlorate, is prepared by passing chlorine through KOH :—



When heated, KClO_3 is decomposed into KCl and oxygen:—



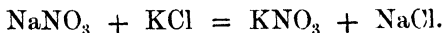
Hence it is used in the preparation of oxygen.

123. KNO_3 , potassium nitrate—*nitre or saltpetre*— is found in the form of efflorescent crystals in the soil of warm countries as India, Egypt, Persia...

To collect it, the earth is taken off to a depth of 2 or 3 inches, and is then treated with water. The solution, on evaporation, yields crystals of potassium nitrate; these are, however, impure and must be refined.

This salt is formed wherever nitrogenised organic matter undergoes decomposition in presence of potash.

The greater part of nitre is now obtained by decomposing Chili saltpetre or sodium nitrate with potassium chloride:—

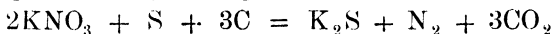


The potassium nitrate thus formed, being less soluble in cold water than sodium chloride, crystallises out first and is removed.

Uses—Nitre is chiefly used in the preparation of nitric acid, and in the manufacture of gunpowder.

124. Gunpowder is a mixture of nitre, charcoal and sulphur, approximately in the proportion of 75, 13 and 12 per cent respectively.

Gunpowder can burn in a space devoid of air, as the burning substance, viz., carbon, gets the necessary oxygen from the potassium nitrate, to form carbon dioxide gas. The chemical change which takes place when gunpowder explodes, may, in the simplest way, be represented by the equation:—



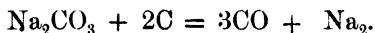
It has been calculated that the volume of the gases formed at the moment of the explosion, is about 1500 times the volume occupied by the solid powder; and the force of explosion of gunpowder is due to this fact.

SODIUM.

Symb. Na (Natrium). At. wt. = 23.

125 Occurrence.—Like potassium, sodium is never found *free*, on account of its great affinity for oxygen or chlorine. *Combined*, it is much more abundant than potassium, and exists chiefly as a chloride NaCl , and as a nitrate NaNO_3 or Chili saltpetre. It is also present in plants growing in the sea and on the sea-shore. When these plants are burnt down, their ashes yield chiefly sodium carbonate Na_2CO_3 .

126. Preparation.—The mode of preparation of sodium is entirely similar to that of potassium:—



127. Properties.—The properties of sodium resemble those of potassium, with this exception that sodium acts in general somewhat less energetically

than potassium. When thrown upon cold water, it floats about and decomposes the water like potassium; but the heat evolved in this case is not enough to ignite the liberated hydrogen. However, if it be thrown upon hot water, or water thickened with gum or starch, or upon wet blotting-paper, so as to prevent the sodium from floating about, the metallic globule becomes sufficiently heated to ignite the hydrogen evolved, which then burns with a yellow flame, the colour being due to vapour of sodium.

128. Uses.—Sodium is much used in the preparation of metallic magnesium and aluminium.

129. Test.—When a sodium salt is held on the loop of a clean platinum wire in a non-luminous alcohol flame, it colours the flame bright *yellow*, whilst potassium salts colour the flame *violet*.

CHIEF SODIUM COMPOUNDS.

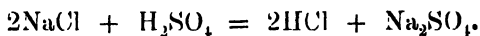
130. NaCl, sodium chloride—or *common salt, rock salt, sea salt, table salt*—is found abundantly in nature. In the *solid* state, it is found in many countries in vast deposits, and is then known as *rock salt*; in the *dissolved* state, it exists in the sea, and forms the greater portion of the solid constituents of sea-water.

Extraction.—Large quantities of rock salt are cut out in the solid state from mines. Sometimes it is obtained from rock salt deposits by allowing water to flow through them; the water, when saturated with salt, is pumped out and evaporated.

Another method of obtaining common salt on a large scale consists in the evaporation of sea-water and also of waters of salt springs. In warm countries sea-water is led into large, shallow basins, dug along the sea-shore. Exposed to the heat of the sun, the brine concentrates and deposits crystals of salt; these are raked off and, if necessary, purified by re-crystallisation.

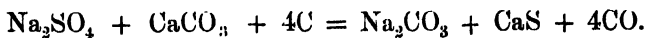
Uses.—The applications of common salt are very numerous, the chief among them being :—to prepare hydrochloric acid; chlorine, and many sodium salts; to glaze earthenware vessels; to preserve food from putrefaction. It is finally a common ingredient in food, serving as a condiment to give it flavour, and is thus one of the necessaries of animal life.

131. Na_2SO_4 , sodium sulphate—(*Glauber's salt*)—is obtained by heating NaCl with H_2SO_4 (see preparation of HCl):—



This salt is prepared chiefly in order to obtain from it the more important Na_2CO_3 .

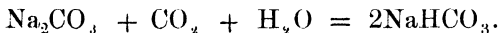
132. Na_2CO_3 , sodium carbonate—*crystals of soda*—*washing soda*—(also improperly called *soda*)—used formerly to be obtained from the ashes of sea-plants, but is now got by roasting together Na_2SO_4 , charcoal, and chalk :—



Uses.—It enters in the manufacture of soap and

glass, and in the preparation of other sodium compounds. It is moreover used in softening water.

133. NaHCO_3 , sodium hydrogen carbonate - sodium acid carbonate, bicarbonate of soda—is obtained by passing CO_2 through a solution of Na_2CO_3 :—

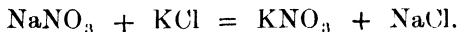


Uses.—In medicine, and in the manufacture of effervescent drinks, and for baking powders.

134. NaOH , sodium hydroxide or soda.—What has been said concerning the preparation, properties and uses of KOH applies equally well to NaOH .

135. NaNO_3 , sodium nitrate—*Chili saltpetre*—is chiefly found in South America. Its properties are similar to those of KNO_3 ; but as it is slightly deliquescent, it cannot be used to make gunpowder.

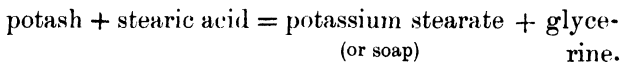
Uses.—It is chiefly used in the preparation of nitric acid. Potassium nitrate is now also largely obtained from it:—



136. Manufacture of soap.—To complete the study of the alkali metals, potassium and sodium, something is to be said about soap.

Soap is an organic salt—an *alkaline stearate*, i. e., a salt in which the basic radical is potassium or sodium and the acid radical is stearic acid. Stearic acid is present in every fatty or oily substance; in order, therefore, to prepare K or Na stearate or soap, it is sufficient to boil for some time caustic potash or soda

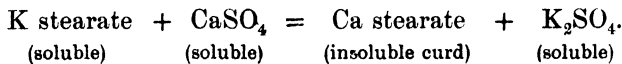
with oil or fat in an iron vessel—any other metal except silver would be attacked by the boiling potash, even glass could not resist its action. By double decomposition, soap is formed. Just as by treating KOH with HCl we get KCl, a salt, and water or hydrogen hydroxide, in the same manner we here obtain an alkaline organic salt and an organic hydroxide called glycerine :—



If now we throw common salt into the mixture containing the dissolved soap, the latter will come out of the solution and float on the surface where it may be collected.

If we use potash, we get soft soap ; if soda, we get hard soap.

Note. 1.—K or Na stearate, or soap, being soluble in water, and Ca or Mg stearate being insoluble, we can easily understand why soap is used as a test for hard and soft water. If the water be soft, that is, free from salts of magnesium and calcium, the soap will simply dissolve producing froth ; but if the water be hard, that is, if it contain salts of magnesium and calcium, then by double decomposition, Ca or Mg stearate is formed, and as this is insoluble, it appears as a precipitate or a curd :—



We can thus also understand why hard water will

froth only after all the calcium or magnesium salt has been precipitated.

2.—The cleansing action of soap depends upon the fact that soap dissolves the oily film which covers the skin; this being washed away, all the remaining dirty matter is easily taken off.

AMMONIUM.

137. Ammonium theory.—In connection with sodium and potassium, it will not be out of place to say a word about *ammonium*, as the name occurs often in this little work.

Ammonia combines with water to form a basic hydroxide, in many respects similar to NaOH and KOH. Again ammonia combines with acids to form salts exhibiting properties similar to the corresponding sodium and potassium salts.

Now the best way to account for this, is to assume the existence of the group of atoms or radical NH_4 , which would behave exactly like the atoms of metallic sodium Na, or potassium K; and to remind us of its behaviour being similar to that of metals, the name of ammonium has been given to it, the ending *um* being common to the names of the metals. The radical NH_4 , however, has never yet been isolated or obtained free, like sodium or potassium.

The analogy that exists between NH_4 , Na and K, may be seen from the following compounds:

NaOH—sodium hydroxide,
 KOH—potassium hydroxide,
 NH₄OH—ammonium hydroxide.

Similarly :

NaCl,	KCl,	NH ₄ Cl,	sodium-,	potassium-,	ammonium chloride,
NaNO ₃ ,	KNO ₃ ,	NH ₄ NO ₃	,,	,,	,, nitrate,
Na ₂ SO ₄ ,	K ₂ SO ₄ ,	(NH ₄) ₂ SO ₄	,,	,,	,, sulphate,
Na ₂ S,	K ₂ S,	(NH ₄) ₂ S	,,	,,	,, sulphide,
Na ₂ CO ₃ ,	K ₂ CO ₃ ,	(NH ₄) ₂ CO ₃	,,	,,	,, carbonate.

138. NH₄Cl, ammonium chloride—commercially known as *sal ammoniac*—is one of the chief compounds of ammonium with which we have to deal. It is formed whenever hydrochloric acid and ammonia come into contact.

It is manufactured by neutralizing with HCl the ammoniacal gas-liquor containing ammonia, and evaporating to dryness.

Uses.—It is used in the manufacture of ammonia; it is also used by tinkers to clean the surfaces of metals which are to be tinned or soldered: *To un*
stallie surfaces by a fusible metall

CHAPTER II. *Mona*

SILVER.

Symb. Ag. (Argentum). At. wt. = 108.

139. Occurrence.—Silver, having no affinity for oxygen, is found *free* in nature; hence it has been

known and used from the remotest times. *Combined*, it occurs more generally as a sulphide, Ag_2S , either alone or associated with galena or PbS . Large silver mines exist in North and South America.

140. Extraction.—(1) The method of obtaining metallic silver from the sulphide is very simple in theory but complicated in practice :—

(a) Silver sulphide Ag_2S is heated with common salt. As silver has a greater affinity for chlorine than for sulphur, by double decomposition AgCl and Na_2S are formed.

(b) Next the silver chloride is shaken up with mercury and scrap-iron : the iron combines with the chlorine, and the silver is dissolved by the mercury, forming an amalgam.

(c) This amalgam is then distilled ; the mercury being very volatile, distils over while metallic silver remains behind.

(2) Silver is often obtained from ores of lead :—

(a) The whole mass is melted and allowed to solidify ; the greater portion of lead separates pure in the form of crystals which are ladled out. This operation is several times repeated.

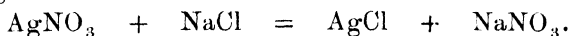
(b) Finally the remaining mass, now rich in silver, is heated in a current of air ; the lead alone oxidizes, melts and flows away, whilst metallic silver is left.

141. Properties. Silver is the whitest and most brilliant of all the ordinary metals. It is the best conductor of heat and electricity known. It is after gold the most malleable and ductile metal. Its density is $10\frac{1}{2}$.

It has a great affinity for sulphur and no affinity for oxygen ; hence it remains unaltered in the air, but in presence of vapours of sulphur it becomes covered with a black film of silver sulphide.

142. Uses—As silver does not become tarnished in the air, and possesses besides a great lustre, it is highly valued and used in jewelry and for coinage. But pure silver being a little too soft for the purpose, it is usually alloyed with some copper (from 7 to 10 per cent).

143. To show that a silver coin contains copper.—Cut a piece off a silver coin, and drop it into nitric acid. The metal will dissolve entirely, giving off brown nitrous fumes in abundance. The solution is but a mixture of silver and copper nitrate. To separate the two metals, add slowly a solution of NaCl, so long as a white precipitate of insoluble AgCl is seen to be formed:—



Remove the precipitate by filtration. The clear filtrate exhibits a greenish-blue tint which already indicates the presence of a copper salt. It can now directly be proved to contain copper, by dipping into it a bright, clean iron wire: brown metallic copper will be deposited upon it.

N. B.—As there may still be some acid remaining in the solution, it should first be neutralized by slowly pouring upon it a solution of potash or ammonia, until a light-blue precipitate begins to be formed, as otherwise the acid would simply dissolve the iron.

If we continue to pour ammonia in excess, a deep-blue solution is obtained, and this is another test for the presence of copper. (See tests for copper).

N. B.—Metallic silver can be recovered from the white precipitate of AgCl by mixing it with dry sodium carbonate and heating the mixture strongly on charcoal, with the blow-pipe.

144. AgNO_3 , silver nitrate, is obtained by dissolving silver in HNO_3 and evaporating the solution to dryness. A white salt is thus obtained very soluble in water. It is used in medicine under the name of *lunar caustic* to burn wounds and to destroy some skin diseases.

Silver salts are acted upon by sunlight, chiefly in presence of organic matter; hence their use in photography.

They are moreover employed in the preparation of indelible inks, as they produce stains which cannot be taken away by ordinary washing; these stains however can be removed by potassium cyanide.

145. AgCl , silver chloride, is a white powder insoluble in water. It is formed when a solution of AgNO_3 is treated with HCl or any metallic chloride in solution. Hence the use of AgNO_3 to detect the presence of sodium chloride or common salt in water.

CHAPTER III.

CALCIUM.

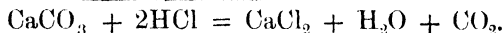
Symb. Ca. At. wt. = 40. *Diatomic*

146. Occurrence.—Calcium is *never* found *free* in nature, on account of its great affinity for oxygen (compare with K and Na). In *combination* it is one of the most abundant and widely distributed elements

in the crust of the earth. It exists chiefly in the form of CaCO_3 , CaSO_4 , and as a silicate.

Metallic calcium is of no importance and is moreover difficult to obtain. It is prepared by the electrolysis of fused CaCl_2 . It acts chemically upon water like sodium and potassium, though less energetically, and has a very great affinity for oxygen and chlorine.

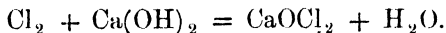
147. CaCl_2 , calcium chloride—is obtained by decomposing chalk or marble with hydrochloric acid :—



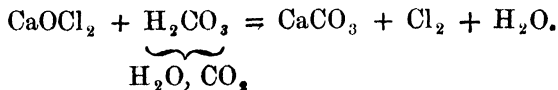
The solution of CaCl_2 thus obtained is filtered, if necessary, and evaporated to dryness.

It has the property of absorbing water to a remarkable degree : hence its use in the laboratory to dry gases.

148. Calcium chloro-hypochlorite—or *bleaching powder*, CaOCl_2 or $\text{Ca}(\text{OCl})\text{Cl}$ —is obtained by the action of chlorine upon slaked lime, as represented by the equation :—

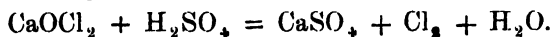
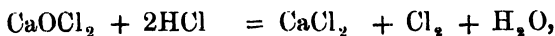


The chlorine it contains is so loosely combined, that even such weak acids as carbonic acid can displace it. Consequently, on merely exposing bleaching powder to the air, the CO_2 present in the atmosphere, combines with it, forming CaCO_3 , while chlorine is set free :—



Now, as has been said previously (Cf. no. 83), the chlorine thus liberated combines, in presence of moisture, with the hydrogen of the water, and sets the oxygen free; the oxygen in the *nascent state* (i.e., at the moment of liberation from the compound) oxidizes the colouring matter and renders it colourless, i.e., bleaches it.

The decomposition of bleaching powder is helped by the addition of dilute HCl or H₂SO₄, thus:—



149 CaCO₃, calcium carbonate—*limestone*—*chalk*—*marble*—*Iceland spar*—*calcspar*—is very widely distributed on the surface of the earth; in many countries it forms large beds of rocks. It is also the chief constituent of most shells. When heated, it is decomposed into quicklime and carbon dioxide:—



It is insoluble in pure water but dissolves in water containing carbonic acid, i.e., carbon dioxide in solution; it thus causes temporary hardness of water.

It is much used in the manufacture of lime and cement.

150. CaO, calcium oxide—*lime*—*quicklime*—is obtained by heating limestone in special furnaces or *lime-kilns* (Fig. 68).

Quicklime is not decomposed under the action of heat; it is one of the most infusible substances known,

and it emits a very brilliant light when raised to a high temperature. Hence its *uses*—

(a) in the manufacture of crucibles, or in lining furnaces where a very high temperature is to be had ;

(b) in the production of *Drummond light* or *limelight*, with the oxyhydrogen blow-pipe.

Quicklime combines with water with evolution of much heat, and formation of calcium hydroxide.

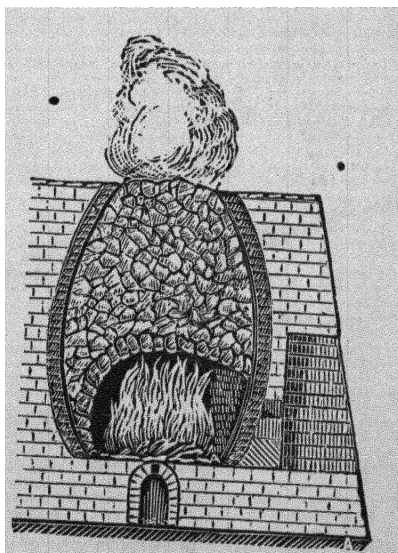
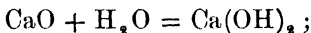


Fig. 68.—Lime-kiln.

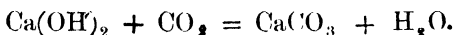
151. $\text{Ca}(\text{OH})_2$, calcium hydroxide.—When we pour just enough of water upon quicklime, so that every

molecule of CaO can combine with a molecule of H₂O, we get this reaction:—



and this is properly *calcium hydroxide* or *slaked lime*. It exists in the form of a white powder. Now this powder of Ca(OH)₂ is soluble in water. A clear solution of it is called *lime-water*. When slaked lime is shaken up with water, so as to form a white, creamy liquid, it is called *milk of lime*, and is used to white-wash walls and houses.

Calcium hydroxide in any of the above forms, has the power of absorbing CO₂, forming CaCO₃, with elimination of water:—



Hence freshly built walls remain for a long time damp.

It is used for very many purposes, chiefly—

(a) in the preparation of mortar—a mixture of slaked lime and sand;

(b) in the manufacture of glass;

(c) in the preparation of bleaching powder, of KOH, NaOH, NH₄OH;

(d) to whitewash buildings and to disinfect.

152. CaSO₄, calcium sulphate—gypsum—plaster of Paris.—Gypsum, properly so called, is CaSO₄. 2H₂O, a crystalline substance. When this is heated, it loses its water of crystallisation and crumbles down to a fine white powder, which can be made into a paste

when treated with water. It hardens very quickly, and increasing slightly in volume, assumes the shape of any object upon which it is spread; hence it is much used to make casts or moulds.

The white powder obtained by heating gypsum, is called *Plaster of Paris*.

Gypsum is slightly soluble in water, and renders it permanently hard. The hardness can be removed to some extent, by adding to the water a solution of Na_2CO_3 , which causes insoluble CaCO_3 to be precipitated (no. 62).

153. $\text{Ca}_3(\text{PO}_4)_2$, calcium phosphate—or *bone-earth*—is found in the soil; it is also one of the chief constituents of bones, and remains behind in the form of *bone-ash*, when bones are calcined in the open air. From this, phosphorus is extracted.

154. Test for calcium.—Calcium salts, when held in a loop of platinum wire in the alcohol flame, colour the flame *brick-red*.

Strontium salts give it a beautiful *crimson-red* hue; hence their use in fire-works.

Barium salts colour it *pale-green*.

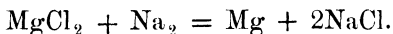
CHAPTER IV.

MAGNESIUM.

Symb. Mg. At. wt. = 24.

155. Occurrence.—Magnesium does not exist *free* in nature, as it has too great an affinity for oxygen. *Combined*, it is widely distributed as a constituent of many rocks. It is found in solution in many spring-waters (particularly at Epsom in England), in the form of MgSO_4 , and in sea-water as MgCl_2 . It is also found as carbonate, and in the mineral dolomite $\text{MgCa}(\text{CO}_3)_2$.

156. Preparation and properties.—It is difficultly obtained in the metallic state. It is prepared by heating MgCl_2 with metallic sodium:—

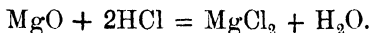


It has a great affinity for oxygen, with which it combines, when heated, with display of a very brilliant, dazzling flame. Hence it is used where very bright light is required, as for signalling or for photography at night or in dark places.

It is a light, soft, silver-white metal, and does not easily tarnish in dry air; in moist air it becomes coated with a film of oxide.

157. MgO, magnesium oxide—or *magnesia*—is the white powder formed when magnesium burns either

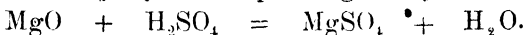
in air or in oxygen. When dissolved in HCl, it forms MgCl_2 :—



Magnesium chloride is a white, crystalline substance, very deliquescent. It is the presence of this substance in ordinary sodium chloride that renders the latter (hygroscopic.) *indicate the hu atmosphere*

158. MgSO_4 , magnesium sulphate—or *Epsom salt*—is obtained by evaporating the spring-waters that contain it dissolved. It has a bitter taste, and is used in medicine as a purgative.

It may be prepared artificially by gently heating MgO with H_2SO_4 , and evaporating to dryness :—



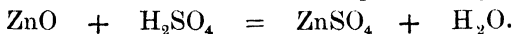
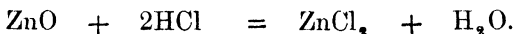
ZINC. *Diatomus*

Symb. Zn. At. wt. = 65.

159. Occurrence—Zinc is rarely found *free*; *combined*, it is found chiefly as a sulphide ZnS , or *blende*; it exists also as ZnO and ZnCO_3 .

160. Preparation and properties.—Metallic zinc is obtained by heating ZnO with charcoal: the metal volatilises and is then condensed. It is a bluish-white metal. In dry air it keeps its metallic lustre; but when exposed to moist air, it tarnishes like magnesium, and the whitish layer of oxide formed on the surface, protects the metal below against further oxidation. When strongly heated in the air, zinc burns

with flame, leaving a white powder of ZnO behind (Compare with magnesium). This, when treated with HCl or H₂SO₄, forms the chloride or the sulphate:—



When pure, zinc is scarcely attacked by HCl or H₂SO₄; but ordinary commercial zinc, which is impure, is energetically attacked by these acids: hydrogen is liberated (see preparation of hydrogen), and liquid ZnCl₂ or ZnSO₄ is left behind; either can be got in the solid form by filtering and evaporating the liquid.

Zinc chloride is, like magnesium chloride, very deliquescent.

161. Uses.—As zinc is not so easily affected by the air, it is used for coating iron vessels, for roofing, for water-pipes, etc.: the iron thoroughly cleaned is plunged into molten zinc, and thus gets covered with a layer of this metal; such iron is said to be *galvanised*.

Zinc is also used to form an alloy with copper, called *brass*, and for electric batteries.

CHAPTER V.

COPPER. *Diatonic*

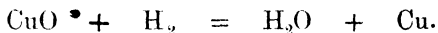
Symb. Cu. (Cuprum). At. wt. = 63.

162. Occurrence.—Copper is found *free* in nature, chiefly near Lake Superior, North America. The

chief ore of copper is copper sulphide, either alone or associated with iron sulphide. There exists also the oxide Cu_2O , and the carbonate CuCO_3 .

163. Preparation.—Metallic copper is easily obtained from the oxide or carbonate by simply heating either of them with charcoal; to obtain it from the sulphide is more difficult.

Pure metallic copper can be obtained by dipping a clean iron or zinc plate into a copper solution; or again by heating copper oxide in a glass bulb and passing over it a current of dry hydrogen:—



164. Properties.—Copper is a metal, of a peculiar deep-red colour, very malleable and ductile, and, next after silver, one of the best conductors of heat and electricity.

It does not oxidize at ordinary temperatures in either dry or moist air, when pure; but it does, when strongly heated, or when in presence of acids, even of carbonic acid. In the latter case it forms a green compound, copper carbonate, called by some authors *verdigris*, which is poisonous.

165. Uses.—It is much used to make boilers, vessels, tubes, wires, etc; also to make alloys, chiefly *brass* (Cu and Zn), and *bronze* (Cu and Sn).

166. CuO , copper oxide—known as *copper scales*—is formed when copper is strongly heated in the air. It is easily reduced to the metallic state, by heating

it with carbon, or by passing a current of hydrogen gas over it while heated (Fig. 69).

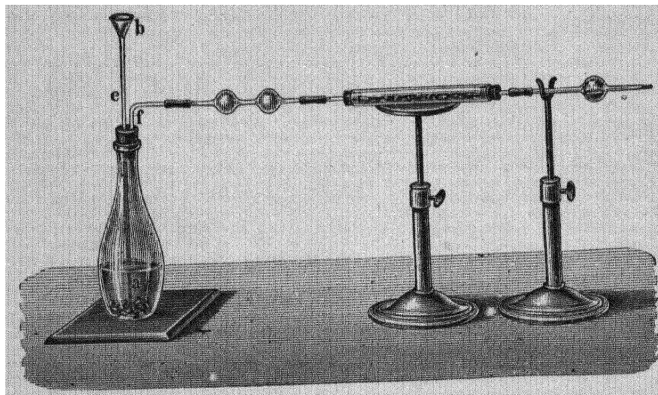


Fig. 69.—Copper oxide reduced by hydrogen.

It is used in the quantitative analysis of water, and also in the manufacture of coloured glass.

167. CuSO_4 , copper sulphate—*blue vitriol* or *blue stone*—can be obtained by heating copper with H_2SO_4 , although in practice other methods are preferred.

When ammonia is added to a solution of copper sulphate, a light-blue precipitate of copper hydroxide is first obtained. If ammonia be added in excess, this precipitate dissolves into a liquid of a beautiful dark-blue colour.

$\text{Cu}(\text{NO}_3)_2$, copper nitrate—is formed as a deep-blue solution when either metallic copper, or the oxide of copper, or copper carbonate, is acted upon by nitric acid.

CuCl₂, copper chloride—may be obtained as a green solution, when copper oxide or carbonate dissolves in hydrochloric acid. It is also formed when chlorine is made to pass over heated copper (Cf. No. 82). When the green solution obtained is diluted with water, it turns blue.

NOTE.—When a piece of clean iron is dipped in a solution of a copper salt, *e.g.* copper sulphate, the liquid gradually loses its colour until at last it assumes a greenish tint. If the piece of iron is now taken out of the liquid, it is found to be coated with a red-brown layer of metallic copper. What happens is this: the iron displaces the copper, atom for atom, from the solution; ferrous sulphate is formed, which gives to the solution its greenish appearance, and an equivalent quantity of copper is set free, *i.e.*, for every 63 parts of copper which are liberated, 56 of iron have passed over to the solution:—



168 Tests for copper salts:

(a) A characteristic deep-blue coloured liquid obtained by adding excess of ammonia to a copper solution.

(b) Metallic copper deposited, when clean iron or zinc is dipped in the solution.

(c) The alcohol flame coloured vivid green or blue by copper salts.

MERCURY.

Symb. Hg. (Hydrargyrum). At. wt. = 200. *Si'a*

169. Occurrence.—Mercury in the *free* state is met with only in small quantities. Its principal ore is the

sulphide, HgS , or *cinnabar*, which is chiefly found in Spain (Almaden) and in Illyria.

§170. **Preparation.**—Metallic mercury is obtained from its ore by heating the latter in a current of air, and condensing the vapour of mercury in a series of chambers (Fig. 70). The reaction may be represented by the following equation:—

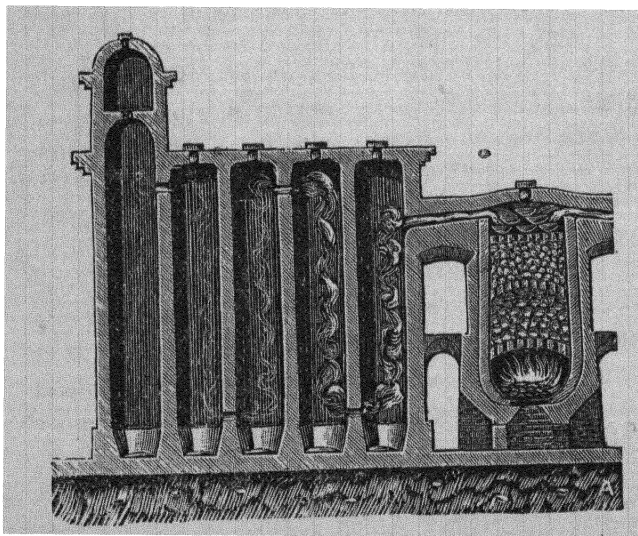
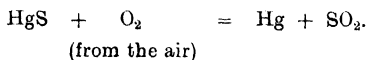


Fig. 70.—Extraction of mercury.

171. **Preparation.**—Mercury, sometimes called *quicksilver* (in allusion to its mobility and bright colour), is a silver-white, heavy ($13\frac{1}{2}$ times heavier than water) liquid metal—the only liquid metal at the ordinary temperature. (1) It is not tarnished in

pure air. } Heated above 300°C . in presence of air or oxygen, it is changed into the red oxide HgO . It boils at 350°C ., solidifies at -40°C .

↳ It is not attacked by HCl and H_2SO_4 at the ordinary temperature; hence it is used to amalgamate the zinc plates of electric batteries. HNO_3 transforms it into the nitrate:—



↳ Copper displaces it from its solutions. Thus when a clean strip of copper is introduced into a solution of mercuric nitrate or chloride, it gets soon covered with a film of metallic mercury and the liquid turns blue. Copper in fact takes the place of mercury in the solution, forming copper nitrate or chloride, and metallic mercury is thrown out of the solution:—



↳ Mercury dissolves metals forming amalgams.

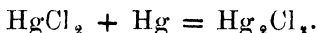
NOTE.—When metals unite with one another (to form compounds or mixtures), they are said to form *alloys*; when one of the metals is mercury, the name of *amalgam* is given to the mixture.

172. **Uses.**—Mercury is used for constructing thermometers and barometers, for glass mirrors, for collecting gases which are soluble in water, for extracting silver and gold from their ores. Mercury compounds are much used in medicine.

173. **HgO, mercuric oxide**—is obtained by heating mercury in the air above 300°C ., but more practically by heating mercury nitrate. When heated at a temperature higher than that of its formation, the

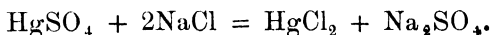
oxide is decomposed into metallic mercury and oxygen, and on this account, is employed in the preparation of oxygen gas.

174. Hg_2Cl_2 , mercurous chloride—or *calomel*—is obtained by heating mercuric chloride, HgCl_2 , with mercury :—



This salt is insoluble in water ; it is used in medicine.

175. HgCl_2 , mercuric chloride—or *corrosive sublimate*—is obtained by heating a mixture of HgSO_4 and NaCl :—



It is soluble in water and is a strong poison.

N.B.—The antidote for mercury poisoning is albumen (white of eggs).

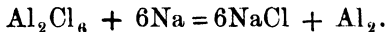
CHAPTER VI.

ALUMINIUM.

Symb. Al. At. wt. = 27. *Tri*

176. Occurrence.—Aluminium is not met with in the *free* state ; but in the form of alumina and aluminium silicate, it is found widely distributed over the earth and is the chief constituent of clay and slate.

¶ **177. Preparation.**—It is obtained by heating aluminium chloride with metallic sodium :—



178. Properties and uses.—Aluminium is a silver-white, shining metal; is very light being only $2\frac{1}{2}$ times heavier than water; and moreover takes a high polish and does not tarnish in the air. Were it not for the difficulty of extraction, aluminium would, on account of its valuable properties, be very much used instead of iron and copper for many purposes, where lightness and strength, together with indifference to atmospheric corrosion, are required.

It is easily attacked by hydrochloric acid and boiling potash.

179. Al_2O_3 , aluminium oxide—or *alumina*—is obtained in the form of an amorphous white powder, by heating crystals of alum, that is, aluminium potassium sulphate, $\text{Al}_2\text{K}_2(\text{SO}_4)_4 \cdot 24\text{H}_2\text{O}$. It is also obtained by strongly heating aluminium hydroxide $\text{Al}_2(\text{OH})_6$, or burning aluminium foil in oxygen.

The native alumina is called *corundum*, and is the hardest substance known next to diamond. Sometimes small quantities of metallic oxides are found mixed with alumina, imparting to it beautiful shades of colour, and forming what are called *precious stones*, as rubies, sapphires. . . .

heavier than water GOLD. *triatonic*

Symb. Au. (Aurum). At. wt. = 197.

180. Occurrence.—Gold occurs *mostly* in the *free* state, either in veins in quartz rocks, or in sand produced by the slow decomposition of such rocks.

The chief gold mines are in California, Australia, South Africa, and in the Ural Mountains.

181. Extraction.—The extraction of gold is a purely mechanical operation. If the ore consists of compact masses, these are previously crushed to a fine powder, and gold is then separated by repeated washings, that is to say, by continually stirring up the fine dust with water in large vessels. On account of its higher specific gravity, the metal sinks to the bottom of the vessels. When finally the gold remains mixed with but a small quantity of sand, it is treated with mercury which dissolves the metal by forming a liquid amalgam that flows away. This is then heated, and the less volatile gold remains behind.

182. Properties.—Gold has a beautiful, yellow colour, and is $19\frac{1}{2}$ times heavier than water. It is soft and the most malleable and ductile of metals; gold leaf can be hammered out until it is no more than $\frac{1}{280000}$ of an inch in thickness.

It is unaltered by the air at any temperature whatever. It is not attacked by any of the acids singly, but dissolves in a mixture of HCl and HNO₃ (Cf. No. 77). It does not get stained with sulphur—a property which gives it an additional advantage over silver.

183. Uses.—Gold is used for coinage and in jewelry; it is then usually alloyed with some copper to increase its hardness. In the form of gold leaf, it is used to gild wood and other substances.

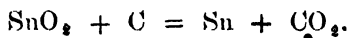
CHAPTER VII.

TIN.

Symb. Sn. (Stannum). At. wt. = 118.

184. Occurrence.—Tin occurs mostly as SnO_2 , called tinstone, which is chiefly found in Cornwall in England, in the Malay Peninsula, and in the island of Banca.

185. Extraction.—Metallic tin is obtained by mixing tinstone with powdered charcoal and heating the mixture; the reaction is represented thus:—



186. Uses.—Tin is a silver-white metal, very malleable, and keeps well in the air; hence its uses:

(1) for *tin-foil*, i.e., tin rolled out into very thin sheets, which are used to protect substances against atmospheric influences;

(2) for *tin-plate*, i.e., iron coated with tin. Sheets of iron perfectly cleaned are dipped into molten tin, which adheres to the iron, and thus protects it against rust;

(3) for *tinning copper vessels*, thus to prevent the copper from forming poisonous copper salts with the acids that enter into the composition of various articles of food;

(4) for *alloys*: *bronze* (Cu and Sn),—*solder* (about equal quantities of Pb and Sn),—*pewter* (a kind of

solder : 4 parts of Sn and 1 of Pb),—*Britannia metal* (an alloy chiefly of antimony and tin, with which Zn is sometimes mixed).

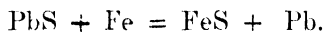
(5) An amalgam of tin is used for *silvering mirrors*. Pins are made with tinned brass wire.

LEAD.

Symb. Pb. (Plumbum). At. wt. = 207. *Tetra*

187. Occurrence.—Lead is found *very rarely free*, and only in very small quantities; usually it is found in the form of lead sulphide PbS , called *galena*, and lead carbonate $PbCO_3$, or *white lead*.

188. Extraction—To obtain metallic lead from $PbCO_3$, it is sufficient to heat it with charcoal; to obtain it from PbS , one method consists in smelting it with iron:—



Generally galena contains a little silver which is to be separated from it (see extraction of silver).

189 Properties.—Lead is a heavy metal ($11\frac{1}{2}$ times heavier than water), soft (it can be scratched with the finger nail), and capable of leaving a greyish mark when drawn across paper (a test for lead). When freshly cut, it exhibits a bluish white surface, that soon becomes tarnished, chiefly in moist air, through the formation of an oxide, which protects the metal against further oxidation.

5) Pure dry air has no action on lead at the ordinary temperature; but when the metal is heated in the air, it is transformed into a yellow oxide called *litharge*.

6) H_2SO_4 and HCl have no action on lead at the ordinary temperature, but HNO_3 attacks it easily with formation of lead nitrate (compare with Hg).

190. Uses.—Because it gets oxidized only on the surface, lead can be used for cisterns, gas or water-pipes, and for roofing.

N.B.—Lead in presence of water and air containing CO_2 , is changed to PbCO_3 , a poisonous substance; hence it is dangerous to keep drinking water in lead vessels.

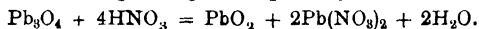
Lead is also used to make solder, type-metal, shot and bullets, because it easily melts and can be poured into moulds.

191. PbO , lead oxide—or *litharge*—is a yellow powder, obtained by heating lead in a current of air.

192. Pb_3O_4 , red lead—or *minium*—is obtained by heating the above PbO at a temperature of about 600°C . in a powerful draught of air.

It is used as a paint and in the manufacture of flint glass.

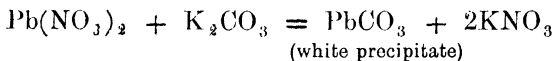
When red lead is warmed with dilute nitric acid, the brown peroxide of lead PbO_2 is formed; this may be separated by filtration, and the filtrate, on evaporating, will deposit crystals of lead nitrate:—



193. PbS , lead sulphide—or *galena*—occurs native in the form of bluish-grey crystals with a metallic lustre. This is the ordinary source of metallic lead.

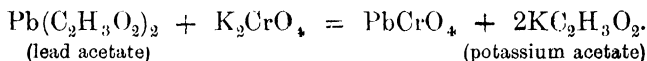
It is obtained in the laboratory as a black precipitate by treating a solution of lead with hydrogen sulphide, H_2S .

194. $PbCO_3$, lead carbonate—or *white lead*—exists usually in the form of a basic salt, $PbCO_3 \cdot Pb(OH)_2$. It is found in nature, and is also prepared artificially by adding an alkaline carbonate to a solution of lead. For example:—



It was formerly much used as a white paint—hence its name of *white lead*.

195. $PbCrO_4$, lead chromate—or *chrome yellow*—is obtained by adding potassium chromate K_2CrO_4 to a solution of lead nitrate or acetate:—



It is much used as a paint.

196. $Pb(C_2H_3O_2)_2$, lead acetate—or *sugar of lead*—is obtained by treating litharge with acetic acid (or vinegar). It is, like the other salts of lead, very poisonous.

When a clean piece of metallic zinc is dipped into a solution of a lead salt, a phenomenon occurs similar to that mentioned above in the case of $CuSO_4$ and Fe, *i.e.*, an exchange takes place between the two metals: for every atom of zinc which dissolves and combines with the acid to form a corresponding salt of zinc,

an atom or an equivalent weight of lead is set free, and is deposited on the remaining zinc, forming shining crystals of lead. This is a test for the presence of lead in a solution, and is known as the *lead tree*, or the *tree of Saturn*—*Arbor Saturni*.

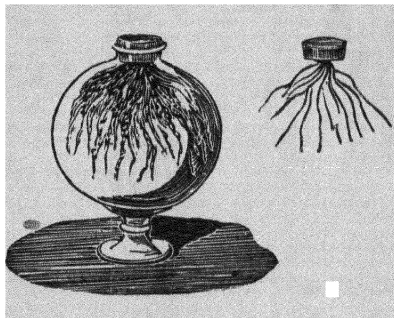


Fig. 71.—Tree of Saturn.

A beautiful lead tree may be obtained by dipping into a dilute solution of lead acetate fine ribbons of zinc spreading out like the branches of a tree. Upon these the metallic lead will be deposited, exhibiting fern-like leaves (Fig. 71.)

CHAPTER VIII.

IRON. *Tetraatomic*

Symb. Fe (Ferrum). At. wt. = 56.

197. Occurrence.—Iron is one of the metals most widely spread on the surface of the earth; it is

present in almost all the rocks ; traces of it are found in the organs of vegetables and of animals, and it forms one of the essential constituents of the blood.

Iron exists *free* in meteorites only. The most important compounds of iron that occur in nature are :—

Fe_2O_3 , iron oxide, or *red hæmatite* ;

$\text{Fe}_2(\text{OH})_6$, iron hydroxide, or *brown hæmatite* ;

Fe_3O_4 , black magnetic iron ore, or *loadstone* ;

FeCO_3 , iron carbonate ;

FeS_2 , iron sulphide, or iron pyrite.

†198. Extraction.—Iron is usually obtained from its oxide ; any other ore being previously roasted to burn off the sulphur or to expel the water or carbon dioxide which it may contain in combination.

Iron smelting, *i.e.*, the extraction of iron by fusing its ore, is performed in very large furnaces (Fig. 72). The ore mixed with limestone, is thrown into the furnace from above in alternate layers with fuel or coal. The fire is kept burning underneath, being at the same time activated by very powerful currents or *blasts* of heated air, directed into the furnace ; hence the name of *blast-furnace*.

The iron oxide is reduced, *i.e.*, deprived of its oxygen, by the coal. The iron gradually sinks in the furnace, fuses and unites with some carbon, thus forming cast-iron. Every now and then the crucible, or the bottom of the furnace, is opened, to allow

the molten iron to flow out: it is then run into cylindrical moulds prepared in the ground, and constitutes what is usually known as *pig-iron*.

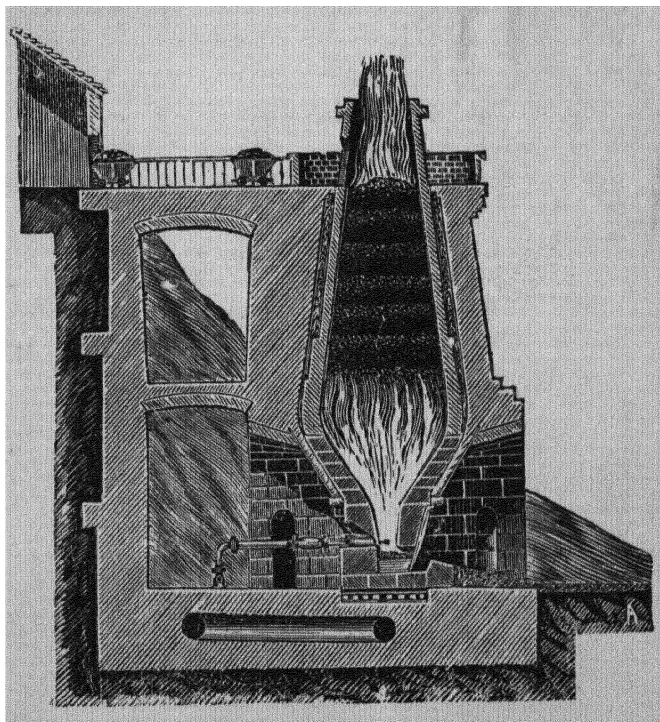
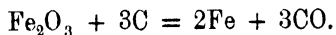


Fig. 72.—The blast-furnace.

The chemical reaction which takes place may be represented by the equation:—



This CO burns to CO₂ at the top of the furnace, or, in more modern arrangements, is utilised for heating the blast.

199. Varieties of iron:—(a) **Cast-iron.**—The iron obtained by reduction in the blast-furnace is called *cast-iron*, because it can be cast into moulds. It contains about 5 p.c. of carbon and several other impurities, as phosphorus, sulphur, and silicon. Cast-iron is very brittle, but melts at a lower temperature than either wrought-iron or steel. It cannot therefore be welded or hammered, but is well suited for casting into moulds. Hence its use for rails, wheels, pillars, gas and water pipes etc. From cast-iron, both wrought-iron and steel are obtained.

(b) **Wrought-iron** or *soft-iron*, is a comparatively pure variety of iron, which may contain but a small fractional percentage of carbon. It is obtained from cast-iron by heating the latter in a special furnace (puddling furnace), in which an arrangement is made to burn away the carbon (and the impurities phosphorus, sulphur, and silicon). It is called wrought-iron because pieces of it can easily be forged or hammered and welded together when red-hot, and can be made to assume any shape we please. It has a fibrous appearance and is not easily broken; hence it is used for many purposes for which cast-iron is unsuitable.

(c) **Steel.**—Steel contains about $1\frac{1}{2}$ or 2 p. c. of carbon, *i.e.*, less than cast-iron, but more than wrought-

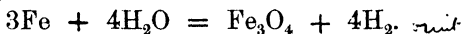
iron. Steel is obtained from wrought-iron by heating the latter in a furnace with carbon. It is obtained from cast-iron by burning away all the carbon and impurities of cast-iron in a special vessel, called *Bessemer's converter*, and then adding to the molten mass a known quantity of cast-iron containing the percentage of carbon required to make steel. This is called *Bessemer steel*.

When heated to redness and suddenly cooled in water, steel becomes intensely hard and brittle and may be used to cut glass. On '*tempering*,' i. e., on carefully reheating this steel to a suitable temperature, and then allowing it to cool slowly, it becomes less hard and more tough, and can be used for cutting tools of various kinds—razors, swords etc.

200. General properties of iron.—Iron is the most important metal for man. In the laboratory pure iron can be got by heating the oxide of iron in a current of hydrogen. Iron is a white metal, remarkable for its hardness; its magnetic power exceeds that of any other substance. Soft-iron retains its magnetism only so long as it is under a magnetic influence; whilst steel keeps magnetism even when removed from the magnetizing agent; hence steel is used for making permanent magnets.

Iron is rapidly oxidized in moist air, becoming covered with a layer of rust, which is ferric hydroxide $\text{Fe}_2(\text{OH})_6$. As this substance is porous, the whole mass of iron is soon converted into rust.

Heated iron decomposes steam (see preparation of hydrogen):—

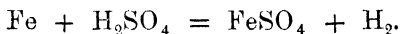


Iron is easily attacked by HCl, dilute HNO₃, and H₂SO₄ (see *ibid*), with formation of the chloride, the nitrate and the sulphate respectively.

Iron forms two series of compounds, *ferrous* and *ferric* :—

FeCl ₂	ferrous chloride	Fe ₂ Cl ₆	ferric chloride
FeSO ₄	„ sulphate	Fe ₂ (SO ₄) ₃	„ sulphate
FeO	„ oxide	Fe ₂ O ₃	„ oxide
Fe(OH) ₂	„ hydroxide	Fe ₂ (OH) ₆	„ hydroxide

201. FeSO₄, ferrous sulphate—or *green vitriol*—is one of the more important iron salts. It can be obtained by treating iron with H₂SO₄ and evaporating :—



Its crystals have a greenish colour; hence the name of green vitriol.

Uses.—It is much used in the manufacture of ink and in dyeing.

202. Tests for iron.—Potassium ferrocyanide and potassium ferricyanide :

Potassium ferrocyanide, or yellow prussiate of potash, gives with ferrous salts a light-blue precipitate.

Potassium ferrocyanide gives with ferric salts, or with ferrous salts to which some nitric acid has been added, a dark-blue precipitate (Prussian blue): if the amount of iron is very small, a blue or greenish colouration only is produced.

Potassium ferricyanide gives with ferrous salts a dark-blue precipitate (Turnbull's blue), indistinguishable by its appearance from Prussian blue.

Potassium ferricyanide gives with ferric salts a green colouration.

PLATINUM. *Tetratomic.*

Symb. Pt. At. wt. = 194.

203. Occurrence.—Platinum is *always* found in the *free* or uncombined state, but is usually alloyed with some rare metals, from which it is separated before use.

204. Properties and uses.—Platinum is a malleable and ductile white metal; it is one of the heaviest substances known ($21\frac{1}{2}$ times heavier than water).

Its infusibility at any temperature of our furnaces (it melts only in the oxyhydrogen flame), the facts that it does not unite with oxygen at any temperature, and that it is not attacked by any acid except by *aqua regia*—all these properties render platinum most valuable and even indispensable to the chemist. It is employed in the form of wire, foil, crucibles, dishes, retorts, etc.

205. PtCl_4 , platinum chloride—is the chief platinum salt, and is obtained by dissolving the metal in *aqua regia*.

Question Papers for Revision.

I.

I. What difference is there between—

- (a) a physical and a chemical change,
- (b) elementary and compound substances,
- (c) a mechanical mixture and a chemical compound,
- (d) atoms and molecules,
- (e) chemical affinity and cohesion,
- (f) metals and non-metals ?

II. Define or briefly explain the following :—

- (a) combining weight of elements ;
- (b) oxide, hydroxide, acid, base, salt ;
- (c) an acid, normal and basic salt.

III. (a) What are the four more important conditions for a chemical change to take place ?

(b) Give examples of other changes that usually accompany chemical changes.

(c) What are the three chief kinds of chemical changes ?

IV. (a) Explain, with the help of examples, the two chief laws of chemical combination.

(b) Illustrate the statement that chemical action takes place more readily between unlike substances, and

(c) that, in a chemical combination, nothing is created and nothing is really lost.

V. (a) Make a drawing of an ordinary candle flame and explain it.

(b) How would you prove that there are unburnt gases in the inner part of the candle flame ?

(c) Explain :—reduction, oxidation, quick and slow combustion.

(d) Why is the hydrogen flame scarcely visible, while the flame of a petroleum lamp is luminous ?

II.

I. (a) Describe one *physical* and two different *chemical* methods of obtaining hydrogen from water.

(b) Mention two different methods of preparing hydrogen, and one of preparing oxygen, *on a large scale*.

(c) How can we obtain pure hydrogen and pure oxygen ?

(d) Mention the three chief uses of hydrogen, and give reasons for its being thus employed.

II. (a) What are the proportions of hydrogen and of oxygen present in water both as regards their volume and their weight ?

(b) Describe an experiment by which the composition of water by weight is ascertained.

(c) What reasons are there for saying that both oxygen and hydrogen are very abundant and very widely distributed in nature ?

(d) What are the chief properties and tests of hydrogen and oxygen ?

III. (a) What are the chief gases dissolved in ordinary water ?

(b) Explain the following : crystallisation, filtration, distillation, solution, deliquescence, efflorescence, For what practical purposes are the first three used ?

(c) Explain what you understand by hard and soft water, permanent and temporary hardness of water.

(d) In what manner would you test whether water is hard or soft? Give a scientific explanation of your experiment.

- IV. (a) What are the chief constituents of sea-water?
(b) Explain the formation of rain-water.
(c) How is common salt extracted from sea-water?

III.

I. (a) How do chlorine and nitrogen occur in nature? Mention some natural compounds in which they are met with?

(b) Describe two methods of preparing chlorine and nitrogen.

(c) What are their chief properties, their tests and their uses?

II. Describe the properties and uses of the allotropic forms of carbon, and indicate the difference between yellow and red phosphorus.

III. (a) Mention the principal substances found in nature containing either sulphur, phosphorus or silicon. Which of these three elements exists *free* in nature?

(b) How is sulphur extracted and purified?

IV. Describe a method of preparing HCl , HNO_3 and H_2SO_4 .

V. (a) Mention the ten more common elementary or compound gases.

(b) Which of them are heavier and which are lighter than air?

(c) Which are combustible, and which incombustible, supporters of combustion?

(d) Which are odourless, and which have a characteristic smell?

(e) How many are there that can be practically collected over water, and which are those that cannot be so collected?

(f) Which of them can be collected by upward or downward displacement of air? For what gases is such a method unpractical?

VI. What happens when blue litmus paper, previously moistened, is introduced into jars containing respectively (1) HCl gas, (2) Cl, (3) SO₂, (4) NH₃, (5) CO₂, (6) N?

IV.

I. (a) How do potassium and sodium occur in nature?

(b) In what form are they found in the ashes of plants? Whence^w and how do they get into the plants?

(c) How are potassium and sodium prepared? Compare their properties.

(d) What are the uses of these two metals, and by what test can you recognise their presence in a compound?

II. (a) Name the four more important compounds of potassium and sodium, with their formulæ and common names.

(b) How are they obtained or how are they manufactured?

(c) Which are found in nature and which are not?

(d) What are the chief uses of each?

III. (a) Compare the affinity of silver for oxygen and for sulphur. What do you conclude from this, as to the mode of existence of silver in nature and as to its uses?

(b) How is silver separated from lead?

(c) What practical advantage is derived from the facts—
and

(1) that silver salts get decomposed by sunlight,

(2) that silver chloride is insoluble?

IV. (a) What are the uses of calcium chloride, calcium hydroxide, bone-earth, magnesium, and zinc ?

(b) What are the various forms and names of calcium carbonate ? In what manner does it render water hard ?

(c) Explain :—slaked lime, milk of lime, lime, lime-water.

V. (a) In what forms are calcium, magnesium and zinc found in nature ?

(b) Compare magnesium and zinc (1) as to their properties, (2) as to the products obtained when these metals are burnt in air, (3) as to the new compounds formed when the products mentioned in (2) are treated with hydrochloric and sulphuric acids ?

(c) Compare the action of water and of hydrochloric acid upon limestone (1) in its ordinary state, (2) after being strongly heated.

V.

I. (a) What are the chief ores of copper and mercury, and how are these metals obtained from them ?

(b) Describe the chief properties and uses of copper and mercury.

II. (a) What happens when CuO is heated (1) with charcoal, (2) in a current of hydrogen ?

(b) What happens when mercury is heated in air and when mercuric oxide is heated ?

III. (a) Describe four different tests by which you can determine the presence of copper in a solution.

(b) How would you prove that a silver coin contains some copper ?

(c) Explain what takes place when a clean iron-rod is dipped in a solution of copper sulphate.

IV. (a) How are the following obtained :—alumina, tin, cast-iron, wrought-iron, steel ?

(b) What is the chemical composition and name of :—blue stone, calomel, corundum, litharge, red lead, white lead, black lead, sugar of lead, chrome yellow, galena ?

(c) What are the uses of tin, lead, platinum, and gold ?

(d) What is brass, bronze, blende, galvanised iron, solder, pewter, Britannia metal, an alloy, an amalgam ?

V. (a) What are the chief ores of iron ?

(b) By what tests can iron be recognised in a solution ?

(c) Describe the differences between the three varieties of iron with regard to *composition, properties, and uses*.

VI. A solution of a salt of each of the following is given :—K, Na, Ag, Ca, Cu, Pb, Fe. How would you find out the presence of each metal in the solution of its salt ?

MISCELLANEOUS.

I. (a) What happens when a lighted taper is introduced into jars filled respectively with hydrogen, oxygen, coal gas, carbon monoxide, carbon dioxide, a mixture of hydrogen and oxygen, of hydrogen and chlorine, of air and marsh gas ?

(b) What are the chief tests for the following gases, and how would you proceed to fill a bottle with each :—Cl, N, NH₃, CH₄, CO₂, HCl, SO₂, H₂S ?

(c) For each gas in (b) mention the more common or practical method of preparation.

II. What is the action of water upon potassium, calcium, slaked lime, plaster of Paris, phosphoric pentoxide, sulphur dioxide, ammonia gas ?

III. Name the elements that are *always* found *free* in nature; those that are *never* found *free*; and those that occur both *free* and *combined*?

(b) For which elements must special precautions be taken to preserve them? State the nature and explain the reason of such precautions.

(c) Classify the non-metals in the order (1) of their increasing atomic weight, (2) of their abundance and wide distribution, (3) of their utility in industry.

IV. (a) Describe a method of preparing:—(1) KOH, (2) soap, (3) phosphorus, yellow and red, (4) coal gas.

(b) Mention five different elementary or compound substances, in the *common practical* preparation of which sulphuric acid is used. Represent the reaction by equations when possible; and for each substance thus obtained mention some important use.

V. (a) What is a base, an acid and an alkali?

(b) What do you understand by the terms:—phosphorescence, allotropy?

(c) Explain briefly and give the chemical composition of:—fire-damp, choke-damp, clay, flint glass, water gas, aqua regia.

VI. (a) How would you distinguish between samples of water containing in solution one of the following:—(1) KCl, (2) CaCl₂, (3) CaCO₃, (4) both CaCl₂ and CaCO₃? Explain your answer. Is the water in each case hard or soft? and if hard, is it temporarily hard or permanently hard?

(b) Give three ways of distinguishing sea-water from distilled water.

VII. Write equations for the reactions that take place when—

1. potassium chromate is added to lead acetate;
2. silver nitrate is added to sodium chloride;

3. potassium carbonate is added to lime-water ;
4. steam is made to pass over heated iron ;
5. copper dissolves in nitric acid ;
6. chlorine is made to pass through potassium hydroxide ;
7. hydrochloric acid is added to bleaching powder ;
8. zinc displaces lead in a solution of lead nitrate.

VIII. (a) How much sodium is required to transform 54 lbs. of water into soda ?—With the hydrogen thus obtained, how much metallic copper can be got from black oxide of copper ?

(b) How much coal containing 80 p. c. of carbon is required to reduce 100 lbs. of tinstone ?

(c) Compare the amount of HgO and $KClO_3$ required to obtain in each case 100 grammes of oxygen.

(d) How many litres of ammonia gas can be obtained from 214 grammes of sal ammoniac ?

(Ammonia is $8\frac{1}{2}$ times heavier than H. One gramme of H occupies 11.16 litres).

IX. Name the chief sulphates and nitrates you know. Give their formula, common name and uses.

X. What are the chief constituents of the atmosphere ? How would you prove the presence of each in it ? What difference is there between pure air and its constituents ? How do you account for the fact that the amount of carbon dioxide in the air remains practically the same ?

University of Madras.

MATRICULATION EXAMINATION PAPERS.

CHEMISTRY.

1907.

1. When dilute sulphuric acid is poured on zinc, the hydrogen of the acid is displaced by the metal and zinc sulphate is formed; mention four other reactions between an element and a compound in which one element is displaced by another, write equations for them, and name the compound which is formed in each case.

II. Name (*a*) four elements which oxidise in air at the ordinary temperature, (*b*) four which oxidise only when heated, and (*c*) two which do not oxidise in air; and state in tabular form with regard to each element named in groups *a* and *b* only:—

- (1) how it occurs in nature,
- (2) in what form or forms it is met with,
- (3) whether it is a metal or a non-metal,
- (4) the name of the compound formed by its oxidation.

III. Describe in detail any experiment which shows that the weight of a metal is increased when it is oxidised, and then describe how the metal may be obtained from its oxide.

IV. (1) Describe all that is *seen* to take place when dilute sulphuric acid is heated with iron filings.

(2) How would you produce Prussian blue from a watery solution of green vitriol?

(3) Calculate the least weight of sulphuric acid which is required in order to dissolve 1.4 grammes of iron.

V. Write the following names of substances in a column, and in parallel columns write opposite to each (1) its

composition, (2) its appearance, (3) the method of preparing it, and (4) its use:—

Caustic soda, epsom salts, copper scales, sodium amalgam, mercuric oxide, calcium chloride.

VI. How do we know that coal is the remains of plants which grew long ago on the surface of the earth?

• What substances are given off, and what are left behind, when coal is heated in retorts?

VII. Mention the points of resemblance and difference between zinc, magnesium, tin, and silver.

1906.

I. What change in composition and properties does limestone undergo when heated? How is lime-water prepared and what happens when air from the lungs is blown through it?

II. 1. Describe experiments which show the composition of water (*a*) by weight, (*b*) by volume.

2. Calculate the percentage composition of water from the following:—

Weight of tube with copper oxide before the experiment	40.957
Weight of the same tube after the experiment	30.125
Weight of water formed	12.197

III. By what tests would you distinguish (*a*) hydrogen from coal gas, (*b*) oxygen from air, (*c*) nitrogen from carbonic acid gas, (*d*) chlorine from hydrochloric acid gas?

IV. Describe the preparation of nitric acid and make a large drawing, covering about a page of your answer paper, of the apparatus employed.

V. Describe the appearance of the following substances and name the elements contained in each:—blende, coke, galena, hæmatite, litharge, quartz.

VI. Give the formula, method of preparation and use of each of the following:—silver nitrate, calcium chloride, lead chromate, caustic soda.

1905.

I. If you were given samples of rain water, sea water, hard chalk water, and gypsum water, what tests would you apply to identify each?

II. Make a drawing of a candle flame, covering one half sheet of your answer paper, and describe what is going on in the different parts of the flame.

III. Describe the preparation and properties of carbonic acid gas, and make a large drawing of the apparatus used to prepare and collect the gas.

IV. Some sand, copper sulphate and alum are thrown into a vessel containing boiling water. Explain what happens, and describe how you would separate each of the substances from the water.

V. Give, in tabular form, the name, chemical formula and method of making the following substances:—Epsom salt, nitre, chrome yellow, green vitriol, and Glauber's salt.

VI. A two-anna piece weighing 22.5 grains was dissolved in nitric acid and a solution of common salt added; the precipitate when dried was found to weigh 27.3 grains. Calculate the percentage of silver in the coin; and state and explain the law upon which the calculation is based.

[Ag. = 108, Cl. = 35.]

1904.

I. (a) Why is it necessary to take care that all the air has been expelled from a hydrogen-generating flask before igniting the hydrogen collected from it?

(b) How is the presence or absence of air in the flask ascertained?

(c) Draw a full-page diagram showing the generating flask and the manner of examining the purity of the gas in the flask.

II. Green leaves are placed in two bottles, one of which is filled with fresh spring water and the other with fresh boiled water; both the bottles are then inverted over basins containing respectively spring water and boiled water, and they are placed in sunlight. Bubbles of oxygen gas collect on the leaves in one bottle but not in the other.

(a) State in which bottle the bubbles are seen.

(b) Give the reason for the difference.

(c) Give a chemical test by which this reason may be confirmed.

III. Describe *what is seen* to take place, and name the products formed when—

(a) flowers of sulphur are heated to boiling with copper turnings;

(b) oxide of copper is heated in a bulb tube through which hydrogen is passed;

(c) mercury oxide is heated in a test tube;

(d) nitric acid is poured on copper turnings;

(e) a piece of yellow phosphorus, and a piece of sodium, are each placed in a separate dish of water.

IV. You are given three samples of water, in one of which gypsum is dissolved, in another chalk, and in the third carbonic acid. Explain how you would distinguish between these three waters.

V. (a) Without describing any experiments, state the experimental facts and the reasons for expressing the composition of water by the formula H_2O .

(b) If 8, instead of 16, is taken as the combining weight of oxygen, what would be the formula for water, and why?

(c) 117 grammes of water are formed when hydrogen is passed over a heated oxide; calculate (1) the loss of weight sustained by the oxide, and (2) the weight of the hydrogen converted into water.

VI. (a) Explain why sulphur is called a non-metal and potassium a metal.

(b) Draw up a list of the most important non-metals and state opposite to each in parallel columns:—

(1) how it occurs in nature;

(2) whether it occurs in more than one form; and if so, name the forms;

(3) whether it is inflammable in air; and if so, name the compounds formed when it burns.

VII. (a) A white powder may be either chalk, alum, or washing soda. How would you identify it without applying chemical tests?

(b) How could the metal be separated from each of the following substances, mentioning a different method for each:—lead acetate, hæmatite, copper sulphate, tinstone?

1903.

I. (1) Which of the following elements are combustible, and which are incombustible:—zinc, chlorine, silver, sulphur, iron, nitrogen, phosphorus, sodium?

(2) Arrange those which are combustible in the order of their combustibility, *i.e.*, the readiness with which they burn in air.

(3) Write the formula of the compound formed when each burns.

II. (1) Make a sketch, occupying one sheet of your answer paper, of the apparatus you have seen used to show that water is produced when hydrogen burns in the air.

(2) If one litre of hydrogen weighs 0.0896 gramme, and oxygen is sixteen times as dense as hydrogen, how many litres of each gas are required to produce fifty grammes of water?

III. (1) Describe (a) the appearance, (b) the properties, of the two common varieties of phosphorus.

(2) What grounds are there for asserting that these two substances are (a) elementary bodies, (b) one and the same element?

IV. When hydrogen is passed over heated copper oxide, it displaces the copper and forms hydrogen oxide. Mention three other reactions in which an element is obtained by displacement by another element. Write equations for two of these reactions.

V. (1) What is meant by the *ore* of a metal?

(2) Write the name and state the chemical composition of one ore of each of the following metals:—iron, lead, zinc, copper, tin.

(3) Explain why the metals aluminium and magnesium should be more costly than lead and zinc, although their ores are far more abundant.

VI. You are given a thin strip of a white metal and told to ascertain whether it is tin or zinc. How would you proceed?

1902.

I. (a) Explain the terms *hardness* and *softness* as applied to water.

(b) What are the different kinds of hardness, and how are they caused?

(c) What kind of hardness can be removed, and how ?

(d) Describe clearly the method of testing a water in order to ascertain whether it is hard or soft ?

II. Describe the experiment by which it can be shown that—

1. chlorine has bleaching properties,
2. vegetable matter contains carbon,
3. metals become heavier by oxidation.

Mention, without describing, the experiments which you have seen performed to show

4. the decomposition of water ;

and, in each case, state what the experiment teaches as to the composition of water.

III. How do you know—

1. that the composition of chalk is quicklime and carbonic acid ; and
2. that it is a compound and not a mixture of these constituents ?

IV. Draw up a list of the more important elements classifying them into metals and non-metals.

Mention the properties of elements to which the term *metallic* is applied.

Point out the differences which you have noticed between any oxides of non-metals and those of metals.

V. Describe all the changes which occur, and name all the compounds formed, when—

- (1) black oxide of copper is heated in a stream of hydrogen,
- (2) a piece of silver coin is warmed with nitric acid,
- (3) water is poured over quicklime,
- (4) powdered antimony is thrown into chlorine.

VI. What is "fire-damp," and under what circumstances does it cause an explosion?

Why are people suffocated by the air in a coal mine immediately after such an explosion?

Explain the principle of Davy's safety-lamp.

VII. Describe the appearances, the preparation, and the use, of each of the following substances:—

Glauber's salt, bone ash, copper oxide, sodium amalgam, red oxide of mercury.

By means of the last two substances illustrate the statement "Chemical combination takes place most readily between those bodies which least resemble one another."

1901.

I. How is clear lime-water prepared? Describe and explain the changes which take place when (1) carbon dioxide is passed into lime-water, (2) the gas is continued to be passed into lime-water for a long time, (3) the resulting liquid in the second case is boiled.

II. What takes place when a piece of sodium is thrown into a basin of cold water? Compare this with what takes place when potassium is thrown on cold water, and account for the difference.

III. (1) Describe a single experiment by which it can be shown that hydrogen burns with a pale-blue flame and also that it does not support the combustion of a taper. (2) Make a half-page sketch of the apparatus as seen in actual use during the experiment.

IV. (1) How can the metal be obtained from iron rust?
(2) How would you produce Prussian blue from a solution of green vitriol? Give the reason for any precaution which may be necessary for the success of the test.

(3) Fifty-three grammes of blue stone are dissolved in water; calculate the least weight of iron required in order to throw down all the copper present in the solution.

(4) How can you tell when all the copper has been thrown down?

V. (1) What is meant by the combining weight of an element?

(2) When hydrogen is passed over an excess of heated oxide of copper, the oxide loses 60 grammes in weight, and 67.5 grammes of water are formed. From these data calculate the proportion by weights of oxygen to hydrogen in water.

(3) State whether the numbers so obtained are the same as the combining weight of these elements; and, if not, give reasons for the difference.

VI. With reference to sulphur and to phosphorus, state (1) how each occurs in nature, (2) in what forms each is known, (3) in what respects the several forms of each element differ from each other, (4) the purposes for which they are used; and (5) describe in detail what takes place when each is heated in a spoon over a flame.

[H=1; O=16; S=32; Fe=56; Cu=63.]

1900.

I. (1) What chemical change takes place when chalk or limestone is "burnt"?

(2) Mention two respects in which the substance left behind differs from the chalk or limestone.

(3) Explain the method of calculating the quantity of this substance which can be obtained from any given quantity of chalk.

II. (1) Make a sketch, occupying one sheet of your answer paper, of the apparatus you have seen used to prepare nitric acid.

(2) By what properties may this acid be distinguished from sulphuric acid ?

III. (1) Name all the substances produced when sulphuric acid is added to each of the following:—common salt, magnesia, iron, chalk, saltpetre, caustic potash.

(2) Write equations to represent any two of these changes.

IV. You are given a powder and told to ascertain whether it consists of bluestone mixed with alum, or of the first substance only. How would you proceed ?

V. (1) What is sodium amalgam ?

(2) How is it prepared ?

(3) What takes place when it is brought into contact with water ?

VI. Mention (without going into details) the experiments you have seen performed to prove—

(1) that oxygen is necessary for the combustion of a candle ;

(2) that carbonic acid gas and water are produced when the candle burns ;

(3) that the carbonic acid and water weigh more than the burnt candle.

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