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AND CHARING CROSS.

THE

ELEMENTS OF CHEMISTRY,

INORGANIC AND ORGANIC.

BY

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LATE LECTURER IN CHEMISTRY AND PHYSICS TO THE EVENING CLASSES
OF THE ROYAL POLYTECHNIC INSTITUTION,
ETC. ETC.

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society, and invest with power a people that have never been educated to use it.

I am not insensible to the objections which are frequently urged against the introduction of new subjects of instruction; but it has often struck me that some of the subjects usually taught in schools might be got rid of with great advantage to the children. The great object of education with the working classes is its practical value, and they rarely rise above this idea; nor do I think it so low and grovelling as some have described it. Unless you can instruct a boy in some of those principles and things which will help him through life, and supply him with new resources in the time of difficulty, infusing a new sentiment into his labour, and relieving him from the monotony of mere mechanical manipulation,—unless this is done, education loses much of its usefulness, and fails in the estimation of those for whose benefit it is intended.

I would not, on this account, ignore everything of an abstract character. A man who can demonstrate the first twenty propositions in Euclid, is raised to an intellectual level above that at which he stood before. But persons who have little capacity for mathematical investigation, often distinguish themselves in the sciences of experiment and observation; and these sciences are not without their logic in the hands of a good teacher. Working men are almost always experimentalists; a difficulty or failure is overcome by trying some other method; they seldom reason out a result,—indeed, results in natural science are rarely so obtained,—and if they reason at all, it is after the result is accomplished.

I have endeavoured to render this book as practical as possible; and am not aware that anything of importance is omitted. In the hands of an intelligent teacher, it will admit of much amplification.

St. John's Hill, Wandsworth.

PREFACE.

I HAVE prepared this little work in the hope that it may be found useful in promoting the scientific education of the senior classes in Middle-class Schools. I trust it will not be without its value to pupil teachers, schoolmasters, and the students of classes in Mechanics' Institutions.

Having been engaged as a teacher of physics, both to boys and adults, I have always found the experimental sciences very popular subjects of instruction; and if this little book should contribute, in however small a degree, towards the scientific industrial education of the working classes, or stir up in some obscure worker a desire to learn, I shall be content to have done something for that class which has always had my warm and earnest sympathy.

For some years I have taken a deep interest in Working Men's Institutions. I believe that many of them contain the elements of great usefulness, and may yet be made to take an important part in the secondary education of the working classes.

The true elevation of the working classes chiefly depends on their own efforts, judiciously aided by the State. A comprehensive and liberal system of education would do much to remove many of those disadvantages under which they have hitherto laboured.

The political liberty of the present day is quite equal to the intellectual condition of the people. To advance the one and neglect the other is to invert the natural order of

In preparing this Edition I have sought the advice of several Certificated Teachers of Elementary Science, and hope I have succeeded in producing a book worthy the acceptance of those engaged in teaching and learning the Elements of Chemical Science.

St. John's Hill, Wandsworth, S.W.
March, 1863.

ELEMENTARY CHEMISTRY.

THE earth is made up of a variety of solid, liquid, and gaseous materials. Chemistry investigates the relationship between the different kinds of atoms or particles of which these materials are composed. If the earth were composed of only one simple substance, there could be no such science as chemistry.

All substances, whether animal, vegetable, or mineral, are either *simple* or *compound.* Those which cannot be reduced into any simpler form are called *elements,* and these elements combine to form an infinite variety of *compound* bodies. We may illustrate this by the letters of the alphabet, which cannot be separated into any simpler form, but by different combinations form a great variety of *words.* These words can be separated into their letters, as can the compound bodies into the elements of which they are composed. It is only compound bodies which can thus be separated. Sulphur, iron, zinc, copper, oxygen, hydrogen, cannot be resolved into any simpler substance: they are therefore called elementary bodies. Water can be separated into oxygen and hydrogen, and is therefore a compound body, and not an element, as was once supposed. Again, chalk, or carbonate of lime, is made up of three elementary bodies,—a metal called calcium, oxygen, and carbon. Common salt is composed of two elementary bodies,—a metal called sodium, and chlorine. The great majority of compounds are formed by the union of two, three, or four elementary bodies. The separation of these compound bodies into simpler forms is called *analysis*, and the formation of compound bodies by the union of simple ones is called *synthesis*.

The Elementary bodies are about seventy in number; but many of them are so rare as not to require special attention. They are divided into metals and metalloids, or non-metallic bodies. The metals have great lustre: are good conductors of heat and electricity. These properties are never associated in a non-metallic body. This division is not based on any exact principles of science, but as a common means of description. The metalloids or non-metallic bodies in the following table are marked with an asterisk. The rare elements are printed in italics.

TABLE OF ELEMENTS.

Names of Elements.	Symbols.	Atomic or Equivalent Weights.	Specific gravities.
Aluminum	Al.	13·7	2·60
Antimony (*stibium*) . .	Sb.	129	6·70
Arsenic	As.	75	5·88
Barium	Ba.	68·5	
Bismuth	Bi.	213	9·82
*Boron	B.	10·9	
*Bromine	Br.	78·26	2·97
Cadmium	Cd.	56	8·60
Calcium	Ca.	20	
*Carbon	C.	6	3·5
Cerium	Ce.	47	
*Chlorine	Cl.	35·5	2·47
Chromium	Cr.	28	5·9
Cobalt	Co.	29·5	8·54
Columbium (*tantalum*) . .	Ta.orCm.	184	
Copper (*cuprum*) . . .	Cu.	31·7	8·89

Names of Elements.	Symbols.	Atomic or Equivalent Weights.	Specific gravities.
Didymium	D.	49·6	
Erbium	E.		
*Fluorine	F.	19	
Glucinum	G.	26·5	3
Gold (*aurum*)	Au.	197	19·26
*Hydrogen	H.	1	0·069
*Iodine	I.	127	4·948
Iridium	Ir.	99	16
Iron (*ferrum*)	Fe.	28	7·8
Lanthanum	La.	47	
Lead (*plumbum*)	Pb.	103·7	11·35
Lithium	Li.	6·52	
Magnesium	Mg.	12·7	1·70
Manganese	Mn.	27·6	6·85
Mercury (*hydrargyrum*)	Hg.	100	13·59
Molybdenum	Mo.	46	8·6
Nickel	Ni.	29·6	8·28
Niobium	Nb.		
*Nitrogen	N.	14	0·9729
Norium	No.		
Osmium	Os.	99·6	10
*Oxygen	O.	8	1·106
Palladium	Pd.	53·3	11·3 to 11·8
Pelopium	Pe.		
*Phosphorus	P.	32	1·77
Platinum	Pt.	98·7	21
Potassium (*kalium*)	K.	39	0·865
Rhodium	R.	52·2	11
Ruthenium	Ru.	52·2	8·6
**Selenium*	Se.	39·5	4·28
*Silicon	Si.	21·3	
Silver (*argentum*)	Ag.	108·	10·47
Sodium (*natrium*)	Na.	23	0·972
Strontium	Sr.	43·8	
*Sulphur	S.	16	2·07
Tellurium	Te.	64·2	6·25
Terbium	Tb.		
Thorium	Th.	59·6	
Tin (*stannum*)	Sn.	59	7·29

Names of Elements.	Symbols.	Atomic or Equivalent Weights.	Specific gravities.
Titanium	Ti.	25	5·3
Tungsten (*Wolfram*)	Tn. or W.	95	17·6
Uranium	U.	60	
Vanadium	V.	68·6	
Yttrium	Y.		
Zinc	Zn.	32·6	7·86 to 7·1
Zirconium	Zr.	22·4	

Some of the elements have a popular name: sulphur is called brimstone; mercury, quicksilver, Carbon when crystallised is a diamond; when prepared artificially, it is known as charcoal, coke, lamp-black, and bone-black. The Latin names of all the metals terminate in *um*. Five of the elementary bodies are gaseous under ordinary pressure and temperature;—hydrogen, oxygen, nitrogen, chlorine, and fluorine. Two are liquid,—mercury and bromine. The others are solid.

The elementary bodies are represented by symbols. These symbols are of great importance in expressing chemical changes.

The student should learn the symbols and combining equivalents of the commonly occurring elements. The elementary bodies forming a compound are frequently called its *constituents*, or *components*, and the act of union is called *chemical combination*, whilst the separation of a compound into its constituents is called *decomposition*.

CHEMICAL AFFINITY.

The power or force by which bodies chemically unite is called *affinity*. It differs from the force of gravitation, which tends to draw all bodies towards the centre of the earth, and operates at all known distances. It also differs from cohesion, which binds like particles of matter together.

The rain falls by gravitation, the particles of iron are held together by cohesion; but in iron rust (or oxide of iron) the oxygen is united to the iron by the force of chemical affinity. Chemical combination may be distinguished from mere mechanical mixture by the formation of compounds having properties different from those of the bodies which have entered into combination. And what is most remarkable, bodies of opposite chemical characters evince the strongest tendency to combination. Chemical affinity may be regarded as that force by which bodies unite, and form compounds having properties totally different to those of the constituents which have entered into combination. This combination is usually attended with an elevation of temperature. Chemical affinity acts with different degrees of force. If sulphuric acid be united with ammonia, and magnesia added to the compound, the ammonia is set free, and the sulphuric acid combines with the magnesia. This union of magnesia and sulphuric acid may be again decomposed with lime; and this compound may be decomposed with soda, then strontia, then baryta. From this it appears that the affinity of sulphuric acid for baryta is strongest, and for ammonia the weakest.

The force of chemical affinity is very much influenced by heat. Some chemical compounds are entirely decomposed by heat; the particles seem to be removed beyond the sphere of chemical attraction by the repulsive force of heat. Heat destroys all organic substances, but other compounds are formed of a more permanent character. Light, also, modifies the force of affinity; if equal volumes of chlorine and hydrogen be mixed in the dark they do not unite, but under the influence of sunlight combination takes place with great violence. The particles of matter attracted are assumed to be in different electrical states, and this is regarded as the basis of all chemical action. Electrical currents are among some of the most powerful causes of chemical decomposition. As chemical combination can only take place when the particles or molecules are brought within the sphere of chemical attraction, which acts at insensible distances, the force of cohesion must, first, be overcome; and this is done by dissolving the substance usually in water: so that the particles may be brought into closer contact. Substances in a solid state rarely combine, even if reduced to powder. Phosphorus with Iodine is an exception. When elements are liberated at the same time, in what is called their nascent condition, compounds are sometimes formed which cannot be effected under other circumstances. If nitrogen and hydrogen be mixed in the same vessel, combination does not take place; but when these gases are set free at the same time, as in the case of manure heaps, ammonia is formed.

LAWS OF CHEMICAL COMBINATION.

Chemical combination takes place according to certain definite laws. These laws are called the *laws of combining proportion:* no circumstances can alter the proportional quantities in which bodies chemically unite.

1. *The same chemical compound invariably consists of the same elements united in the same proportion by weight.* Pure water, no matter where obtained, contains 8 parts of oxygen by weight united to 1 part of hydrogen. Common salt, whether from Poland or Cheshire, contains 23 parts by weight of sodium, to 35·5 parts by weight of chlorine. Carbonate of lime, whether crystallised or amorphous, has the same elements united in the same proportion by weight.

2. *When one body is capable of uniting with another, in several proportions, these proportions bear a simple relation to each other.* This is called the law of multiple proportions. The best illustration of this law is afforded by the five oxygen compounds of nitrogen.

$$\underbrace{N = 14 \quad . \quad O = 8 \quad NO = 14 + 8 = 22}_{\text{Protoxide of nitrogen.}}$$

$$\underbrace{N = 14 \quad . \quad O_2 = 16 \quad NO_2 = 14 + 8 + 8 = 30}_{\text{Binoxide of nitrogen.}}$$

$$\underbrace{N = 14 \quad . \quad O_3 = 24 \quad NO_3 = 14 + 8 + 8 + 8 = 38}_{\text{Nitrous acid.}}$$

$$\underbrace{N = 14 \quad . \quad O_4 = 32 \quad NO_4 = 14 + 8 + 8 + 8 + 8 = 46}_{\text{Hyponitric acid.}}$$

$$\underbrace{N = 14 \quad . \quad O_5 = 40 \quad NO_5 = 14 + 8 + 8 + 8 + 8 + 8 = 54}_{\text{Nitric acid.}}$$

In this example, while the nitrogen remains constant, the oxygen increases by multiples of eight.

Sometimes the members of a group are wanting: if these members should be discovered, they will, doubtless, follow the law just illustrated. Chlorine affords a good example.

$$\underbrace{Cl = 35{\cdot}5 \quad . \quad O = 8 \quad ClO = 35{\cdot}5 + 8 = 43{\cdot}5}_{\text{Hypochlorous acid.}}$$

$$\underbrace{Cl = 35{\cdot}5 \ . \ O_3 = 24 \ ClO_3 = 35{\cdot}5 + 8 + 8 + 8 = 59{\cdot}5}_{\text{Chlorous acid.}}$$

$$\underbrace{Cl = 35{\cdot}5 \ . \ O_4 = 32 \ ClO_4 = 35{\cdot}5 + 8 + 8 + 8 + 8 = 67{\cdot}5}_{\text{Hypochloric acid.}}$$

$$\underbrace{Cl = 35{\cdot}5 \ . \ O_5 = 40 \ Clo_5 = 35{\cdot}5 + 8 + 8 + 8 + 8 + 8 = 75{\cdot}5}_{\text{Chloric acid.}}$$

$$\underbrace{Cl = 35{\cdot}5 \ . \ O_7 = 56 \ ClO_7 = 35{\cdot}5 + 8 + 8 + 8 + 8 + 8 + 8 + 8 = 91{\cdot}5}_{\text{Perchloric acid.}}$$

Here the compounds between the first and second, and the fourth and fifth are wanting. While the chlorine is a constant quantity, the oxygen increases by multiples of eight.

The oxygen compounds of sulphur afford another illustration of this law.

$$\underbrace{S_2 = 32 \ . \ O_2 = 16 \ S_2O_2 = 16 + 16 + 8 + 8 = 48}_{\text{Hyposulphurous acid.}}$$

$$\underbrace{S = 16 \quad . \quad O_2 = 16 \quad SO_2 = 16 + 8 + 8 = 32}_{\text{Sulphurous acid.}}$$

$$\underbrace{S_2 = 32 \quad . \quad O_5 = 40 \quad S_2O_5 = 16 + 16 + 8 + 8 + 8 + 8 + 8 = 72}_{\text{Hyposulphuric acid.}}$$

$$\underbrace{S = 16 \quad . \quad O_3 = 24 \quad SO_3 = 16 + 8 + 8 + 8 = 40}_{\text{Sulphuric acid.}}$$

3. *If one body, A, unite with other bodies, B C D, the quantities of B C D which unite with A represent the proportions in which they unite among themselves, in the event of union taking place.*

No element has a greater range of affinity than oxygen, and many of its compounds are capable of exact analysis. Its combining equivalent is 8.*

Let us take a few illustrations.

O = 8
H = 1 . . . S = 16
I = 127 . . . K = 39
P = 32 . . . Fe = 28
Cl = 35·5 . . . Ag = 108
N = 14 . . . C = 6

These numbers represent the proportions in which these elements unite: thus, 8 of oxygen unites with 1 of hydrogen, with 14 of nitrogen, with 6 of carbon, with 16 of sulphur, 39 of potassium, 28 of iron, and 32 of phosphorus. 1 part by weight of hydrogen goes as far in saturating 8 of oxygen, as 108 of silver, 39 of potassium, 6 of carbon, 32 of phosphorus, and 28 of iron. 16 of sulphur combines with

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* These numbers representing only combining *proportions*, it is necessary to fix upon a standard. For many reasons, amongst which may be mentioned its very low combining proportion, hydrogen has been fixed upon as the standard, and its combining equivalent is taken as unity. And upon analysing water, its most common compound with oxygen, it is found that the weight of the oxygen is always eight times that of the hydrogen. Hence, if the equivalent of hydrogen be taken as 1, that of oxygen will be 8.

28 of iron, 35·5 of chlorine. This law is never departed from in any case. It is very common to speak of such a number of equivalents of one body united to one or more equivalents of another. Sulphuric acid, SO_3, contains one equivalent of sulphur to three equivalents of oxygen. Potash, KO, one of potassium to one of oxygen. Phosphoric acid, PO_5, one of phosphorus to five of oxygen.

4. *The combining equivalent of a compound is always the sum of the combining equivalents of its constituents.* The combining equivalent of sulphuric acid is 40, 16 of sulphur, and 3 times 8 of oxygen; $16 + 8 + 8 + 8 = 40$. The combining equivalent of potash, KO, is 47, 39 the equivalent of potassium, and 8 the equivalent of oxygen; $39 + 8 = 47$. 40 parts of sulphuric acid neutralize 47 parts of potash, forming the sulphate of potash, which has a combining equivalent of $47 + 40 = 87$. Nitric acid, NO_5, has a combining equivalent of $14 + 8 + 8 + 8 + 8 + 8 = 54$; the proportion of nitric acid necessary to neutralize and combine with 47 parts of potash. From this it will be evident that when neutral salts decompose each other, the resulting compound will be neutral. It will be understood that these proportions are by weight.

COMBINATION BY VOLUME.

It is often more convenient to *measure* a certain volume of gas than to weigh it. Gases and volatile vapours combine by volume according to a definite and simple law. Equivalent weights of hydrogen and chlorine occupy at the same temperature and pressure equal volumes. Oxygen occupies only half the volume. If the specific gravity of a gas or vapour, and its combining equivalent, be known, its combining volume may be determined by dividing its equivalent, by its specific gravity. The combining equivalent of oxygen (8) divided by its specific gravity (1·106) $= \frac{8·000}{1·106} = 7·23$, its combining volume. Chlorine equivalent 35·5 divided by its sp. gr. which is 2·470 = $\frac{35{\cdot}50}{2{\cdot}470} = 14{\cdot}37$, its combining volume. Sulphur vapour, sp. gr. $\frac{16{\cdot}000}{6{\cdot}654} = 2{\cdot}40$.

From this it appears that the combining volume of chlorine is double that of oxygen; and sulphur one-sixth of chlorine. Hydrogen, nitrogen, chlorine, bromine, iodine, carbon, and mercury, in a gaseous state, have the same combining volume; oxygen, phosphorus, and arsenic, one-half; sulphur, one-sixth. Compound bodies in a state of vapour yield the same results.

WEIGHT.

By the term weight is meant the whole amount of matter a body contains, and this is generally expressed in grains, ounces, pounds, &c. *Specific weight*, or *specific gravity*, is the number which expresses the ratio which the weight of a cubic inch of the body bears to the weight of a cubic inch of distilled water at a temperature of 60°.

The specific gravity of—

Platina	21
Gold	19
Lead	11
Copper	8
Iron	7

So that in the same bulk we have three times more matter in platina than iron, and iron is seven times heavier than its own bulk of water. In gases, air is taken as 1; and the specific gravities of gases are referred to that standard. Great skill is required in determining the specific gravity of gases or vapours, and nothing short of practice can give accuracy in the results.

Atomic weight, or *equivalent* weight, is a term used to express the proportion in which bodies combine chemically. The atomic theory was first promulgated by Dr. Dalton. It assumes the existence of indivisible particles, molecules, or atoms; that compounds are formed by the union of these atoms, one to one, one to two, one to three, and so on. The compound atom thus formed unites itself to the compound atom of another kind, and combination of the second order results. If it be allowed that the relative weights of these atoms are in the proportion of their equivalent numbers, the hypothesis renders satisfactory reasons for the laws of combining proportion. Chemical compounds must be definite in their composition, always containing the same number of atoms. The law of multiple proportion is also easily explained. An atom of nitrogen unites with one, two, three, four, and five of oxygen. If the black circle represent the nitrogen atom, and the open circles the oxygen, we may help our understanding by the following diagram :—

NO .	Protoxide of Nitrogen . .	
NO_2 .	Binoxide of Nitrogen . .	
NO_3 .	Nitrous acid . . .	
NO_4 .	Hyponitric acid . . .	
NO_5 .	Nitric acid	

CHEMICAL NOMENCLATURE.

In early days, fanciful names were given to chemical compounds; many of these names have been retained in common use, such as oil of vitriol, aqua fortis, spirits of hartshorn, spirit of salt, cream of tartar, sugar of lead, blue stone, flower of zinc. Lavoisier introduced a new nomenclature, which satisfied for some time all the requirements of science. In organic chemistry, which is a new department of the science, the defects of this nomenclature are mostly felt. Some of the elementary substances have names referring to some peculiarity of the body: such as barium, antimony, chromium, chlorine, &c. When any of the non-metallic elements combines with a metal, the compound ends in *ide*. So we have ox*ides*, brom*ides*, chlor*ides*, iod*ides*, carb*ides*, phosph*ides*, and sulph*ides*. Carb*uret*, sulph*uret*, and phosph*oret* are sometimes used.

A compound of oxygen and iron, FeO, is the oxide of iron; by replacing the oxygen with chlorine, sulphur, or iodine, we get the chloride, sulphide, or iodide of iron. If the non-metallic bodies combine with each other, they are similarly named; as, sulphide of carbon, chloride of sulphur, phosphide of hydrogen.

The oxides are generally divided into three classes. The oxides which resemble potash, soda, and lime, are called *basic oxides*. The oxides of the second class are the very opposite in chemical relationship to the first, and are called *acids*. Phosphorus, sulphuric, chloric, and nitric acid may represent this class of oxides. There is a third class called *neutral oxides*, of which water and the binoxide of manganese are good illustrations. It frequently happens that a metal unites with oxygen in several proportions: in such cases, one oxide has generally a strongly marked basic character; to this the term *prot*oxide is given. The compounds next are called *bin*oxide or *deut*oxide, *ter*oxide or *trit*oxide; the highest oxide, not having acid properties, is called the *per*oxide. Any compound containing less oxygen than the protoxide is called a *sub*oxide. If an oxide should occur between the protoxide and binoxide, it is called the *sesqui*oxide.

If the letter M represent a metal, and if oxygen unite with it in several proportions, we shall have—

M_2O suboxide,
MO protoxide,
M_2O_3 sesquioxide,
MO_2 binoxide,
MO_3 teroxide,
MO_4 quadroxide

By substituting chlorine, or any of the other metalloids for the oxygen, we shall have the subchloride, protochloride, bichloride, &c.

When the metalloids unite with oxygen, compounds are generally formed which have acid characters. The acid compound containing most oxygen has the termination in *ic;* as, sulphur*ic*, nitr*ic*, chlor*ic*, phosphor*ic*. The compound containing the next less quantity of oxygen terminates in *ous;* phosphor*ous*, chlor*ous*, sulphur*ous*. When other compounds were discovered the prefixes *hyper* and *hypo* were introduced.

Hypophosphorous acid	PO
Phosphorous acid	PO_3
Phosphoric acid	PO_5
Sulphurous acid	SO_2
Sulphuric acid	SO_3
Hyposulphurous acid	S_2O_2
Hyposulphuric acid	S_2O_5
Nitrous acid	NO_3
Hyponitric acid	NO_4
Nitric acid	NO_5
Hypochlorous acid	ClO
Chlorous acid	ClO_3
Hypochloric acid	ClO_4
Chloric acid	ClO_5
Perchloric acid	ClO_7

The prefix *per* is usually given to indicate *the highest* or *most intense*. In the case of metals it is given to the highest oxide destitute of acid properties.

The nomenclature of a salt is derived from that of the acid. When the acid ends in *ic* the salt ends in *ate*—sulphate of lime. When it terminates in *ous* the salt ends in *ite*—nitrite of soda. Hyponitrate, hyposulphate, hyponitrite, hypophosphite, &c.

BASIC OXIDES.

It has been stated that the metallic oxides are divided into bases, acids, and indifferent bodies. Several of the metals combine in different proportions with oxygen, and the chemical character of these compounds is determined in a great measure by the quantity of oxygen they contain. The oxides of manganese afford a good illustration. The protoxide, MnO, is a powerful base, the sesquioxide, Mn_2O_3, is feebly basic. The binoxide or peroxide, MnO_2, is an indifferent body; while MnO_3, manganic acid, and Mn_2O_7, permanganic acid, are acid bodies. Very few of the metals form acid oxides, or indifferent bodies; but the next oxide after an indifferent body, has acid properties. The most important basic oxides are the following. The symbols of these bases should be committed to memory.

LIST OF BASIC SUBSTANCES.

Names.	Colour.	Symbols.
Potash	White	KO
Soda	Ditto	NaO
Ammonia	Colourless	NH_4O
Baryta	White	BaO
Strontia	Ditto	SrO
Lime	Ditto	CaO
Magnesia	Ditto	MgO
Alumina	Ditto	Al_2O_3
Sesquioxide of chromium	Green	Cr_2O_3
Protoxide of iron	Black	FeO
Sesquioxide of iron	Brownish red	Fe_2O_3
Oxide of zinc	White	ZnO
Protoxide of manganese	Greenish grey	MnO
Protoxide of nickel	Grey	NiO
Protoxide of cobalt	Ditto	CoO
Oxide of silver	Brown	AgO
Suboxide of mercury	Black	Hg_2O
Protoxide of mercury	Red	HgO
Oxide of lead	Yellow or reddish yellow	PbO
Oxide of cadmium	Brown or yellowish brown	CdO
Oxide of copper	Black	CuO
Teroxide of bismuth	Yellow	BiO_3

Names.	Colour.	Symbols.
Protoxide of tin	Black	SnO
Binoxide of tin	Light straw colour	SnO_2
Teroxide of antimony	Greyish white	SbO_3
Teroxide of gold	Brown	AuO_3
Binoxide of platinum	Ditto	PtO_2

Most of these bases are insoluble in water: those which are soluble restore the blue colour of litmus, which has been reddened by an acid. Soda, potash, and ammonia are called *alkalies:* the latter is called the volatile alkali. These bodies are soluble in water, and neutralise the strongest acids. Baryta, strontia, lime, and magnesia, are called *alkaline earths;* except magnesia, they are all soluble in water, but less so than the alkalies. Alumina is insoluble in water, and belongs to the earths.

It may appear strange that the terms protoxide of potassium, protoxide of sodium, protoxide of ammonium, are not applied instead of potash, soda, ammonia, baryta, and lime. There can be no doubt but protoxide of potassium is the correct term for potash, but when these names were given they were considered simple bodies, and the terms potash, lime, &c. still continue in use.

In writing out simple chemical formulæ, figures frequently occur. A little figure to the right merely shows the number of equivalents of the element taken, which stands to the left of the figure. Carbonic oxide, CO, means one equivalent of carbon and one of oxygen. Carbonic acid, CO_2, means one equivalent of carbon and two of oxygen. Sesquioxide of iron, Fe_2O_3, means two equivalents of iron and three of oxygen. The suboxide of mercury, Hg_2O, means two equivalents of mercury and one of oxygen. The subiodide of mercury, Hg_2I, means two equivalents of mercury and one of iodine. The terchloride of gold, $AuCl_3$, means one equivalent of gold and three of chlorine. The terhydride of arsenic, AsH_3, means one equivalent of arsenic and three of hydrogen. Hydrides must not be confounded with hydrates. The former is a union of hydrogen with some other body; the latter is a union of water with some substance, as the hydrate of lime, CaO,HO. When compounds of the metalloids and metals are written symbolically, the metal is placed first;—Chloride of calcium, $CaCl$.

Exercise.

Write out the names of these compounds:—

CaS.
Fe_2Br_3.
AgI.
$SnCl_2$.
$PbCl$.
NH_4Cl.
CaF.
$SnCl$.
NH_4S.
$AuBr_3$.
Cr_2O_3.
FeS.
Mn_2O_3.
Fe_2S_3.
CaF.

When the metalloids combine, the hydrogen is placed before the other metalloid, except in the case of phosphorus and carbon, as C_4H_4,PH_3.

ACID BODIES.

An acid body was formerly defined as an oxidised compound which had a sour taste, turned blue litmus red, and neutralised alkalies. This definition is not sufficient: because we know many compounds without oxygen that have all the properties of an acid, as hydrochloric acid, HCl; and some bodies, such as carbonic and boracic acid, cannot neutralise the alkaline reaction of potash. Chemists now regard as acids all those compounds which unite with potash and soda, and give rise to compounds similar in their composition and character to the salts which sulphuric, nitric, or some admitted acid forms when combined with these alkalies.

Acids were formerly, and are still in many chemical works, divided into two classes, viz. (1) Oxygen acids, or those which contain oxygen as the acidifying element, and (2) Hydrogen acids, those which contain hydrogen, but have no oxygen in their composition. Of the former class, or those considered as oxyacids, the following may be given as illustrations:—

Sulphurous acid	SO_2
Sulphuric acid	SO_3
Nitric acid	NO_5
Chloric acid	ClO_5
Oxalic acid	C_2O_3
Phosphoric acid	PO_5
Carbonic acid	CO_2
Silicic acid	SiO_3

These were termed "anhydrous acids," as being destitute of water in their composition. It is, however, a remarkable fact that most, if not all, of the so-called "anhydrous acids" either do not exhibit acid properties at all, or only to a very doubtful extent. For instance, the bodies SO_3 and PO_5 are white snowy-looking solids, which may be handled or worked about with the fingers, and which do not in the slightest degree affect the blue colour of litmus paper. In fact they have no more acid properties than the body they resemble. Similar remarks might be made of all the others. The term "acid" therefore as applied to these compounds, is erroneous, and for this reason "anhydride" has been substituted. The anhydrides are characterised in general by an intense attraction for the elements of water, with which they combine, forming true acids: thus if sulphuric anhydride SO_3, be thrown into water it dissolves with a loud hissing sound, and the production of ordinary sulphuric acid or oil of vitriol; in like manner nitric anhydride NO_5, forms nitric acid; phosphoric anhydride PO_5, forms phosphoric acid, &c. By those who consider the anhydrides as anhydrous acids, the real acids are called "hydrated acids," and are supposed to consist of the anhydrous acid in combination with water; thus,

Hydrated sulphuric acid	HO,SO_3
,, nitric acid	HO,NO_5
,, phosphoric acid	$3HO,PO_5$
&c.	&c.

Of the "hydrogen acids," the following may be taken as examples:—

Hydrochloric acid	HCl
Hydrobromic acid	HBr
Hydriodic acid	HI
Hydrocyanic (prussic) acid . .	HCy or HC_2N

Both "oxygen acids" and "hydrogen acids" are thus seen to contain the element hydrogen as a necessary constituent, and as they all possess exactly similar properties, and cannot, except by analysis, be distinguished from each other, there can therefore be no doubt that the real acidifying element is *hydrogen*, not oxygen.

For this and many other reasons founded on the study of the actions of acids on metals and metallic oxides, the unphilosophical division of acids into two classes is very generally discarded, and all acid bodies are now considered by the most eminent chemists to be constituted according to one type, the type of the so-called hydrogen acids, and to consist essentially of hydrogen in union with a body called an "acid radical." By the term "acid radical" is to be understood *any simple or compound body capable of forming with hydrogen a substance possessed of acid properties.* As examples of *simple* radicals we may mention chlorine Cl, bromine Br, sulphur S, &c., while cyanogen C_2N or Cy, and the compounds represented by SO_4, NO_6, PO_8, &c. may be taken as examples of *compound* radicals.

An acid then may be defined as *the hydrogen compound of a simple or compound acid radical.* The following table represents the composition of some of the most important acids;—

Acids with Compound Radicals.

Sulphurous acid*	Solid (cryst) .	H,SO_3+8aq
Sulphuric acid	Fluid . . .	H,SO_4

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* Sulphurous anhydride is a colourless gas of suffocating odour. Carbonic and chromic acids are unknown in a free state; their

Nitric acid (aquafortis)	Fluid	H,NO_6
Chloric acid	Fluid	H,ClO_6
Oxalic acid	Solid	H,C_2O_4 or H,O
Phosphoric acid (ordinary)	Fluid	H_3,PO_8
Carbonic acid (choke damp)*		
Arsenious acid	Solid	H_2,AsO_5
Arsenic acid	Solid	H_3,AsO_8
Chromic acid	Solid	
Boracic acid	Solid (cryst.)	H,BO_4
Silicic acid (silica, quartz)*	Solid	
Hydrocyanic acid	Fluid	HCy

Acids with Simple Radicals.

Hydrochloric acid	Gas	HCl
Hydrobromic acid	Gas	HBr
Hydriodic acid	Gas	HI
Hydrosulphuric acid	Gas	HS
Hydrofluoric acid	Gas	HF

Some acids contain more than one equivalent of hydrogen, such are termed *polybasic* acids; *e.g.* monobasic phosphoric acid H,PO_6; bibasic phosphoric acid H_2,PO_7; tribasic arsenic acid H_3,PO_8, &c. These will be again referred to under "salts."

SALTS.

The division of salts into the two classes of oxysalts and haloid salts, corresponding to oxygen acids and hydrogen acids, is still retained in most elementary works, and it therefore is necessary to say a few words about them. Oxysalts are formed by the union of a base with an oxygen acid (anhydride), as KO,SO_3, sulphate of potash. Haloid salts are formed by the radical of a hydrogen acid united to a metal, as NaCl, chloride of sodium. There are the same objections to this division of salts as to the corresponding division of acids; and therefore most chemists now view all

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anhydrides CO_2 and CrO_3, are well known; CO_2 is a gas, CrO_3 a crystalline solid. Gelatinous silica is supposed to be the real silicic acid, and its composition H,SiO_4.

salts as formed on the hydracid type exactly similar to the acids. Accordingly sulphate of potash is represented as K,SO_4; nitrate of copper, Ca,NO_6; carbonate of soda, Na,CO_3; chloride of sodium, $NaCl$, &c.

A salt may therefore be defined as the metallic compound of a simple or compound salt—or acid—radical.

According to this theory (which is sometimes called the binary theory), an acid must be considered as a salt in which hydrogen occupies the base of the metallic element. Most chemists in fact regard hydrogen as a metal in the gaseous state, and its affinities certainly favour this supposition.

A salt is formed from an acid by the simple *displacement of the hydrogen* of the acid, and the *substitution* in its place of a metal, or some compound which will act the part of a metal. When an acid acts on a metal, a simple case of substitution occurs, as may be seen in the following instances where HCl or H,SO_4 acts on sodium, Na:—

$$\begin{array}{llll} Na + HCl & = NaCl & + H. \\ Na + H,SO_4 & = Na,SO_4 & + H. \end{array}$$

or, substituting zinc, Zn, for the sodium,

$$\begin{array}{llll} Zn + HCl & = ZnCl & + H. \\ Zn + H,SO_4 & = Zn,SO_4 & + H. \end{array}$$

Whenever then an acid acts on a metal its hydrogen is displaced, and, in most cases, escapes in a free state, while a salt of the metal is formed. When, on the other hand, an acid acts on a base, a different reaction ensues; the displaced hydrogen at the moment of its liberation, entering into combination with the oxygen of the base; forming water thus:—

$$\begin{array}{llll} NaO + HCl & = NaCl & + HO. \\ NaO + H,SO_4 & = Na,SO_4 & + HO. \\ ZnO + HCl & = ZnCl & + HO. \\ ZnO + H,SO_4 & = Zn,SO_4 & + HO. \end{array}$$

Polybasic acids require as many equivalents of base to form neutral salts as they contain of water; or to speak according to the language of the binary theory, they require as many equivalents of metal as they contain of replaceable hydrogen; *e.g.*, neutral bibasic phosphate of soda is $2NaO,PO_5$, or Na_2,PO_7; neutral tribasic phosphate of soda is $3NaO,PO_5$, or Na_3,PO_8, &c. Many of the bases unite with water, forming *hydrates*; such as the hydrate of potassium, KO,HO; the hydrate of calcium, CaO,HO. By some chemists, the water is in these cases supposed to act the part of a feeble acid, and they consequently represent such compounds as salts; as hydrate of potassium KHO^2; hydrate of calcium, $CaHO_2$, &c.

Nearly all salts are solid at the common temperature, and have different affinities for water. The nitrate of lime and the nitrate of magnesia, the chloride of calcium and the silicate of soda, attract sufficient moisture from the atmosphere to become liquid: others undergo scarcely any change. Salts differ greatly in their degrees of solubility; some dissolve in less than their weight of water, while others are quite insoluble. Soluble salts crystallize out when their solutions are evaporated. The form which they assume in crystallising is always constant, and forms a character by which they are frequently distinguished. Some salts, in crystallising, unite with a definite quantity of water, called water of crystallisation. This water of crystallisation is held by a very feeble affinity: a moderate heat is sufficient to expel it. A portion of water is frequently retained with such force that it cannot be expelled without decomposing the salt. When this is the case, it is called essential water, and is expressed by the symbol HO. When the water is non-essential, it is often expressed by the Latin word *aqua*. The following formulæ of salts must be studied. The letter M stands for a metal.

General Formulæ of the more important Salts,

Sulphites . . .	MO,SO_2	or	M,SO_3
Sulphates . . .	MO,SO_3	or	M,SO_4
Nitrates	MO,NO_5	or	M,NO_6
Chlorates . . .	MO,ClO_5	or	M,ClO_6
Oxalates* . . .	$MO,\bar{O}$	or	$M\bar{O}$
Phosphates . . .	$3MO,PO_5$	or	M_3,PO_8
Carbonates . . .	MO,CO_2	or	M,CO_3
Arsenites . . .	$2MO,AsO_3$	or	M_2,AsO_5
Arseniates . . .	$3MO,AsO_5$	or	M_3,AsO_8
Chromates . . .	$MOCrO_3$	or	M,CrO_4
Borates	MO,BO_3	or	M,BO_4
Silicates	MO,SiO_3	or	M,SiO_4

Chlorides (muriates).	MCl	Iodides	MI
Sulphides (sulphurets)	MS	Bromides	MBr
Fluorides	MF	Cyanides	MCy

It is seen by these formulæ how the termination *ic* and *ous* in the acids change into *ite* and *ate* in salts.

Write out the following salts:—

$NaO,SO_3(Na,SO_4)$.
$BaO,SO_3(Ba,SO_4)$.
$MgO,SO_3(Mg,SO_4)$.
$FeO,SO_3(Fe,SO_4)$.
$KO,SO_2(K,SO_3)$.
$NH_4O,SO_2(NH_4,SO_3)$.
$SrO,ClO_5(Sr,ClO_6)$.
$KO,ClO_5(K,ClO_6)$.
$NH_4O,C_2O_3(NH_4,\bar{O})$.
$NiO,SO_3(Ni,SO_4)$.
$BaO,NO_5(BaNO_6)$.

NH_4O,SO_3 (NH_4,SO_4).
$CaO,SO_3(Ca,SO_4)$.
$ZnO,SO_3(Zn,SO_4)$.
$CuO,SO_3(Ca,SO_4)$.
$NaO,SO_2(Na,SO_3)$.
$AgO,CO_2(Ag,CO_3)$.
$HgO,CrO_3(Hg,CrO_4)$.
$KO,NO_5(K,NO_6)$.
$NiO,C_2O_3(Ni,\bar{O})$.
$CdO,SO_3(Cd,SO_4)$.
$KO,SiO_3(K,SiO_4)$.

HgI. ZnS. HgS. FeCl. KCy or KC_2N. $AuCy_3$.
HgCy. Fe_2Cy_3. Hg_2Cy. PbI. AgBr.

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* On the old theory $\bar{O}$ stands for C_2O_3, on the new it stands for C_2O_4, the radical of oxalic acid.

There are three kinds of salts; neutral salts, acid salts, and basic salts. All salts are now regarded as neutral when they contain as many equivalents of acid or rather of the acid radical as there are equivalents of oxygen in the base, as KO,SO_3 (K,SO_4), $Al_2O_3 3SO_3$ $(Al_2,3SO_4.)$ If no prefix is employed to indicate the quantity of acid or base, a neutral salt is always meant; thus, protosulphate of iron FeO,SO_3; sulphate of the sesquioxide of iron, $Fe_2O_3,3SO_3$ $(Fe_2,3SO_4)$. It is thought advisable to retain at present the old nomenclature of salts, and the student can transpose the formulæ for himself.

Exercises.

Write out sulphate of alumina, sulphate of protoxide of mercury, sulphate of teroxide of antimony, chlorate of suboxide of mercury, nitrate of the sesquioxide of iron, sulphate of the sesquioxide of chromium.

An acid salt contains more equivalents of acid than there are equivalents of oxygen in the base. There are two kinds of acid salts, one in which the dry acid is in excess, the other in which it is in excess in the hydrated state, as KO,CO_2,HO,CO_2; the bicarbonate of potash, in which the hydrated acid is in excess; $KO,2CO_2$, the bicarbonate of potash, in which the dry acid is in excess. The hydrated acid salts are now regarded as double salts, in which one base is water. The prefix *bi* or *bis* is employed when the salt contains two equivalents of acid to one of base, as bisulphate of soda, $NaO,2SO_3$. Biborate of magnesia, $MgO,2BO_3$. *Sesqui* is employed to indicate two of base to three of acid, as sesquicarbonate of ammonia, $2NH_4O,3CO_2$. The prefix *ter* is used to indicate three of acid to one of base, as terborate of magnesia, $MgO,3BO_3$.

Basic salts contain an excess of base. The terms *bibasic* and *tribasic* are used to indicate this excess; as bibasic phosphate of lime, $2CaO,PO_5$ (Ca_2,PO_7), tribasic arseniate of soda, $3NaO,AsO_5$ (Na_3,AsO_4). All salts are basic which contain more than one atom of base for one of acid: we have bibasic arseniate of lime, $2CaO,AsO_5$ (Ca_2,AsO_7), bibasic arseniate of soda, $2NaO,AsO_5$, tribasic phosphate of soda and ammonia, NaO,NH_4O,HO,PO_5 (Na,NH_4,H,PO_8). In *polybasic* salts the bases may be, as we have just seen, of different kinds. There are double oxygen salts and double haloid salts, as sulphate of soda and lime, $NaO,SO_3 + CaO,SO_3$.

Compounds of two elements are called binary compounds, as all the oxides and haloid salts, potash, KO, chloride of sodium, NaCl. When two binary compounds unite, a ternary compound is formed, sulphate of potash, KO,SO_3. A quaternary compound is formed when two ternary compounds combine together, as all double oxygen salts, $Al_2O_3,3SO_3 + KO,SO_3$.

The plus sign is used in formulæ to show the arrangement of the elements and their degree of affinity in the compound. The comma is used to separate ternary compounds into binary, and to indicate the force of affinity which binds the elements or compounds together, as KO,NO_5. The potassium and oxygen in this case are united by a greater affinity than the nitric acid and potash, and it would be much easier to separate the salt into nitric acid and potash, than into nitrogen, oxygen, and potassium. When a large figure is placed on the line it multiplies all the elements on the right of it as far as the first comma or plus sign; if before a bracket it multiplies all in that bracket; $2SO_3$ = two equivalents of anhydrous sulphuric acid (sulphuric anhydride); $3CO_2$ = three equivalents of carbonic acid; $3FeO,CO_2$, means three equivalents of protoxide of iron and one of carbonic acid; but $3(FeO,CO_2)$ signifies three equivalents of the protoxide of iron, and three equivalents of carbonic acid (anhydride)—in short, three equivalents of the proto-carbonate of iron. When more than one equivalent of any ternary compound is expressed, it should always be placed in brackets with the figure outside:—

$3AuCl_3$.—Three equivalents of the terchloride of gold.

$2Cr_2Cl_3$.—Two equivalents of the sesquichloride of chromium.

$2(PtCl_2,KCl)$.—Two equivalents of the chloride of platinum and potassium.

$2FeI$.—Two equivalents of the proto-iodide of iron.

$3(NH_4O,NO_5)$ or $3(NH_4,NO_6)$.—Three equivalents of the nitrate of ammonia.

$3(Fe_2O_3,3\bar{O})$ or $3(Fe_2O_3,3C_2O_3)$ or $3(Fe_2,3C_2O_4)$.—Three equivalents of the oxalate of the sesquioxide of iron.

$2MgO,3SiO_3 + 2HO$.—One equivalent of the sesquisilicate of magnesia, and two of water.

$3(Al_2O_3,3SO_3 + KO,SO_3 + 24aq)$.—Three equivalents of the sulphate of alumina and potash and 24 equivalents of non-esential water.

Interpret the following formulæ:—

$3AsH_3$.
BaO,CO_2+CaO,CO_2.
Cu_2I.
NaO,HO.
$3(MgO,CO_2)+aq$.
KO,C_2O_3+HO,or$K\bar{O},+HO$.
$CaO,3SiO_3$.
$2CuO,HO,PO_5$.
$2MgO,NH_4O,PO_5$.
MnO,SO_3+NaO,SO_3.
$Cr_2O_3,3HO$.
$3(2CuO,AsO_3)$.
$2NaO,HO,PO_5$.
CaO,C_2O_3+2HO.
$CaO,CO_2+3(Mg,CO_2)$.
$Al_2O_3,3HO$.

APPARATUS AND MANIPULATION.

Expertness and neatness in the construction and use of apparatus can only be acquired by practice.

Corks are of constant use, and they should be of the best kind. Before fitting them into apparatus they should be softened by rolling them under the foot on a piece of paper, so as to keep them free from dust.

When glass tubes have to be passed through the cork, holes should be made with a cork-borer rather smaller than the tube to be inserted, and nicely eased to the proper size with a rat-tail file. If the cork should prove porous, the exposed parts must be covered with a solution of sealing-wax in spirits of wine.

Glass tubing may be cut into the required lengths by marking round the tube with a three-cornered file; the tube is then easily broken. The depth of the mark made by the file must depend on the thickness of the glass. If the tube has to be bent this should be done before cutting it. To do this the tube must be heated gradually in the flame of a gas or spirit lamp, by turning it round in the flame and moving it a little to the right and left. As soon as the tube becomes red-hot it can be easily bent into the required shape. A little practice with a few pieces of glass tube over a lamp will soon give the experimenter more knowledge than any written instructions.

India-rubber tubing, which can be purchased any size, is used for connecting glass tubes. In connecting glass tubes by means of india-rubber, the ends of the tube should not be brought so close together as to interfere with the flexibility of the arrangement. If the india-rubber tubing does not fit the glass tubes perfectly air-tight, it must be tied with a piece of twine.

A good supply of Florence flasks will be found extremely useful, as they can be fitted up to suit a variety of purposes.

Gas Jars.—Plain cylindrical jars should be from twelve to sixteen inches in length and about two and a half to three inches in diameter, and about one-eighth of an inch in thickness. The open end should be ground, so as to be closed air-tight with a piece of flat glass. A little grease should be placed on the glass plate or edge of the gas jar, so as to insure perfect tightness. Stoppered jars usually have glass stoppers like bottles, and should be ground at the top, so as to be closed with a glass plate in the manner described for cylindrical gas jars. These jars are removed from the pneumatic trough by passing a plate or saucer under the bottom. The plate or saucer should always contain a little water, so as to prevent the access of air and escape of the gas.

When substances have to be burned in a gas jar they are introduced by means of a deflagrating spoon. This spoon should be fitted with a plate, so as to cover the aperture of the jar and allow it to pass the required depth into the gas. In burning phosphorus a piece not larger than a split pea should be used; it should be carefully dried by gently wrapping it in blotting-paper. The phosphorus may then be ignited by touching it with a red-hot wire and quickly introducing it to about two-thirds the depth of the jar.

In collecting gases over water, where heat is required, the heat should in all cases be applied gradually, and the first portions of gas allowed to escape, till all the common air in the apparatus has been expelled.

The *pneumatic trough* should be about two feet long, eighteen inches wide, and a foot deep, provided with a shelf about three inches below the edge of the trough. The water should always stand about one inch over this shelf. Holes are made in this shelf, over which the jars intended to receive the gas are placed. The delivery-tube is brought under the open end of the jar, and as the gas rises in the jar the water is displaced. A spout should be provided to the trough, so as to allow the surplus water to pass into a pail. An ordinary gas jar is filled with air, and, as no two bodies can occupy the same space at the same time, it is necessary to expel this air before the jar can be filled with gas. This is done by immersing the jar sideways in the water of the pneumatic trough, turning it over with the open end downwards, and gently raising it to the shelf of the trough. If the stopper of the jar be air-tight, the water, although above the level of that in the trough, will stand for any length of time in the jar. A little acquaintance with pneumatics will enable the pupils to understand the reason.

It is sometimes necessary to dry gases. This is effected by passing them through a series of Wolfe's bottles containing sulphuric acid, or passing the gas through a tube, from a foot to eighteen inches long and three-quarters of an inch in diameter, filled with the chloride of calcium. Tubes in the form of a U are frequently used. The sulphuric acid and the chloride of calcium are called desiccating agents, because they deprive the gases of water.

In the generation of carbonic oxide, from oxalic acid and hydrated sulphuric acid, carbonic acid, and carbonic oxide, are evolved simultaneously. The presence of the carbonic acid interferes with a careful examination of the carbonic oxide. This carbonic acid must be removed; and this is effected by passing the gases through a strong solution of caustic potash or lime water, which retains the carbonic acid. By this method pure carbonic oxide is obtained, which is passed into gas jars filled with water at the pneumatic trough.

When no heat is required to liberate gases or gaseous compounds, a two-necked Wolfe's bottle may be used. It should be fitted up with two glass tubes. The delivery-tube, from which the gas escapes, may be about a quarter of an inch in diameter, and pass just through the cork of the bottle; another tube, passing to within a quarter of an inch of the bottom of the bottle, and about half an inch in diameter, is used for pouring in acid or water. This saves the trouble of taking out the cork, which must be done every time fresh acid or water is added, unless the bottle be provided with two tubes in the manner described. The materials in the bottle should never occupy more than a third, and the acid should be added carefully, a little at a time, and the materials in the bottle agitated by a little shaking. A Florence flask may be fitted up with two tubes; but the neck is rather too small, unless great care be taken in boring the corks.

The whole of the apparatus for the evolution of gases must fit air-tight. Too much pains cannot be taken to secure this object; it will save much trouble and disappointment. The expression, "That's near enough," will not do in chemistry; it must be exactly, or a satisfactory result cannot be obtained.

Gases are sometimes collected by *displacement*. If the gases are lighter than common air, as hydrogen and ammonia, or heavier, as chlorine and carbonic acid, the pneumatic trough may be dispensed with. But gases of nearly the same specific gravity of common air must be collected over water. The same generating apparatus as that just described will do, provided the flask or bottle will stand heat if required. The delivery-tube must pass to the bottom of the jar it is intended to fill with gas. This is often done by blocking the jar up with pieces of wood of different thicknesses. The top of the jar may be covered with a piece of card-board, with a hole for the delivery-tube to pass through. As the gas collects at the bottom it gradually expels all the air, exactly the same as if water were poured into the gas jar. See fig. 1. In raising bottles or other apparatus let the blocking be firm, the largest and heaviest pieces at bottom. When the gases are lighter than common air the receiver may be supported by a retort-stand, the delivery-tube passing upwards to the extremity of the vessel intended to receive the gas. When large quantities of gas are required, a gasometer is used. The most convenient is the one known

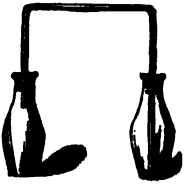


Fig. 1.

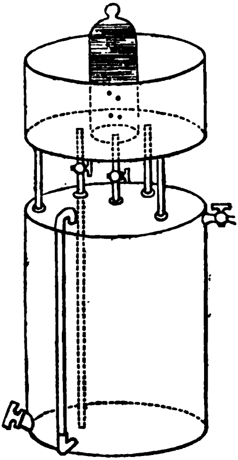


Fig. 2.

as Pepys' gas-holder. See fig. 2. I shall not describe the action of this piece of apparatus, because five minutes' examination of one of these gasholders will give the student a much better idea of its action than any written description. The most careful explanation of any piece of mechanism or apparatus must fall short of actual observation; and the teacher who attempts to teach without the objects, when they can be obtained, teaches words and not things.

In chemistry the terms *solution*, *soluble*, and *solvent* are frequently employed. Many solid bodies, when placed in water, dissolve. Salt and sugar dissolve in water. These substances are said to be soluble in water; and the water is called the solvent or menstruum. The more a solid is reduced to powder the easier it dissolves. If sugar be placed in alcohol, it does not dissolve. Here, then, we have an example, which may be multiplied to a great extent, of a body soluble in one liquid and insoluble in another. A solution may be either mechanical or chemical. The sugar in water is a case of mere mechanical solution; no chemical change has taken place, and, by evaporation, the sugar can be obtained in its original condition. Salt dissolved in water is another illustration of the same kind. In chemical solutions the solid and liquid combine together, forming a new compound, from which the original solid and fluid cannot be obtained by evaporation or any mechanical operation. When chalk is dissolved in hydrochloric acid, or copper dissolved with nitric acid, we have examples of chemical solution. A liquid can only dissolve a certain quantity of a solid substance. This quantity will vary according to the solid and liquid employed. When the solvent or liquid contains as much of the solid matter as it is capable of dissolving, it is said to be saturated, as all additional quantities of the solid fall to the bottom.

The solubility of most substances is increased by an increase of temperature.

100 pts. of water at	60°F.	dissolve	11 pts.	of sulphate of potash.	
" "	212°	"	25	"	"
" "	60°	"	32 pts.	of dry sulphate of magnesia.	
" "	212°	"	74	"	"

Common salt is about as soluble in hot water as cold. Some bodies attain their maximum solubility at a particular temperature. The solubility of the dry sulphate of soda is at its maximum at a temperature of 93°; but in the nitrate of soda we have the order reversed,—water, at 32°, dissolves 80 parts; at 50°, only 23 parts; but from this point the solubility increases up to 200 parts at a temperature of 212° F.

Substances dissolved in alcohol, ether, oils, and water are called simple solutions. The most important solvent is water; and when the substance is soluble in this liquid no other is used.

Elevation of temperature accelerates all chemical solutions: heat promotes the action of bodies on each other. The quantity of a dissolved body in a chemical solution is always the same. When bodies enter into chemical combination, they unite in definite and fixed proportions irrespective of temperature. Solutions are made either in test tubes, Florence flasks, evaporating dishes, or beakers, the size being regulated by the quantity of the substance to be dissolved.

Filters and Filtration.—Filtering-paper should be of uniform texture, and sufficiently porous to admit a free passage of the fluid. Filtration is purely a mechanical operation. The object is to separate the solid from the liquid, the former collecting on the filter, and the latter passing through, which is called the filtrate. The filtering-paper is supported by a glass funnel, and the

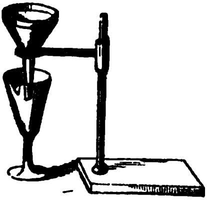


Fig. 3.

funnel may be supported by one of the rings of the retort stand. See fig. 3. A filter is easily made by folding the paper in such a manner as to bring the four corners together, and, by opening one corner from the other three, a conical cavity is formed. The filter may now be placed in the funnel, taking care that it is the proper size. It should fit nicely against the sides of the funnel, and the substance to be filtered added by gently pouring it on the side. If projected violently into the centre of the filter, the paper is liable to be broken. Nothing short of actually doing the thing will give the proper knowledge. It would be just as reasonable to expect a man to learn the trade of a joiner by reading books on the construction of tools, as to expect a person to learn chemistry without proper apparatus; but this apparatus need not be expensive, if the student have (as he ought to have) some inventive and constructive power.

NON-METALLIC BODIES.

The metalloids or non-metallic bodies may be divided into five classes, each division being characterized by a great resemblance in their chemical properties :—

First Division.

Oxygen,
Sulphur,
Selenium.

Second Division.

Chlorine,
Bromine,
Iodine,
Fluorine.

Third Division.

Nitrogen,
Phosphorus.

Fourth Division.

Carbon,
Silicon,
Boron.

Fifth Division.

Hydrogen.

FIRST DIVISION.

OXYGEN.

Symbol O ; *Equivalent*, 8.

The non-metallic elements, or metalloids, constitute a large proportion of the material world. Oxygen gas was discovered by Priestley, in 1774, who called it dephlogisticated air. It was also discovered about the same time by Scheele, a Swedish chemist. The term oxygen was given to this gas by Lavoisier. This element is largely distributed throughout nature, both in the organic and inorganic kingdom. It constitutes eight parts by weight in every nine of water, which covers three-fourths of the globe. One-third by weight of the solid crust of the globe is oxygen, as will be seen when we consider the vast quantities of silica, lime, and alumina of which it is composed.

The atmosphere is composed by weight of 23 parts of oxygen, and 77 parts of nitrogen. By volume it contains 79 parts of nitrogen, and 21 parts of oxygen.

Preparation.—Oxygen may be prepared in a variety of ways. 1. By exposing to heat in a small glass retort the red oxide of mercury. In this substance the affinity which holds the oxygen to the mercury is so feeble that exposure to heat is sufficient to cause a separation into oxygen gas and metallic mercury. The mercury collects in small globules in the cold part of the retort, and the oxygen passes over as a colourless, inodorous, invisible gas: $HgO = Hg$ and O free. The chief interest connected with this method is that it was the one adopted by Priestley: it is seldom adopted by chemists, as being too expensive. 2. It may also be prepared by exposing to a red heat, in an iron bottle or retort, the binoxide, or peroxide as it is sometimes called, of manganese, MnO_2. The oxide of manganese yields up half an equivalent of oxygen, reducing it to the sesquioxide, $MnO_{1\frac{1}{2}}$, which doubled makes Mn_2O_3. This method is sometimes adopted when large quantities of oxygen are required, but the oxygen is seldom pure. See fig. 4. 3. Oxygen may also be prepared by distilling in a glass retort the binoxide of manganese and sulphuric acid. These materials should be well mixed, so that no dry place appears in that part of the retort where

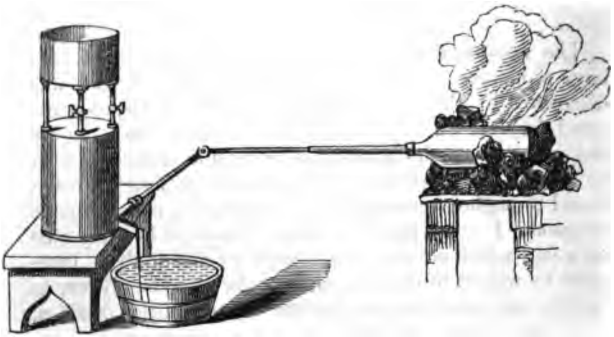


Fig. 4.

the materials are placed. If this is not attended to, fracture of the retort is likely to take place. Sulphuric acid will not unite with the binoxide of manganese, but it unites with the protoxide readily. The decomposition is easily illustrated : $MnO_2 + SO_3 = MnO,SO_3 + O$ free. 4. Oxygen is also prepared by the decomposition of water by voltaic electricity. See fig. 5. This is interesting as determining the composition of water, the hydrogen being liberated at the negative electrode, and the oxygen at the positive, in the proportion of double the volume of hydrogen to that of oxygen, which is the proportion in which these gases combine by volume to form water. 5. The usual method of preparing oxygen is by heating a mixture of equal parts of the binoxide of manganese and the chlorate of potash, which should be well dried in a capsule. See fig. 6. The manganese in this case undergoes no change, but the gas is given off at a much lower temperature than when the

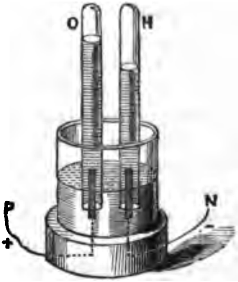


Fig. 5.

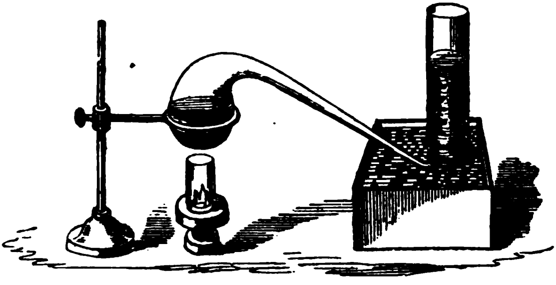


Fig. 6.

chlorate is used alone. In the preparation of gases the first part is a mixture of air, and should always be allowed to escape. The chlorate of potash is composed of 1 equivalent of potassium, 1 of chlorine, and 6 of oxygen; on the application of heat the whole of the 6 equivalents of oxygen is given off. Decomposition, $KO,ClO_5 = KCl + 6O$, or $KClO_6 = KCl + 6O$.

Oxygen at all observed temperatures is a gas. It is not combustible; that is, it will not burn, but it is a powerful supporter of the burning of other bodies. It has no acid or alkaline reaction, although it enters into the composition of all oxygen acids and alkalies. Bodies which burn in air, burn with increased brilliancy in pure oxygen. A strip of wood, if only a spark remain, is instantly rekindled if plunged into a jar of this gas. Small iron wire, twisted into a spiral like a cork-screw, and the end dipped in sulphur, burns with beautiful little sparks. See fig. 7. Phosphorus burns with a white light of great brilliancy; sulphur, with a beautiful blue light; and charcoal burns with beautiful scintillations. In all these cases of combustion, the products are oxides of the burning bodies. Some of these are acid oxides or *acids*. When iron wire is burned, the product is the black oxide of iron, Fe_3O_4; the burning phosphorus produces phosphoric acid (anhydrous) in solid white flakes, which dissolve in the water: sulphurous acid gas, SO_2, is produced from the sulphur, and carbonic acid gas, CO_2, from the burning charcoal. Both the SO_2 and the CO_2 dissolve in the water. In all these cases chemical combination takes place between the oxygen and the burning body; and the weight of these bodies is increased by the exact weight of the oxygen consumed. All the common cases of combustion which take place in air

Fig. 7.

are precisely the same as those which take place in oxygen. The action in pure oxygen is very energetic, owing to the absence of nitrogen. In the atmosphere these changes take place in a more subdued manner. The respiration of animals is, in effect, the same as common combustion. The substances in the blood burn away slowly by the aid of the oxygen introduced into the system. The activity of oxygen never ceases, although the phenomena of combination is not so marked as those just described. An ounce of iron rusting in the air, produces just the same amount of heat as if it had been burnt like the wire in oxygen. This slow combustion; unattended by light, extends over several years, while the same result could be produced by pure oxygen (assisted at first by heat) in a few seconds.

The decay of animal and vegetable substances is a process of slow oxidation. The offensive impurities of the air and water are destroyed in a great degree by a slow process of burning. The lungs of land animals, and the gills of fish, are adapted to absorb oxygen, the animal from the air, and the fish from the dissolved oxygen of the water. An animal confined in a given bulk of air dies as soon as the oxygen is consumed; but if confined in the same bulk of pure oxygen, the circulation quickens, and fever and death are the results, not, however, as in the previous case, from deprivation of oxygen, but from an excess of that body, causing the animal to live *too quickly*. Such is a brief outline of the relation of oxygen to animal life; its action seems destructive, yet it is the sustainer of life, the mainspring of all vital action. Oxygen may be distinguished from every other gas, except nitrous oxide, by plunging into it a taper which has just been blown out; if the smallest part is incandescent, it immediately bursts into flame.

SULPHUR. BRIMSTONE.

Symbol, S; *Equivalent*, 16.

At ordinary temperatures this element is a pale yellow crystalline solid. It is often found in a free state. Most of the sulphur employed in this country comes from Sicily and other volcanic districts. In combination with iron, antimony, zinc, lead, copper, and other metals, as sulphides, it occurs abundantly in nature; in union with oxygen as sulphuric acid, it is found combined with lime, baryta, strontia, &c., as sulphates; it is present in small quantities in all animal and vegetable substances. It exists in the albumen of eggs. The blackening of silver spoons, used with eggs, is owing to the formation of the sulphide of silver.

Sulphur melts at 230°, becoming a thin amber-coloured liquid; when heated to 430° it thickens like treacle. In this state it may be poured from a flask into cold water, when it retains for some time a soft, flexible condition, like gutta percha; this is the amorphous condition of sulphur. If it be raised to a temperature of 790° it boils; and if the vapour be conducted into a cold chamber it condenses in crystals known as sublimed sulphur, or *flowers of sulphur*. The roll sulphur of commerce is formed by running it when in a melted state into wooden moulds.

Sulphur in many of its chemical relations bears a close analogy to oxygen. Many of the oxides of the metals have corresponding sulphides; these sulphides often unite, forming compounds analogous in composition to salts. There are probably seven oxygen compounds of sulphur—

Sulphurous acid	SO_2
Sulphuric acid	SO_3
Hyposulphurous acid	S_2O_2
Hyposulphuric acid	S_2O_5
Sulphuretted hyposulphuric acid	S_3O_5
Bisulphuretted hyposulphuric acid	S_4O_5
Trisulphuretted hyposulphuric acid	S_5O_5

The two first only of these compounds will engage our attention.

Sulphurous acid, SO_2.—When sulphur is burned in oxygen, sulphurous acid, SO_2, is formed. It may also be prepared by heating equal weights of sulphuric acid and mercury or copper turnings. The sulphurous acid comes off as a colourless gas. It must be collected by displacement as its great solubility in water will not admit of its collection at the pneumatic trough. The decomposition may be represented as follows:—

$$2(HO,SO_3) \text{ and } Hg \text{ (or } Cu) = HgO,SO_3 \text{ (or } CuO,SO_3) + SO_2 + 2HO; \text{ or,}$$

$$2(H,SO_4) + Hg = Hg,SO_4 + 2HO + SO_2.$$

Sulphurous acid has bleaching properties, and is used for bleaching straw used in plaiting, and woollen goods. The plaiters in Bedfordshire suspend their plait in a tea chest, and set fire to some sulphur in an iron dish at the bottom; the top of the chest is then covered over with a sack. If blue violets, or any coloured flowers, be suspended in a jar of this gas, the colour is entirely removed. Litmus paper, dipped in a solution of sulphurous acid, is first reddened, then bleached. It has also the power of arresting animal decomposition.

Sulphuric acid, HO, SO_3.—This is a most important compound. It is extensively used in the manufacture of soda from common salt, also in the manufacture of chlorine, for bleaching, for dissolving indigo, dyeing, calico printing, and numerous other purposes. The hydrated acid, HO,SO_3 (H,SO_4), has been known from the fourteenth century. It is now prepared by two methods:—1, from the distillation of the sulphate of iron, FeO, $SO_3 + 7HO$, (green vitriol); and 2, by the oxidization of sulphurous acid with the vapour of hyponitric acid, NO_4. The first method is that adopted on the Continent. The sulphate of iron is well dried, and placed in a stoneware retort, which is raised to a red heat. The sulphuric acid comes over in combination with water. This acid, when strongest, contains two equivalents of anhydrous acid (anhydride) to one of water, $HO,2SO_3$. It is much stronger than the English acid, and fumes in the air. It is called *Nordhausen sulphuric acid.*

Anhydrous sulphuric acid (SO_3) is a solid, and may be obtained by gently heating the strongest (Nordhausen) acid in a retort provided with a receiver, surrounded by a freezing mixture. The vapour, on reaching the receiver, is condensed into white silky crystals. It has no acid properties; but, when thrown into water, it hisses like red-hot iron, and becomes converted into the ordinary acid. On exposure to the air, it liquefies by the absorption of moisture.

The method usually adopted in the manufacture of sulphuric acid depends on the formation of sulphurous acid, and its subsequent oxidation into sulphuric acid. Sulphur is kindled in a furnace freely supplied with air; the flame of the sulphur heats a crucible containing a mixture of nitre and sulphuric acid, which evolves the vapour of nitric acid, (HO,NO_5); this is conducted into a large leaden chamber accompanied with sulphurous acid gas formed from the burning sulphur, and a large quantity of unaltered atmospheric air. (See fig. 8.) The floor of the chamber is covered with water, and a jet of steam plays into it. At the end of the chamber an exit tube is provided for the spent gases. By this arrangement a constant supply of sulphurous acid, nitric acid, water, and atmospheric air is provided.

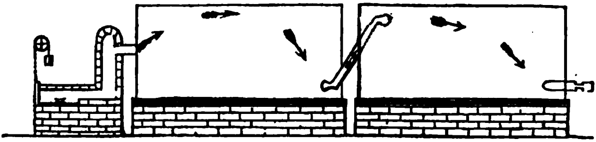


Fig. 8.

The following changes take place in the leaden cistern:—At first a portion of the vapour of nitric acid gives up part of its oxygen to a portion of sulphurous acid, converting it into sulphuric acid, and becoming itself reduced to binoxide of nitrogen. This change may be thus represented:—

$$HO,NO_5+3SO_2+3HO=3(HO,SO)_3+NO_2+HO.$$

This NO_2 in contact with the current of atmospheric air continually passing through the chamber, takes two equivalents of oxygen, and becomes converted into hyponitric acid, NO_4 (not again to NO_5). The NO_4 in turn parts with two equivalents of oxygen to a fresh portion of SO_2, oxidizing it into sulphuric acid, and becoming again reduced to NO_2.

$$NO_4+2SO_2+2HO=2(HO,SO_3)+NO_2.$$

A small quantity of NO_4 is thus sufficient to convert a large quantity of sulphurous acid into sulphuric, the presence of water being in all cases necessary to the change. The NO_4 plays the part of carrier between the oxygen of the air, and the sulphurous acid. At intervals the acid is drawn off from the floor of the leaden chamber, and is deprived of any excess of water by evaporating in shallow leaden pans.

Sulphuric acid has a great affinity for water. Gases are frequently dried by being transmitted through this acid. The nitrate of baryta, or the chloride of barium, is a test for sulphuric acid, or any soluble sulphate. A white precipitate is formed, which is insoluble in nitric acid.

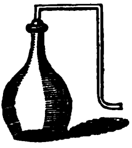


Fig. 9.

Sulphuretted hydrogen, or *hydrosulphuric acid*, HS.—This compound of sulphur and hydrogen is very useful in the laboratory. It may be prepared by the action of dilute sulphuric acid on the protosulphide of iron. The apparatus for hydrogen may be used. (See fig. 9.) The change may be expressed thus:—

$$FeS + HO,SO_3 = HS + FeO,SO_3;$$

$$\text{or, } (FeS + H,SO_4 = HS + Fe,SO_4).$$

Sulphuretted hydrogen, at the common temperature, is a gas very soluble in water, having a most disagreeable odour, and very deleterious to animal life. If a jar of this gas is kindled, it burns with a blue flame, depositing sulphur. When mixed with chlorine, the whole of the sulphur is liberated. Cold water dissolves its own volume of this gas, and the solution is kept as a test. For this purpose the apparatus shown in fig. 10 may be used. It soon spoils by absorption of oxygen, and is best prepared as wanted. Jars of this gas may be collected over warm water; but the pneumatic trough, or any painted vessel, must not be used, as it will be blackened by the formation of a metallic sulphide. This gas is a frequent natural product. It occurs in the Harrogate waters, and is a constant product of the decay of animal and vegetable matter. It has the properties of an acid; reddens litmus, and is called *hydrosulphuric acid*, which must never be confounded with *hydrated sulphuric acid*, HO,SO_3. A piece of white blotting paper, dipped in a solution of acetate of lead, is immediately blackened by the smallest trace of this gas. This method is frequently adopted to test the purity of coal gas. When it occurs in apartments or drawers, it may be entirely destroyed by chlorine. Silversmiths usually burn their gas outside, owing to the sulphuretted hydrogen always present in small quantities in coal gas, blackening the silver by the union of the sulphur with that metal.

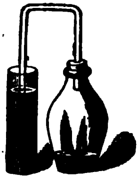


Fig. 10.

SELENIUM.

Symbol, Se; *Combining Equivalent*, 39·5.

This is a rare substance, analogous to sulphur, and forming seleniates, which exactly resemble the sulphates. It forms three compounds with oxygen:—

Oxide of selenium, SeO.—A gas; has the smell of horse radish; not acid; and slighty soluble in water.

Selenious acid, SeO_2.—This is analogous to sulphurous acid.

Selenic acid, HO,SeO_3.—The hydrated acid resembles sulphuric acid.

SECOND DIVISION.

CHLORINE.

Symbol, Cl; *Combining Equivalent*, 35·5.

This element is very abundantly diffused throughout nature. In combination with the metal sodium, it forms common salt. It is never found in its free or uncombined state. At ordinary temperatures, it is a gas of a greenish-yellow colour; and in very small quantities has an odour like sea water; but in large quantities, it produces violent coughing and irritation. It is prepared by mixing, to the consistency of cream, strong hydrochloric acid, and the black oxide of manganese. Care should be taken that every part is well mixed, so that no dry place appears inside the retort or flask, or fracture is most likely to result. A gentle heat should now be applied, when the gas comes over in abundance. It should be collected over warm water, as cold water absorbs about twice its volume; or it may be collected by displacement, like carbonic acid. The reaction is easily understood:—

$$2HCl + MnO_2 = 2HO + MnCl + Cl.$$

Chlorine has little affinity for oxygen; it combines readily however with hydrogen and the metals. If a lighted taper be immersed in a jar of the gas, it burns with a dull red light, with an abundant evolution of carbon, in the form of smoke. If a piece of blotting-paper be soaked in turpentine, and placed in a jar of chlorine it bursts into flame. In both these cases, the chlorine combines with the hydrogen to form hydrochloric acid, to the exclusion of the carbon, which falls as soot—a very good illustration of the superior affinities of chlorine for hydrogen.

Phosphorus combines with chlorine spontaneously, but the flame is very feeble. Copper-leaf, powdered antimony, and arsenic, all undergo spontaneous combustion in chlorine. If equal volumes of hydrogen and chlorine be introduced into a suitable apparatus, and an electric spark sent through the mixture, combination takes place with great violence. The direct rays of the sun produce a similar union. The most remarkable property of this gas is its power of destroying organic colouring matters. A solution of indigo, which is not affected by strong sulphuric acid, is changed into a brown liquid by chlorine, to which the original colour cannot be restored. In a perfectly dry state, chlorine has no bleaching properties. When in solution, it may be poured into test-glasses, containing infusions of litmus, purple cabbage, turmeric, carmine, cochineal, &c.; the colour of these substances will be removed. When bleaching is practised on a large scale, a combination with lime is used called *bleaching powder*. The outline of this process can be easily imitated. The bleaching powder is mixed with warm water, and the solution filtered. Any piece of coloured calico may be taken, providing its tint is due to some organic colouring matter. It may then be washed in the filtered solution. The liquid bleaches but slightly; but if it be removed from the bleaching bath, and rinsed in a bath of dilute sulphuric acid, and this process carried on a few times alternately, the whole of the colour is removed. The chlorine, when in combination with the lime, is comparatively inactive; but when dipped in dilute acid the lime combines with the acid, and the chlorine in the presence of water is set free, and manifests its bleaching properties. An experiment illustrating this property may be performed in the following manner. A piece of cotton, dyed with madder, or any organic colouring matter, may have any pattern or device stamped upon it, or painted with a mixture of gum and tartaric acid. The cloth is dried, and afterwards rinsed, in a hot solution of bleaching powder. A white pattern on a coloured ground is developed in consequence of the acid mixed with the gum liberating the chlorine, and destroying the colour in those parts upon which the pattern was printed. The other parts of the cloth are not sensibly affected.

Chlorine is used as a *disinfectant* and *deodorizer*, that is, it destroys infectious and unpleasant and offensive odours. The chlorine is usually evolved from chloride of lime, which is dissolved in hot water, and a little vinegar or dilute sulphuric acid added to liberate the gas.

Hydrochloric acid, HCl.—This is a most important acid. Hydrochloric acid gas is prepared by heating in a properly-arranged flask a mixture of common salt, a little water, and sulphuric acid. It must be collected by displacement, or over mercury. It is colourless, fumes when in contact with the atmosphere, and is very soluble in water; one cubic foot of that liquid dissolving about 480 cubic feet of gas. A solution of this gas is the acid usually known as muriatic or hydrochloric acid. It may be prepared by mixing equal weights of dry common salt and sulphuric acid in a flask, and applying heat; the gas after being purified by passing it through a wash-bottle, is conducted from thence into distilled water, by which the gas is absorbed. The reaction is easily understood:—

$$NaCl + HO,SO_3 = NaO,SO_3 + HCl;$$
$$\text{or, } (NaCl + H,SO_4 = Na,SO_4 + HCl).$$

A mixture of nitric and hydrochloric acid is called *aqua regia*, from its property of dissolving gold.

The presence of any soluble chloride, or hydrochloric acid, may be detected by adding a solution of the nitrate of silver, when a white curdy precipitate is produced, insoluble in nitric acid:—

$$NaCl + AgO,NO_5 = Na,O,NO_5(\text{soluble}) + AgCl\ (\text{insoluble}),$$
$$\text{or, } (NaCl + Ag,NO_6 = Na,NO_6 + AgCl).$$

There are five compounds of oxygen with chlorine:—

Hypochlorous acid	ClO
Chlorous acid	ClO_3
Hypochloric acid	ClO_4
Chloric acid	ClO_5
Perchloric acid	ClO_7

ClO_3, ClO_4, and ClO_7 are obtained from chloric acid. The other compounds are obtained by the action of chlorine on metallic oxides.

Hypochlorous acid, ClO, is a yellow gas, has very powerful bleaching properties, and may be prepared by adding the powdered red oxide of mercury to a bottle of chlorine containing a little water. The bottle should be briskly shaken for a few moments. Part of the oxygen of the HgO combines with part of the Cl to form ClO, which dissolves in the water, while the remaining Cl forms, with the Hg and HgO, an insoluble compound, called the oxychloride of mercury; thus—

$$2HgO + 2Cl = HgCl,HgO + ClO.$$

Chlorous acid, ClO_3.—This compound is very difficult of preparation, and not very important.

Hypochloric acid, ClO_4, sometimes called the peroxide of chlorine.—This acid may be prepared by carefully mixing concentrated sulphuric acid and the chlorate of potash into a thick paste. After it has cooled it should be cautiously introduced into a green glass retort, and very slowly heated by placing the retort in water, and gently raising the temperature of the water by a lamp or other source of heat; a green-yellow gas comes off, which is hypochloric acid, ClO_4. It must be collected by displacement, as it is very soluble in water, and mercury decomposes it.

This compound is very explosive. At a temperature below boiling water it is resolved with great violence into oxygen and chlorine; hence the care required in its preparation. Decomposition:—

$$3(KO,ClO_5)+2(HO,SO_3)=2ClO_4+KO,ClO_7+2(KO,SO_3)+2HO.$$

If a mixture of chlorate of potash and loaf sugar, both in powder, be placed on a plate and touched with a glass rod previously dipped in sulphuric acid, it is instantly set on fire. In this experiment the hypochloric acid, ClO_4, is disengaged with such violence as to cause combustion.

If half a teaspoonful of chlorate of potash and a few chips of phosphorus be placed in a deep test or ale-glass nearly filled with water, and a little sulphuric acid added by a small funnel which reaches to the bottom of the glass, the phosphorus will burn under water from the oxygen of the disengaged hypochloric acid.

Chloric acid, ClO_5, forms, in combination with bases, salts of considerable importance, especially the chlorate of potash. If a current of chlorine be sent through a moderately warm solution of caustic potash, and the liquid, when saturated with the gas, concentrated by evaporation, crystals of chlorate of potash are obtained. In this reaction a part of the potash, KO, is decomposed, its oxygen uniting with part of the chlorine to form chloric acid, and the potassium forming, with another portion of chlorine, chloride of potassium.

$$6KO+6Cl=KO,ClO_5+5KCl;$$
$$\text{or, } (6KO+6Cl=K,ClO_6+5KCl).$$

From the chlorate of potash chloric acid is obtained, by boiling with a solution of hydrofluosilicic acid. The process is tedious, and requires care.

Perchloric acid, ClO_7, is obtained from the perchlorate of potash, KO,ClO_7.

The chlorates in solution give no precipitate with nitrate of silver. When heated they evolve oxygen; and if treated with sulphuric acid they evolve the characteristic yellow gas as mentioned under hypochloric acid.

BROMINE.

Symbol, Br; *Equivalent*, 78·26.

This element is found in sea-water. At ordinary temperatures it is a brownish-red liquid, having a disagreeable odour, closely resembling chlorine and iodine in character. It is found in combination with magnesium in sea-water, as the bromide of magnesium, MgBr. It is obtained by agitating any solution of bromine with ether. After the less soluble salts have been separated, a stream of chlorine is sent through the solution, which decomposes the bromide of magnesium, and ether being added, dissolves the liberated bromine. The liquid is transferred to a retort with a receiver, and heat applied, when the bromine condenses in the cold part of the retort and receiver. Bromine, when associated with water has bleaching properties.

Hydrobromic acid, HBr, is analogous to hydriodic acid, has nearly the same properties, and may be prepared by substituting bromine for iodine, in the preparation of hydriodic acid.

Till lately there was only one known oxide of bromine, *bromic acid*, BrO_5, which is prepared by adding bromine to a solution of caustic potash, until all alkaline properties are destroyed. The change is the same as that described under chloric acid.

A few months ago a compound resembling ClO was discovered, and on analysis found to have the formula BrO—hypobromous acid.

IODINE.

Symbol, I; *Combining Equivalent*, 127.

At ordinary temperatures, iodine is a bluish-black metallic-looking substance, having the appearance of scales of black lead. It occurs chiefly in combination with sodium and potassium in sea-water. It is a constituent of marine plants and animals. It is chiefly prepared from the ashes of sea-weeds, called kelp. The ashes are mixed with water, filtered and evaporated, until the solution is reduced to a very small quantity: the salts of sodium, potassium, and magnesium being removed as they crystallize. The mother-liquor left after the removal of these salts, contains the iodine in combination with magnesium and sodium. This is placed in a leaden retort with sulphuric acid and the binoxide of manganese. On the application of heat the iodine distils over as a dark purple vapour, which is condensed. This process may be imitated by placing in a retort connected with a receiver, a solution of the iodide of potassium, a little powdered binoxide of manganese, and sulphuric acid: on the application of heat, a beautiful violet vapour passes over, which condenses in the cold part of the retort and receiver in small shining scales. Decomposition:—

$$KI+MnO_2 + 2(HO,SO_3) = KO,SO_3 + MnO,SO_3 + 2HO + I;$$

$$\text{or, } (KI + MnO_2 + 2(H,SO_4) = K,SO_4 + Mn,SO_4 + 2HO+I).$$

Iodine stains the skin; it acts violently on the animal system, and is used in medicine.

When iodine is free, it produces a splendid blue colour in contact with starch. If any soluble iodide be first liberated with a little chlorine water, it is then capable of producing the characteristic colour with starch.

Iodine forms with oxygen two acids, *iodic acid*, IO_5, and *periodic acid*, IO_7.

In combination with hydrogen it forms an acid known as *hydriodic acid*, HI. This acid is a gas, but it may be obtained in solution by passing sulphuretted hydrogen through water in which powdered iodine is suspended. In this case, the sulphur of the sulphuretted hydrogen is liberated, and the hydrogen combines with the iodine, forming hydriodic acid. $HS + I = HI + S$. The liquid is warmed, filtered, and a colourless solution of the acid is obtained, which soon becomes of a brown colour on exposure to air, by its decomposition into free iodine and water, $HI + O = HO + I$.

Iodine dissolves freely in alcohol. If a polished silver plate be held over the vapour of iodine, it first becomes yellow, then violet, and then a deep blue, owing to a combination of the iodine with the silver. The iodide of silver thus formed is decomposed by light.

FLUORINE.

Symbol, F; *Equivalent*, 18·9.

Owing to the powerful affinities of this element for the metals, also for silicon and hydrogen, it has never been isolated. Its most frequent combination is with calcium, as the fluoride of calcium, CaF, or fluor-spar. It exists, also, in a state of combination, in the bones and teeth of animals. So far as it has been examined, it is said to be a gas, like chlorine.

There is no known oxygen compound of fluorine; but in combination with hydrogen, it forms an acid known as *hydrofluoric acid*, HF. To obtain this compound, the powdered fluoride of calcium is heated with concentrated sulphuric acid in a platinum or lead retort connected with a receiver of the same metal, kept cool by being surrounded with a freezing mixture. A volatile, colourless liquid is obtained, which emits suffocating white fumes in the air. It unites with water, and very dilute solutions attack glass. It has, therefore, to be kept in leaden, or gutta-percha bottles.

The decomposition is easily explained:—

$$CaF + HO,SO_3 = HF + CaO,SO_3;$$

$$\text{or, } (CaF + H_2SO_4 = HF + Ca,SO_4).$$

If the concentrated acid be dropped upon the skin, it occasions ulcers of a very malignant and dangerous kind. Hydrofluoric acid is used to etch or engrave letters on glass. The glass to be operated upon is first covered with bees'-wax, and the lines to be etched scratched through with a needle. The glass is then exposed to the action of the vapour of hydrofluoric acid, or the acid is poured upon it in a liquid state. Where the glass is unprotected with wax, it is corroded. It is in this manner that thermometers, and other glass vessels, are graduated. The power of the acid to etch glass may be observed by having a leaden saucer, into which some powdered fluor-spar and sulphuric acid are placed. The glass to be etched is first covered with a coating of wax, and the device drawn upon it with a penknife so as to remove the wax from those parts to be etched. The glass which should cover the leaden saucer is now placed on the top, with the waxed side downwards. The saucer should now be gently heated, care being taken that the temperature does not rise sufficiently to melt the wax. In a few minutes the glass may be removed and cleaned with a little hot turpentine, when it will be found to be etched.

THIRD DIVISION.

NITROGEN.

Symbol N; *Combining Equivalent,* 14.

This gas is without taste, odour, or colour, rather lighter than common air, and distinguished from all the other gases by its negative properties. It forms by volume $\frac{4}{5}$ of the atmosphere, and enters into a variety of combinations. Nitrogen exists in the vegetable kingdom. It is probably derived from the ammonia of the atmosphere. It is a constituent of muscular tissue to the extent of about 17 per cent. It is not found in the mineral crust of the earth, except in some kinds of coal.

Fig. 11.

Nitrogen may be prepared, 1. By burning a small piece of phosphorus in a confined portion of common air under a gas jar standing over water, by which phosphoric acid is produced, and the nitrogen liberated. The white solid flakes of phosphoric acid are dissolved in the water over which the jar is placed. (See fig. 11). It must stand till the jar has become clear. Theory of change:—

Air {Oxygen
Nitrogen — —Nitrogen.
Phosphorus——Phosphoric acid.

2. By passing chlorine through a solution of ammonia:

$$3Cl + NH_3 = 3HCl + N \text{ free.}$$

3. By mixing iron filings and sulphur with water, and allowing the mixture to remain under a gas jar for twenty-four hours. The oxygen of the atmosphere is absorbed and the nitrogen set free. 4. By treating nitric acid with lean butcher's meat in a retort. This change is not well understood.

The atmosphere, which is the great storehouse of nitrogen, is composed of a mixture of nitrogen and oxygen, with variable proportions of carbonic acid, aqueous vapour, and traces of ammonia.

The analysis of common air may be effected in the following manner:—Introduce, by means of a wire, into a graduated tube standing over water, a piece of phosphorus. The known quantity of air can be determined by observing the water point in the graduated tube. The arrangement should now be left for twenty-four hours. The whole of the oxygen will be absorbed by slow combination with the phosphorus. The remaining gas may now be read off, which is nitrogen. Another plan is to mix air with hydrogen, and explode by an electric spark; the hydrogen and oxygen combine, and the remaining volume of gas is read off.

Nitrogen is not combustible, nor is it a supporter of combustion or respiration; it is destitute of any acid or alkaline reaction. It is estimated that six tons of air pass through a blast furnace every hour. While the oxygen is most active in forming new compounds the nitrogen escapes unchanged. Its chief function is to modify the violent action of the oxygen, and render the atmosphere suitable to the wants of life.

There are five compounds of oxygen and nitrogen, which may be taken as a good illustration of the law of multiple proportions in chemistry:—

Protoxide of nitrogen	NO
Binoxide of nitrogen	NO_2
Nitrous acid	NO_3
Hyponitric acid	NO_4
Nitric acid	NO_5

The *protoxide of nitrogen* NO, is sometimes called laughing-gas, from the peculiar exhilarating effects produced when inspired. It is prepared by heating moderately in a retort the nitrate of ammonia. Four ounces of salt produce one cubic foot of gas. It should be collected over warm water, and care should be taken that the salt is pure. The change which takes place is illustrated by the following equation:—

$$NH_4O,NO_5 = 4HO + 2NO.$$
$$(NH_4NO_6 = 4HO + 2NO).$$

Nitrous oxide is a supporter of combustion. The inhalation of this gas is sometimes attended with serious results.

The *binoxide of nitrogen* NO_2, may be prepared by introducing into the hydrogen apparatus some copper turnings, a little water, and nitric acid, added by the funnel. The gas may be collected over water:—

$$3Cu + 4NO_5 = 3(CuO,NO_5) + NO_2.$$

$$(3Cu + 4H,NO_6 = 3Cu,NO_6 + 4HO + NO_2).$$

The gas obtained in this manner is colourless and transparent; but in contact with air or oxygen it produces deep red fumes, which are sufficient to distinguish it from all other gases. This may be beautifully illustrated by passing some binoxide of nitrogen into a gas jar containing oxygen, standing over the pneumatic trough, or allowing the fumes to pass into the atmosphere.

Nitrous acid, NO_3.—When four volumes of dry binoxide of nitrogen are mixed with one volume of dry oxygen, and exposed to as low a temperature as can be obtained by freezing mixtures, a thin mobile liquid of a green colour is produced. This is nitrous acid, NO_3.

Hyponitric acid, NO_4.—If nitrate of lead be heated in a retort of hard glass it is decomposed into the protoxide of lead, oxygen, and hyponitric acid. It may be collected in a receiver surrounded with a freezing mixture.

Great obscurity hangs over these two compounds of nitrogen and oxygen. The red fumes produced when the binoxide of nitrogen comes in contact with oxygen are chiefly due to the formation of hyponitric acid. Other oxygen compounds are also formed.

Nitric acid, HO,NO_5 (H,NO_6).—Known as aquafortis, and one of the most important compounds of nitrogen. In hot dry climates, the soil is sometimes covered with a white, downy-looking salt. When this salt is collected, dissolved, filtered, and crystallized, it produces the nitre or saltpetre of commerce: it is a compound of potash and nitric acid. To obtain nitric acid from this salt, mix equal weights of sulphuric acid and nitrate of potash in a glass retort, cover the receiver, which may be a Florence flask, with a wet cloth, and apply heat. Red fumes at first rise in the retort, and pass to the receiver; these fumes gradually disappear, and a pale yellow liquid collects in the receiver, which is nitric acid. Decomposition:—

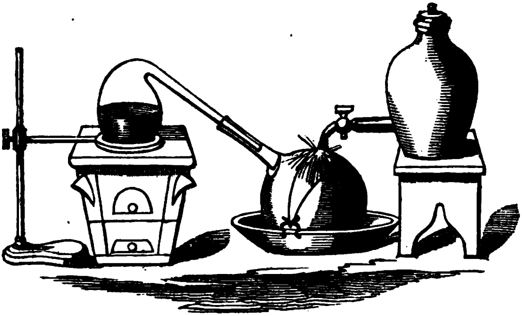


Fig. 12.

$$KO,NO_5 + HO,SO_3 = KO,SO_3 + HO,NO_5.$$
$$(K,NO_6 + H,SO_4 = K,SO_4 + H,NO_6).$$

On a large scale earthen retorts and condensing vessels are used, and nitrate of soda substituted for nitrate of potash.

Pure nitric acid is a colourless liquid. It fumes when exposed to air. It is partially decomposed by light. It is intensely acid, and reddens vegetable blues. It stains the skin and nails yellow, and is used for producing yellow colours on silken fabrics. It is also used for assaying, and as a solvent for tin in calico printing. It is a powerful oxidising agent, attacking most of the metals with great violence.

The anhydrous acid (anhydride) NO_5, is a volatile solid, crystallizing in brilliant colourless prisms. It decomposes at a temperature a little above that of boiling water; yielding hyponitric acid and oxygen. It does not, however, possess acid properties till it is dissolved in water, when it forms ordinary nitric acid.

$$NO_5 + HO = HO,NO_5 \text{ or, } (H,NO_6).$$

Nitric acid forms a series of salts, all of which are soluble in water. After a thunderstorm, nitric acid, in small quantities, has been detected in rain water. It is supposed by some, to be produced in the air by electricity.

When nitrates are heated with an excess of hydrochloric acid they are converted into chlorides, with the formation of hyponitric acid and the disengagement of free chlorine.

$$KO,NO_5 + 2HCl = KCl + 2HO + NO_4 + Cl.$$

When paper is soaked in a solution of potash, and dipped in nitric acid it becomes touch-paper. Cotton, soaked in nitric acid, washed and dried at a temperature not exceeding 120°, forms gun-cotton. The low temperature at which it explodes has prevented its substitution for gunpowder.

PHOSPHORUS.

Symbol, P; *Combining Equivalent*, 32.

This substance was discovered by the alchemists two hundred and fifty years ago. At ordinary temperatures it is a solid, waxy-looking substance, easily cut with a knife. It is poisonous. In the dark it shines with a faint blue light, and is highly combustible, often taking fire with the slightest touch or friction. Phosphorus is always kept under water, and great care is required in experiments with it.

Phosphorus is prepared from bones; they are first well calcined, reduced to powder, and mixed with two thirds their weight of dilute sulphuric acid; after some time the mixture is filtered. The sulphate of lime which has been formed by the sulphuric acid is a very insoluble compound and easily separated on the filter. The liquor is then evaporated to the consistence of treacle; mixed with charcoal powder, and transferred to an iron vessel which is raised to a high temperature till all moisture is expelled. After this mixture has been thoroughly dried it is heated in a stoneware retort, from which a large tube dips into water contained in the receiver. The phosphorus distils over in a soft condition, and is moulded in hot water into cylindrical sticks. The change which takes place in the retort is easily understood. Phosphoric acid (anhydrous) contains one equivalent of phosphorus and five equivalents of oxygen. The oxygen combines with five equivalents of carbon, forming carbonic oxide, which escapes:—

$$PO_5 + 5C = P + 5CO.$$

The phosphates play a very important part in the solid structure of animals, giving stiffness and rigidity to the bones. The dried skeleton of a man weighs about 10 or 12 lbs., and contains about 2 lbs. of phosphorus. Traces of phosphorus have been found in the brains, flesh, and blood of all animals. There is an important modification of phosphorus discovered by Professor Schrotter, called *allotropic* phosphorus, in which state the active properties of ordinary phosphorus are suspended; it is of a brick-red colour, insoluble in bisulphide of carbon, does not become luminous in the dark, has no tendency to combine with oxygen, is not poisonous, and may be carried about without fear.

Phosphorus is largely used for the manufacture of lucifer matches. As the phosphorus would be likely to ignite by exposure to the atmosphere, it is mixed up with glue or gum water. The ends of the matches are first dipped in sulphur, then into the phosphorus solution, and afterwards into a thick mixture of chlorate of potash or nitre, or some other substance rich in oxygen; they are then removed to a stove and carefully dried. About 200,000 to 250,000 lbs. of phosphorus are used annually in London in the manufacture of these matches. When the surface of the match is broken by friction the phosphorus first takes fire; this inflames the sulphur, which ignites the wood: the cracking noise is owing to the chlorate of potash.

There are three well-known compounds of phosphorus—

Hypophosphorous acid	PO
Phosphorous acid	PO_3
Phosphoric acid	PO_5

The first of these compounds is not of much interest; it is a powerful deoxidising agent, and is rather difficult of preparation.

Phosphorous acid (anhydride PO_3) is formed by the slow combustion of phosphorus in a limited supply of air.

Take a glass tube, about one foot long and half an inch in diameter; at about two inches from one end bend the tube at an obtuse angle, drawing the end of the shorter limb out into a conical shape, leaving an aperture about the size of a pin's point. Into the bent end of this tube, at the angle, place a small piece of phosphorus carefully dried by folding it in blotting-paper. Gently heat the tube over where the phosphorus lies by means of a spirit lamp, when the phosphorus will take fire and burn with a greenish flame. Only a small quantity of air enters the tube through the small aperture. The acid condenses as a white powder in the upper cold part of the tube. This is solid anhydrous phosphorous acid (phosphorous anhydride PO_3). It may also be prepared in combination with water from the teriodide of phosphorus, PI_3. To prepare this iodide of phosphorus, place a small piece of dry phosphorus in a test-tube, cover it with about eleven times its weight of iodine; the two bodies combine energetically without any application of heat. This is a singular experiment, in which two solid bodies combine chemically. After the action is over, water is added, when the following decomposition takes place:—$PI_3 + 3HO = PO_3 + 3HI$; phosphorous acid and hydriodic acid. The liquid is heated till all the hydriodic acid has volatilized, and a solution of phosphorous acid is left, $3HO,PO_3$, or H_3,PHO_6.

Phosphoric acid (anhydride, PO_5) is obtained when phosphorus is burned under a bell jar or in oxygen. It is like snow, but, in consequence of its attraction for water, it soon dissolves, and, when it once becomes the hydrated acid, the water cannot be separated. Liquid phosphoric acid may be prepared by adding little bits of phosphorus to nitric acid in a retort provided with a receiver; by this process the glacial phosphoric acid of the pharmacopœia is prepared: it is rather a troublesome and dangerous method.

Phosphoric acid may also be obtained by adding sulphuric acid to bone earth, precipitating with the carbonate of ammonia, and separating the lime salt by filtration, then evaporating and igniting in a platinum crucible; the sulphate and phosphate of ammonia volatilize; hydrated phosphoric acid remains behind. The most advantageous manner is the first process; namely, to burn phosphorus in a stream of dry air, and dissolve the solid PO_5 in water.

Phosphoric acid is not poisonous; it is, however, a powerful acid, reddens vegetable blues, and unites with bases to form a large and important class of salts.

There are three distinct modifications of this acid, supposed by some chemists to be different hydrates of the body PO_5, and hence called respectively the monohydrate, HO,PO_5, the bihydrate $2HO,PO_5$, and the tritohydrate, $3HO,PO_5$. There are, however, difficulties in the way of this hypothesis. It is more probable that the view put forward by Liebig, and now held by very many eminent chemists, is the true one. Instead of these bodies being regarded as different hydrates of the same compound, they are looked upon as the hydrogen compounds of three distinct compound radicals (see p. 18), which may, under certain circumstances, become converted into each other; and their constitution is expressed by the formulæ H,PO_6; H_2,PO_7; H_3,PO_8; respectively. Bearing in mind what was said of the theory of acids and salts at p. 19, 20, it will be seen at a glance that the hydrogen in these acids may be replaced by one, two, and three equivalents respectively of basic radicals; hence they are designated *monobasic*, *bibasic*, and *tribasic* phosphoric acid. The salts of the *monobasic* acid have the general formula MO,PO_5, or M,PO_6; the *bibasic* acid forms two series of salts, one containing two equivalents of metal, $2MO,PO_5$, or M_2PO_7, and the other containing one equivalent of a metal and one of hydrogen, MO,HO,PO_5 or M,H,PO_7. The *tribasic* acid forms three series of salts, as follows:—$3MO,PO_5$, or M_3,PO_8; $2MO,HO,PO_5$, or M_2H,PO_8; $MO,2HO,PO_5$, or M,H_2,PO_8. The following table exhibits the constitution of the phosphates of soda according to both views:—

Monobasio Acid.

Old View.	New View.
NaO,PO_5;	Na,PO_6.

Bibasic Acid.

Old View.	New View.
$2NaO,PO_5$	Na_2,PO_7.
NaO,HO,PO_5	Na,H,PO_7.

Tribasic Acid.

Old View.	New View.
$3NaO,PO_5$	Na_3,PO_8.
$2NaO,HO,PO_5$	Na_2H,PO_8.
$NaO,2HO,PO_5$	Na,H_2,PO_8.

The hydrogen in the bibasic and tribasic acids may be replaced wholly or partially by different metals; the phosphate of magnesia and ammonia, for example, contains both magnesium and ammonium, $2MgO,NH_4O,PO_5$, or Mg_2NH_4,

PO_8; microcosmic salt contains both sodium, ammonium, and hydrogen, NaO,NH_4O,HO,PO_5, or Na,NH_4,H,PO_8.

Salts of the monobasic acid produce in silver solutions a white granular precipitate; bibasic salts give a white flaky precipitate, and tribasic salts a yellow precipitate in the same circumstances.

Phosphuretted Hydrogen, PH_3.—This gas has a most disagreeable odour, resembling garlic; it is poisonous, slightly soluble in water; it precipitates several of the metals, but has no acid or alkaline properties. It may be prepared by heating in a small retort hydrated phosphorous acid, which is decomposed into phosphuretted hydrogen and hydrated phosphoric acid—

$$4(3HO,PO_3) = PH_3 + 3(3HO,PO_5);$$
$$\text{or, } 4(H_2,PHO_6) = PH_3 + 3(H_3,PO_8).$$

It is more commonly prepared by boiling in a small retort sticks of caustic potash, water, and phosphorus; the retort should be nearly filled, and provided with a long tube which should dip into water from the commencement of the operation, and heat gently applied. (See fig. 13). The gas soon comes off with a slight explosion in the retort, attended

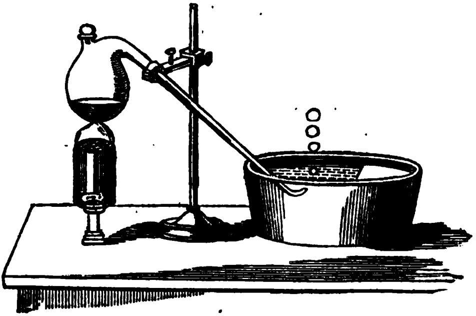


Fig. 13.

with the rushing of the water along the tube. It is to prevent any accident that the tube should be long enough to prevent the water entering the retort: as soon as the oxygen is consumed all danger is over. As the gas escapes through the water into the air, it takes fire and changes into wreaths of white smoke. The preparation of this gas is a very beautiful experiment. The potash acts merely by its presence in determining the decomposition.

The change may be represented thus—

$$4P + 3(KO,HO) = PH_3 + 3(KO,PO).$$

FOURTH DIVISION.

CARBON.

Symbol, C ; *Equivalent*, 6.

This element, in one state or another, has been known from the earliest time. It occurs in nature as the diamond, and graphite or black lead. Charcoal, coke, bone black, soot, and lamp black, are artificial varieties of carbon. In combination with oxygen it forms two compounds, carbonic oxide and carbonic acid. The latter compound in union with lime, magnesia, &c., as carbonates, forms rocks which constitute a large portion of the crust of the globe. In combination with nitrogen, hydrogen, and oxygen, it enters largely into the composition of all animal and vegetable substances.

The purest form of carbon is obtained by exposing animal or vegetable substances to heat, carefully excluding the atmosphere during the operation. This process is now usually performed in large iron retorts, somewhat similar to those which are used for making common gas. The specific gravity of the charcoal varies with the wood from which it is manufactured.

Charcoal is very combustible, and is extensively used on the Continent for fuel. When perfectly pure it burns without flame. It is a very indestructible and unchangeable substance. Wooden piles driven into the earth are usually charred to prevent decay. Charcoal has the power of condensing in its pores different gases. It has been found to absorb ninety times its bulk of ammonia, and thirty-five times its bulk of carbonic acid. It is this property which makes it useful as a deodoriser. Charcoal when newly prepared restores meat slightly tainted, either by being boiled with it or rubbed over it. It thus removes all offensive odours, and checks putrefaction. Rats buried in powdered charcoal yielded no unpleasant effluvia when opened after three months. Water is purified by passing it through filters constructed of layers of charcoal and sand. Black crayons are made from the charcoal of the willow. The charcoal used in the manufacture of gunpowder is made from dogwood or elder. Common bituminous coal is not carbon: it is a combination of many complex unions of carbon and hydrogen. Heat expels the volatile parts, and we have left coke, which is a mixture of carbon, with small quantities of metallic oxides. In the manufacture of gas these volatile products are purified and used for purposes of illumination.

Ivory- and bone-black are varieties of charcoal which result from the heating of bones in a retort, till all the volatile products are evolved. This bone black, however, only contains about 10 per cent. of carbon. The remainder is chiefly phosphate and carbonate of lime. Ivory black is used as ink in printing from copper and steel engravings. The preparation of bone black is now an important branch of chemical manufacture, as it is largely used in suga rrefining, owing to its bleaching or decolorising properties in common with all varieties of charcoal.

It may not be altogether out of place to say a few words on sugar refining. If the saccharine juices of any vegetable substance be allowed to remain at rest, especially if the temperature be high, fermentation commences, and the sugar is changed into alcohol; it is therefore desirable to convert it into a crystalline state as speedily as possible. The juice of the sugar cane is raised to a temperature of about 130°, and lime water added. The whole mass is now raised in temperature, but not boiled. A thick scum rises to the surface. When this scum breaks it is either taken off, or the juice is drawn off by a tap at the bottom of the copper. By this continual heating the water evaporates, and the liquid thickens. Sugar cannot be heated in the presence of acids or alkalies without undergoing decomposition, and molasses may be considered as the product of this decomposition. After a sweet solution of sugar and water has been obtained, a portion of lime water is added, and bullock's blood. It is said that the lime is employed to neutralize acidity, but this is a mistake: no acid exists in unfermented sugar. The proper use of lime water is this: the colouring matter of some varieties of sugar unites with lime, and forms a precipitate which gets enveloped in the albumen of the blood, and rises to the surface in scum, which is removed. The solution is now allowed to escape into a filtering apparatus. This apparatus consists of a number of bags thrust one into the other, and fastened to the bottom of a tray, into which the sugar flows. As the liquor flows through the filter bags, a portion is examined from time to time, as to its brightness. When the solution has become clear it is of a dark port wine colour. The solution is now passed through bone black or animal charcoal. This decolouring property may be shown by passing a litmus solution through a charcoal filter. The water passes through quite clear. The whole theory of this decolorising effect is however involved in some obscurity. In sugar refining, a large cylinder of iron, six or eight feet in diameter, and twelve or fifteen feet high, is packed with granulated charcoal, within two or three feet of the top. The solution percolates through, and is drawn off at the bottom. The sugar solution which entered was of a dark colour; it passes from the filter as colourless as water. In time, however, the solution becomes tinted, the operation is stopped, and the charcoal washed.

Lamp-black results from the imperfect combustion of tar, oil, resin. The greater part of the carbon is carried up in a state of finely-divided powder. When these particles come in contact with a cold body, they form a deposit of carbon, which may be scraped off. This forms the lamp-black of commerce. The operation is performed with a limited supply of air in a high chimney, containing a number of plates, upon which the lamp-black collects like soot in our own chimneys.

Plumbago or black lead was formerly thought to be a carbide of iron. The best specimens are altogether free from iron. Lead is never present in black lead. Plumbago is used for the manfacture of pencils. That of Borrowdale in Cumberland was the best in the world, but the vein is now exhausted. Most of the ordinary pencils are manufactured from a paste made of powdered plumbago, crude antimony, and sulphur.

Diamonds are carbon in a crystallized state. They were formerly obtained in India and the island of Borneo, but for the last hundred years they have been brought from Brazil. If a diamond be burned in oxygen, the sole result is carbonic acid, the same as if we had burnt a piece of charcoal. It is the opinion of some, that the diamond is of vegetable origin. Within the crystalline structure of many diamonds, the remains of insects are seen. Sir D. Brewster thinks the diamond a drop of fossilized gum.

For diamonds of moderate size, and the same quality, the value is estimated in this manner:—The weight is given in carats. These carats are subdivided into halves, quarters, eighths, sixteenths, and thirty-second parts. If four diamonds weigh 1, 2, 3, and 4 carats each, their respective value would be as 1, 4, 9, and 16: their value is in the ratio of the squares of their respective weights. The Koh-i-Noor weighed 186 carats: since it has been recut, it weighs 102 $\frac{3}{4}$ $\frac{1}{16}$ carats.

There are two well-known compounds of oxygen and carbon, carbonic acid, and carbonic oxide.

Carbonic acid, CO_2, is present, in small quantities, in the atmosphere; it occurs in combination with lime and magnesia: it is a constituent of all marbles, shells, and coral. It is also present in all spring water; it is evolved from volcanoes, and is a constant product of respiration. Fermenting liquids, decaying animal and vegetable substances, and the combustion of coal, wood, and gas, are continual sources of carbonic acid at the ordinary temperature and pressure. It is a gas having slightly acid properties: under pressure it may be liquified and solidified.

Carbonic acid may be prepared *synthetically* by burning carbon in oxygen, or *analytically* by decomposing a carbonate. Carbonate of lime is generally used for this latter purpose. When carbon is burnt in oxygen no change of volume takes place, although there is an increase of weight: 6 of carbon combining with two equivalents of oxygen, $6 + 8 + 8 = 22$.

Any carbonate yields carbonic acid when treated with any of the stronger acids. We use for this purpose hydrochloric acid and pieces of marble, which is a carbonate of lime, and a little water. The decomposition is as follows:—

$$CaO,CO_2 + HCl = HO + CaCl + CO_2;$$

$$\text{or, } (Ca,CO_3 + HCl = HO + CaCl + CO_2).$$

Lime water is a test for carbonic acid,—a milky, insoluble precipitate of carbonate of lime is formed when carbonic acid is passed through a solution of lime water.

Carbonic acid is heavier than common air; it may be poured from one vessel to another; it is not a supporter of combustion or respiration, nor is it combustible. It extinguishes flame, not for the same reason as nitrogen, by excluding oxygen; for if we mix one volume of oxygen with four of carbonic acid, it does not support combustion, although the quantity of oxygen is as great as in common air. If a living animal be immersed in an atmosphere of nitrogen, it dies the same as if it were immersed in water. The nitrogen kills by suffocation; carbonic acid kills by poison: all the other conditions of life may be present, but the inhalation of carbonic acid produces death. In winter, people burn charcoal in braziers,—a most unhealthy and dangerous practice where there is not a proper exit for the carbonic acid produced. During the fermentation of beer or wine, carbonic acid is evolved. During respiration, it is given off. To illustrate this, take a glass tube and blow into lime-water: the water becomes milky, owing to the formation of carbonate of lime.

Carbonic Oxide, CO.—The blue lambent flame sometimes seen on the top of brick-kilns and coke-ovens, and in our own fireplaces, is from carbonic oxide (CO). This compound contains one equivalent less of oxygen than carbonic acid. It may be prepared by passing carbonic acid through a red-hot iron tube. Thus,—$Fe + CO_2 = FeO + CO$. It may also be prepared by passing carbonic acid through a porcelain tube filled with charcoal. The carbonic acid, CO_2, uniting with another equivalent of carbon, becomes converted into two equivalents of carbonic oxide, $CO + CO$. A similar change to this takes place in charcoal and coke fires: the charcoal combines with the oxygen, forming carbonic acid, CO_2; it passes through the red-hot part of the fire, and is converted into carbonic oxide, CO, and burns at the surface with a blue flame, which is sometimes said to be a sign of frost. A low temperature favours the production of this compound. Carbonic oxide is generally prepared by heating sulphuric acid with oxalic acid. As soon as the temperature is sufficiently raised, the gas comes off rapidly. The decomposition which takes place may be represented thus :—oxalic acid contains two equivalents of carbon, three of oxygen, and one of water, C_2O_3HO : this equals one equivalent of CO_2, one equivalent of CO, and one of HO. When heated, the sulphuric acid unites with the equivalent of water, and the carbonic acid, CO_2, and carbonic oxide, CO, come off together. By passing the gases through a solution of caustic potash, the carbonic acid is absorbed, and the carbonic oxide, CO, passes over, which may be collected at the pneumatic trough in the usual way.

This gas is invisible, and lighter than common air. It is highly combustible, but does not support either combustion or respiration. Dumas has shown that this gas is a hundred times more poisonous than carbonic acid,—a fact which till very lately was unknown.

The compounds of carbon and hydrogen, are very complex, and are all highly combustible; they belong properly to organic chemistry. We shall describe two of these compounds.

Heavy Carburetted Hydrogen, C_4H_4, generally called olefiant gas, because it forms an oily compound when mixed with chlorine. It is lighter than common air. This gas is prepared by mixing alcohol with five or six times its weight of sulphuric acid, H,SO_4. The acid should be added a little at a time. As it is easier to measure liquids than to weigh them, one volume of alcohol and three of sulphuric acid, HO,SO_3 may be used. The flask should be shaken so as to thoroughly mix the liquids.

Alcohol, $C_4H_6O_2 = C_4H_4 + 2HO$.

$$C_4H_6O_2 + H,SO_4 = C_4H_4 + H,SO_4 + 2HO.$$

Properties.—Olefiant gas is highly combustible, but not a supporter of either combustion or respiration: when burnt it forms carbonic acid, CO_2, and water HO. Its combustibility may be shown by burning some from a jet. If a light be applied to a mixture of chlorine and olefiant gas in equal volumes, HCl is formed, and the carbon liberated in a finely divided state.

Light Carburetted Hydrogen, C_2H_4.—This gas is disengaged abundantly from coal mines and from stagnant water. The miners call it *fire-damp;* it is also called *marsh gas.* When mixed with common air it forms an explosive compound. It is rather difficult of preparation: 40 parts of acetate of soda, 60 parts of quicklime, 40 parts of hydrate of potass are mixed together and heated in a retort. The lime merely prevents the hydrate of potass attacking the glass. The reaction consists in the conversion of the acetic acid, by the aid of the elements of HO, into carbonic acid, CO_2, and light carburetted hydrogen, CH_2.

$$\text{Acetic acid, } \left.\begin{array}{r} C_4H_3O_3\\ HO\\ \hline C_4H_4O_4\end{array}\right\} = \begin{array}{l} C_2O_4 = 2CO_2\\ C_2H_4\\ \hline C_4H_4O_4.\end{array}$$

This gas is colourless, invisible, inodorous; does not support combustion or respiration. It burns with a white flame, similar to olefiant gas, but less brilliant.

Coal Gas.—Coal gas is prepared by exposing coal to a red heat in iron retorts. The variety of coal known as cannel coal is usually employed. As the coal becomes heated a variety of complex gaseous products are given off. As these products issue from the retorts they are conducted into the hydraulic main (a large horizontal cistern half filled with water). The gas passes from the hydraulic main into the refrigerator, and from the refrigerator into the purifier, whence it is conducted into a large metallic gasometer—from which the distribution pipes proceed.

In the hydraulic main the gas is partially cooled and a portion of the ammonia absorbed. In the refrigerator the gas is cooled and the tar and ammoniacal liquor deposited. In the purifier, which is a large vessel containing slaked lime mixed with water and kept in agitation, the gas parts with its sulphuretted hydrogen and carbonic acid.

Ordinary coal gas, although subjected to this purification, usually contains a variable amount of noxious substances.

The compounds of carbon and hydrogen are the only valuable constituents of coal gas. It contains about fifty per cent. by volume of olefiant gas, upon which its illuminating property chiefly depends; from twelve to twenty per cent. of light carburetted hydrogen (marsh gas), with variable quantities of nitrogen, hydrogen, carbonic oxide, sulphuretted hydrogen, volatile oils, and prussic acid.

The Davy Lamp.—A mixture of fire-damp and air forms an explosive compound. Common explosions of gas are analogous to those which occasionally take place in coal mines. To prevent these accidents, a lamp was invented by Sir H. Davy, known as the Davy lamp.

Fire-damp alone is not explosive; but when mixed with from eight to ten volumes of air it is highly dangerous. With three or fifteen volumes of air it is not explosive. The small quantity of oxygen in the one case is not sufficient to burn the gas; in the other the large quantity of air conducts away the heat and lowers the temperature below the point necessary to ignite the gas. It is on this principle that the Davy lamp is constructed. Fire-damp does not take fire unless raised to a white heat. When the fire-damp issues from a seam of coal, meeting the white flame of a lamp or candle, explosion immediately takes place.

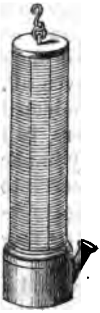


Fig. 14.

The safety lamp consists of an ordinary oil lamp, surrounded by a cylinder of wire gauze, the apertures of the gauze not being more than the 400th of an inch in diameter (See fig. 14). When this lamp is carried into an explosive mixture, no explosion takes place. The gases traverse the apertures of the gauze, and burn inside the lamp; but the flame does not pass outside to kindle the atmosphere of gas which surrounds the lamp. The lamp makes explosion impossible, not by preventing the gas entering inside, but by preventing the flame reaching the gases outside. In traversing the wire gauze the temperature is lowered, heat is conducted away, and the temperature falls below the point necessary to kindle the gases. These effects may be studied by suspending a Davy lamp in a jar, and gradually admitting a stream of coal gas from below. The flame of the lamp increases in length, in the interior of the gauze cylinder, the gas and air burn, and the flame of the lamp is extinguished; but as the atmosphere in the bell jar becomes purer, the lamp again rekindles. These phenomena are so remarkable that the lamp becomes a true indicator of the state of the atmosphere in the mine; and this is the proper and true use of the Davy lamp.

Structure of Flame.—The flame of a candle is not, as it appears, a solid cone of flame. If a piece of wire gauze be pressed from the top over the flame of a candle or common lamp, there appears in the centre a dark spot surrounded by a luminous ring. This dark spot is the combustible hydrogen and carbon of the tallow (See fig. 15). These gases cannot burn until they pass to the surface, to obtain oxygen from the atmosphere, when they are converted into carbonic acid and water. The hydrogen burns with a blue flame, with not much illuminating power, but with great heat. The oxygen of the air has a greater affinity for hydrogen than for carbon. The hydrogen burns first, and this is the blue flame around the central dark spot; the carbon now liberated from the hydrogen passes upward, and, being intensely heated by the combustion of the oxygen and hydrogen, burns with a yellow light; this is the luminous part of the flame, and is due to the solid white hot particles of carbon previously contained


Fig. 15.

in combination with hydrogen in the tallow. The conical form of the flame is owing to the ascent of heated currents of air. A tall glass chimney is placed over lamps, so as to increase the upward current of air and supply a larger amount of oxygen. It is remarkable that the heat evolved by combustion is due to the quantity of oxygen consumed, and not the amount of combustible matter with which it unites, as the following Table illustrates :—

1 lb. of wood charcoal raised 78 lbs. of water from 32° to 212°.
1 lb. of alcohol raised 68 lbs. of water from 32° to 212°.
1 lb. of oil or tallow raised 90 lbs. of water from 32° to 212°.
1 lb. of hydrogen raised 236 lbs. of water from 32° to 212°.

The quantity of oxygen consumed under these circumstances is very different, while the weight of the combustible body was in each case one pound.

If a jet of air be introduced into the flame of a candle or lamp, by means of a blowpipe, a double combustion takes place. Two long pointed flames are observed, the outer one being of a yellowish colour, and the inner one blue (See fig. 16). The space between these flames is filled with combustible matter at a very high temperature; this is called the deoxidising or reducing part of the flame. If the oxide of any metal be placed in this part of the flame, fixed on the end of a fine platinum wire, or supported on a piece of charcoal, it is reduced to a metallic state.

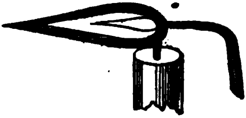


Fig. 16.

Observations of great importance are often made with the blowpipe. The skilful use of this instrument can only be obtained by constant practice.

SILICON.

Symbol, Si; *Combining Equivalent*, 21·35.

This element is sometimes called silicium. It does not occur in a free state; but in union with oxygen it forms silica, SiO_3. It enters largely into the composition of flints, sand, quartz, and other mineral substances, forming, it is thought, a sixth part of the surface of the earth. Silicon is a brown powder, difficult to produce, and is of no importance, except to the scientific chemist.

Silica, SiO_3, is the only known oxide. When nearly pure, as in rock crystal, it is colourless and transparent. It may be prepared by mixing equal parts of fluor-spar, CaF, powdered glass, and sulphuric acid together in a flask. A glass tube, at least half an inch in diameter, should proceed from the flask and dip three quarters of an inch into mercury. The jar containing the mercury is now nearly filled with water; a gentle heat is applied to the flask, when the silica, in a gelatinous condition, separates, and rises through the mercury to the surface of the water. The decomposition is rather complex. The first effect is the disengagement of hydrofluoric acid, which in contact with the silica of the glass, undergoes decomposition, forming water and the terfluoride of silicon, which escapes from the flask as a gas; but when it comes in contact with water, after it has passed through the mercury, it is decomposed into silica and hydrofluosilicic acid.

Decomposition in the flask:—

$$3HF + SiO_3 = 3HO + SiF_3.$$

Decomposition in contact with the water, after passing through the mercury:—

$$3SiF_3 + 3HO = SiO_3 + \underbrace{3HF + 2SiF_3}.$$

The silica is collected on a cloth filter, dried, and raised to a red heat to expel water.

Silica is an acid. It neutralizes many bases, forming salts, known as silicates. Glass, porcelain, and pottery, are silicates of potash, alumina, and other bases. Some silicates are soluble in water: if the bases with which the acid has united be in excess, the salts are soluble; but if the acid be in excess, then they are insoluble. Silica in its crystallized state, coloured with some metallic oxide, forms precious stones, such as blood-stone, agate, jasper, onyx, carnelian, and opal. Silica exists abundantly in many plants. It gives stiffness to the straw of wheat, barley, and oats.

BORON.

Symbol, B; *Equivalent*, 10·90.

This element is analogous to silicon. It may be obtained by heating boracic acid with potassium in a platina crucible. It is a brown powder, and, when heated, burns, forming boracic acid, BO_3. It is found in the lagoons of Tuscany, in combination with oxygen, as boracic acid, BO_3, from which source it is chiefly derived. This acid is also easily prepared by decomposing a hot solution of borax with sulphuric acid. The biborate of soda, or borax, is employed for soldering metals, forming the glazes for porcelain, and as a flux for examining metallic compounds before the blowpipe. A loop made of platinum wire is strongly heated and dipped into borax powder: a series of experiments can now be made. A salt of manganese heated in the outer flame of the blowpipe, with the borax on the wire, gives a violet-coloured bead; cobalt, blue; chromium, green; copper, red.

FIFTH DIVISION.

HYDROGEN.

Symbol, H; *Combining Equivalent*, 1.

Hydrogen, at all known temperatures, is a gas, colourless, without odour, about 14½ times lighter than common air, and highly combustible. This gas exists largely in the organic and inorganic kingdoms; it is a constituent of most substances used either for fuel or illumination; it is most probably an essential constituent of all acids, hydrochloric, HCl, hydrobromic, HB; hydriodic, HI, hydrosulphuric, HS, sulphuric acid, H,SO_4, &c. In combination with animal or vegetable substances it is found chiefly as water and ammonia.

Preparation.—This gas is prepared by deoxidizing water, either by the decomposing agency of voltaic electricity, as explained under oxygen, or by adding small pieces of granulated iron or zinc to one part by measure of sulphuric acid and four or five parts of water in the apparatus figured on page 41. Neither iron nor zinc combines directly with sulphuric acid, but the oxides of these metals combine easily. Decomposition:—

$$Zn + HO + SO_3 = ZnO,SO_3 + H \text{ free.}$$

Or, as is most probably the case, the action is simply a displacement of H by the Zn:—

$$Zn + H,SO_4 = Zn,SO_4 + H \text{ free.}$$

Or hydrochloric may be used instead of sulphuric acid; then—

$$HCl + Zn = ZnCl + H \text{ free.}$$

By passing steam through a red-hot iron tube, the oxygen combines with the iron, and the hydrogen escapes as gas, thus:—

$$\text{(Steam)}, HO + Fe = FeO + H \text{ free.}$$

By passing a small piece of sodium or potassium up into an inverted tube, filled with water and standing on the pneumatic trough:—

$$Na + HO = NaO + H.$$

The sodium or potassium dissolves in the water, and the hydrogen collects in the upper part of the tube.

Hydrogen from its extreme lightness was formerly used for filling balloons; its use is now superseded by employing coal gas and larger balloons.

The property of diffusion, which is common, in a greater or less degree, to all gases and vapours, is well seen in the following experiment with hydrogen and carbonic-acid gas, two bodies of very different specific gravities. If two bottles, the one containing hydrogen, and the other carbonic acid, be joined together with a glass tube about a foot long and ⅜ of an inch in diameter, and the hydrogen placed in the upper bottle and the carbonic acid in the lower, we shall find after a few hours, that the carbonic acid, contrary to its gravity, and the hydrogen, have uniformly mixed throughout. This diffusive power of gases is of great importance in the economy of the universe. Were it not for this diffusion we should have the different gases and vapours arranging themselves on the earth's surface according to their respective weights. This diffusive power of hydrogen renders it impossible to keep it for any length of time in bladders, or the best constructed gas holders, and when mixed with air from any leaking, the mixture becomes explosive.

Hydrogen is not a supporter of combustion: if a lighted taper be plunged into a jar of hydrogen, it is extinguished; the hydrogen burns only at the top of the jar, where it comes in contact with the oxygen of the air. Hydrogen burns with a pale blue flame without much light, but intense heat. The temperature may be increased by mixing hydrogen and oxygen in the proportion of two volumes of the former to one of the latter, and burning it from a safety jet. This is called the oxyhydrogen blowpipe. It is, perhaps, best to use two separate gas-holders for this experiment. When hydrogen is mixed with common air or pure oxygen, it is highly explosive; great care should therefore be taken to allow the first portions of gas to escape before applying a light. Soap bubbles blown with hydrogen by means of a common pipe attached to a bladder provided with a stop-cock to regulate the emission of the gas, rise in the air with considerable velocity: if a lighted taper be applied to the bubble as it ascends a slight explosion takes place. If a mixture of one volume of oxygen and two volumes of hydrogen be taken, and soap bubbles blown on a small cheese-plate, so that a good head of bubbles is formed and a light applied, the explosion is very violent: a small bladder filled with a mixture in the same proportion, may be exploded by pricking a hole with a pin and applying a light. If a jet of hydrogen is burned in a dry glass jar, the sides become covered with moisture, owing to the formation of water by the union of the inflamed hydrogen with the oxygen of the air. If a tube of glass about one inch in diameter, and two feet long, be slowly passed downwards over a burning jet of hydrogen, a musical note is produced. If a piece of spongy platinum be brought near a jet of hydrogen, it is at once kindled, and the platinum becomes red hot. It is thought that the hydrogen along with atmospheric oxygen are so condensed in this experiment that they are brought within the sphere of attraction necessary for combination. The same takes place with a mixture of the two gases.

The chemical relations of hydrogen assign to it the character of a metal; and although it has never as yet been either liquefied or solidified, it is not improbable that these difficulties in the way of its being considered a highly elastic gaseous metal, will disappear as the subject is more investigated.

There are two compounds of hydrogen and oxygen; *water*, HO, and the *binoxide of hydrogen*, HO_2.

The composition of water was first discovered by Mr. Cavendish, but its exact composition has occupied the attention of chemists for some years. There are two methods of determining the composition of any compound body; these are, *analysis*, by which the compound is resolved into its constituents; and *synthesis*, by which the elements or constituents are made to unite to produce the compound. Both these methods have been brought to bear on the determination of the proportions of oxygen and hydrogen in water. The most elegant illustration of the analysis of pure water is its decomposition by voltaic electricity, alluded to under the head of oxygen. In this experiment it appears that the combining volume of hydrogen is twice that of oxygen, and the combining proportion by weight, as 8 of oxygen to 1 of hydrogen.

Water, when pure, is colourless, except when viewed in a large mass, when it appears of a beautiful blue colour. It is inodorous, tasteless, and transparent. It attains its maximum density at 40°; it boils under ordinary atmospheric pressure at 212°. It evaporates at all temperatures. It is 815 times heavier than air, and an imperial gallon weighs 10lbs. avoirdupois. It is compressible in a very slight degree.

Water never occurs in nature perfectly pure. Filtration will remove suspended impurities, and distillation will free the liquid from those dissolved. Water unites with acids and bases, forming hydrates; this combination is sometimes very energetic, producing great heat, as in the slaking of lime, or by mixing sulphuric acid with water. The solvent properties of water are far greater than those of any other liquid. Many of the salts are soluble in water, increasing with an increase of temperature, and depositing crystals as the solution cools. Common salt and the salts of lime are, however, an exception to this general rule.

Water dissolves most of the gases, although in very unequal quantities. Hydrogen, oxygen, and common air are absorbed to a small extent, while ammonia and hydrochloric acid are dissolved to a very great extent; between these extremes we have sulphuretted hydrogen, chlorine, carbonic acid, and many other gases. The colder the water the more gas is dissolved; and on boiling, these gases are liberated.

The *deutoxide of hydrogen*, HO_2, or oxygenated water, is difficult of preparation. It has very few practical applications, and is only regarded as a chemical curiosity.

RÉSUMÉ.

The Law of Definite Proportion.—Those bodies which combine chemically, combine in fixed and definite proportions.

The Law of Multiple Proportion.—When one body is capable of uniting with another in several proportions, these proportions bear a simple relation to each other, as O, O_2, O_3, O_4, &c.

Equivalent numbers are an expression of the relationship by weight of the different bodies which enter into chemical combination.

OXYGEN.

Symbol, O; *Equivalent*, 8.

Discovered by Priestley in 1774; called dephlogisticated air, afterwards vital air.

Physical Properties.—Permanently gaseous, without colour, odour, or taste.

Density.—1·106.

Chemical Properties.—A simple body very slightly soluble in water. It is present in all common cases of combustion. The term *combustion* is used to express chemical action usually attended with disengagement of light and heat. It is also present in all those cases of slow combustion in nature which consist in a slow absorption of oxygen, and in which the temperature does not rise sufficiently to cause any sensible evolution of either light or heat. All cases of combustion are attended with the formation of new combinations.

Preparation.—First Method.—By the decomposition of the binoxide of manganese in an iron bottle raised to a red heat—

$$2MnO_2 = Mn_2O_3 + O.$$

Second Method.—Decompose by heat in a Florence flask the chlorate of potash. An equal part of the binoxide of manganese should be mixed with the salt. The manganese may be disregarded in the decomposition, as it undergoes no change—

$$KO,ClO_5 = KCl + 6O,$$

$$\text{or, } (KClO_6 = KCl + 6O).$$

Third Method.—Decompose by heat the red oxide of mercury—

$$HgO = Hg + O.$$

Fourth Method.—By distilling in a glass retort the binoxide of manganese with sulphuric acid. The oxygen in this case is not very pure—

$$MnO_2 + HO,SO_3 = MnO,\ SO_3 + O + HO;$$

$$\text{or, } (MnO_2 + H,SO_4 = Mn,SO_4 + HO + O).$$

Fifth Method.—By the voltaic decomposition of water.

SULPHUR.

Symbol, S; *Equivalent*, 16.

Physical Properties.—At ordinary temperatures sulphur is solid, of pale yellow colour, without taste, having a peculiar and characteristic odour. It is opaque, and exists both in an amorphous and crystalline state. It is a bad conductor of heat and electricity. At a temperature of 230° it melts into a thin liquid; at a temperature of 430° it becomes thick and black, like treacle; and at 790° it again becomes fluid and volatilizes. The condensed vapour forms sublimed sulphur.

Specific Gravity.—2·07.

Chemical Properties.—Insoluble in water, but soluble in the bisulphide of carbon and essential oils.

Sulphur burns with a blue flame, forming sulphurous acid, SO_2. Sulphur combines with carbon at a red heat, forming the bisulphide of carbon.

Sulphur is found in volcanic districts, and is imported to this country chiefly from Tuscany and Naples.

Uses.—For the manufacture of gunpowder, lucifer matches, sulphuric acid, and in medicine.

Sulphurous Acid.

Symbol, SO_2; *Equivalent*, 32.

Physical Properties.—A colourless gas, with sharp odour and acid taste. It liquefies under a pressure of two atmospheres.

Specific Gravity.—2·234.

Chemical Properties.—Composed of one equivalent of sulphur and two of oxygen, very soluble in water, and exhibits a strong acid reaction on test-paper. It is not combustible or decomposable by heat, nor is it a supporter of combustion or respiration.

Preparation.—By burning sulphur in air or oxygen. By heating in a flask copper or mercury with sulphuric acid.

Decomposition :—

$$\left.\begin{matrix}Cu\\Hg\end{matrix}\right\} + 2(SO_3,HO) = \left\{\begin{matrix}CuO,SO_3\\HgO,SO_3\end{matrix}\right. + 2HO + SO_2;$$

$$\text{or, }\left(\left.\begin{matrix}Cu\\Hg\end{matrix}\right\} + 2(H,SO_4) = \left\{\begin{matrix}Cu,SO_4\\Hg,SO_4\end{matrix}\right. + 2HO + SO_2\right).$$

In preparing large quantities of solution of the acid, carbon is commonly substituted for copper.

Uses.—As a bleaching agent.

SULPHURIC ACID.

	Symbols.	Equivalents.
Anhydride*.	SO_3	40
Real or hydrated acid	HO,SO_3; or, (H,SO_4).	47

Known for a long time as oil of vitriol; analysed by Lavoisier.

Physical Properties.—Sulphuric acid in its anhydrous state, SO_3, is a white, solid, crystalline body. The hydrated acid is a colourless liquid with an oily appearance, without odour.

Specific Gravity of Sulphuric Acid.—As HO,SO_3, 1·843.

Chemical Properties.—The anhydride is composed of one equivalent of sulphur and three of oxygen. The ordinary acid has the following composition, HO,SO_3, or (H,SO_4). It has a great affinity for water, and is used frequently as a desiccating agent. It carbonizes organic substances, combining with the elements of water and liberating the carbon. It is a powerful acid.

At a red heat the anhydride yields up an equivalent of oxygen, becoming sulphurous acid, SO_2.

The nitrate of baryta will give a white precipitate in any solution containing sulphuric acid.

Preparation of Sulphuric Acid.—Fumes of hyponitric acid, NO_4, and sulphurous acid, SO_2, pass simultaneously into a leaden chamber, into which a jet of steam plays.

Two equivalents of sulphurous acid act on one of hyponitric acid, NO_4, reducing it to the binoxide of nitrogen, NO_2, which takes two equivalents of oxygen from the air in the chamber, converting it again into hyponitric acid, to be acted on by two more equivalents of sulphurous acid:—

$$2SO_2+2HO+NO_4=2(SO_3,HO)+NO_2.$$

The acid is afterwards reduced to its proper strength by concentration.

Nordhausen Acid.— This acid is prepared from the sulphate of iron or green vitriol by distillation. The vapours of sulphuric acid pass over, which condense into a heavy dark fuming liquid.

The anhydrous acid may be prepared by distilling the Nordhausen acid, and surrounding the receiver with a freezing mixture.

Sulphuric acid has a very extended application in the arts; it is used in fancy dyeing, and the manufacture of many chemical products.

SULPHURETTED HYDROGEN—HYDROSULPHURIC ACID.

Symbol, HS; *Equivalent*, 17.

Discovered by Scheele in 1776.

Physical Properties.—A colourless gas, having a strong, disagreeable odour, like rotten eggs. It liquefies under a pressure of 16 atmospheres.

Specific Gravity.—1·177.

Chemical Properties.—It is composed of one equivalent of hydrogen, and one of sulphur, HS. Very soluble in water, in which state it is usually employed in the laboratory. On exposure to air, the sulphur is precipitated. Sulphuretted hydrogen is a feeble acid; it is decomposed by sulphurous acid into water and sulphur:—

$$2HS + SO_2 = 2HO + 3S.$$

It is also decomposed by chlorine into hydrochloric acid and sulphur:—

$$Cl + HS = HCl + S.$$

The inhalation of this gas is very injurious, and chlorine is employed to decompose it. It frequently exists as a natural exhalation from drains and decomposing matter.

Preparation.—By decomposing the protosulphide of iron with sulphuric acid:—

$$FeS+HO,SO_3=FeO,SO_3=HS;$$

$$\text{or, } (FeS+H,SO_4=Fe,SO_4+HS).$$

By decomposing the tersulphide of antimony with hydrochloric acid:—

$$SbS_3+3HCl=SbCl_3+3HS.$$

Sulphuretted hydrogen has no use in the arts.

CHLORINE.

Symbol, Cl; *Equivalent*, 35·5.

Discovered in 1774 by Scheele, and named dephlogisticated marine acid, muriatic oxygen acid, oxymuriatic gas.

Properties.—A gas of greenish colour, having a strong suffocating odour; somewhat resembling seaweed. It liquefies under a pressure of five atmospheres.

Specific Gravity.—2·47.

Chemical Properties.—A simple body, very soluble in water. Chlorine has a great affinity for hydrogen. Many of the hydrogen compounds are decomposed by this element. A mixture of equal volumes of chlorine and hydrogen explodes with great violence on exposure to sunlight.

Moist chlorine destroys organic colouring matters. It is also used as a disinfectant. Chlorine has a feeble affinity for oxygen.

Preparation of Chlorine Gas.—1st method.—Heat gently in a retort or flask the peroxide of manganese and hydrochloric acid. Let them be well mixed before applying the spirit lamp:—

$$2HCl+MnO_2=2HO+MnCl+Cl.$$

The gas may be collected by displacement. If a solution of the gas be required, it must be passed into Wolfe's apparatus.

2nd method.—By gently heating in a flask common salt, sulphuric acid, and the peroxide of manganese:—

$$MnO_2 + NaCl + 2(HO,SO_3) = NaO,SO_3 + MnO,SO_3 + 2HO + Cl;$$

$$\text{or, } (MnO_2 + NaCl + 2(H,SO_4) = Na,SO_4 + Mn,SO_4 + 2HO + Cl).$$

Uses.—As a bleaching agent, and for disinfection.

HYDROCHLORIC ACID.

Symbol, HCl; *Equivalent*, 36·5.

Discovered in 1650, and called spirit of salt, marine acid, muriatic acid.

Physical Properties.—A colourless gas, having a sharp, acid odour, and fumes on exposure to air. It is liquefied under a pressure of 40 atmospheres.

Specific Gravity.—1·245.

Chemical Properties.—A body composed of one equivalent of chlorine united to one of hydrogen. It is very soluble in water. The ordinary acid of Sp. Gr. 1·100 is a solution of this gas, and consists of ($HCl + 16HO$).

Preparation of Hydrochloric Acid Gas.—Heat common salt in a flask with concentrated sulphuric acid:—

$$NaCl + HO,SO_3 = NaO,SO_3 + HCl;$$

$$\text{or, } (NaCl + H,SO_4 = Na,SO_4 + HCl.)$$

Uses.—For the extraction of gelatine from bones, and for disinfectants. It is one of the most important reagents in the possession of chemists.

AQUA REGIA.

This is a mixture of one volume of nitric acid with three volumes of hydrochloric acid. Aqua regia is used to dissolve gold:—

$$3(HO,NO_5)+6HCl+4Au=3NO_2+9HO+2Au_2Cl_3.$$

Bromine and iodine are analogous in their chemical properties to chlorine.

AZOTE.—NITROGEN.

Symbol, N; *Equivalent*, 14.

Discovered by Lavoisier in 1775.

Physical Properties.—Permanently gaseous, without colour, odour, or taste.

Density.—0·972.

Chemical Properties.—A simple body, slightly soluble in water. It is neither a supporter of combustion nor animal respiration, nor is it combustible. Under ordinary circumstances its properties are remarkably inert, but in combination it forms compounds of a very energetic and powerful character.

Preparation.—1st method.—By burning phosphorus in a confined portion of common air. This is best done in a bell gas jar standing over water. The phosphorus combines with the oxygen of the air, forming phosphoric acid (anhydride), PO_5, which dissolves after a short time in the water.

2nd method.—By passing a current of air through a tube of hard glass loosely filled with copper turnings. The tube is raised to a red heat. Oxide of copper is formed, CuO, and nitrogen liberated.

ATMOSPHERIC AIR.

First analysed by Lavoisier in 1777.

Physical Properties. — Permanently gaseous, without colour, odour, or taste.

Density, compared with water, 0·0013.

Composition of air: 100 parts by weight contain—

23 of oxygen
77 of nitrogen
100

By volume— 79·19
20·81
100·00

The air contains vapour of water, carbonic acid, and traces of ammonia.

Air may be analysed by the process described in the second method for preparing nitrogen, by allowing the nitrogen to pass into an exhausted glass globe. Previous to the application of heat the glass tube and copper turnings are carefully weighed, and also the glass globe. The increase of weight of the globe after the experiment gives the nitrogen, whilst the increase in the weight of the copper is oxygen.

A less exact method is to introduce a piece of phosphorus, by means of a wire, into a graduated glass tube standing over water. The quantity of air in the tube is easily determined, by holding the tube upright and noticing the point to which the water rises. The whole is now left for twenty-four hours. The oxygen combines with the phosphorus, forming phosphoric acid, which dissolves in the water. The phosphorus may be withdrawn and the residual gas read off, which is nitrogen.

Air may also be analysed with the eudiometer.

PROTOXIDE OF NITROGEN (LAUGHING GAS).

Symbol, NO; *Equivalent*, 22.

Physical Properties.—A gas, without colour, having a sweet taste. Liquefied by a pressure of 50 atmospheres.

Specific Gravity.—1·525.

Chemical Properties.—It supports combustion almost equal to oxygen. Cold water dissolves nearly its own volume of the gas. Warm water must be used in the pneumatic trough. The most remarkable feature of the gas is its exhilarating effects on the animal system. When inspired the gas should be quite pure, and only inhaled for a short time.

Preparation.—By decomposing the nitrate of ammonia by heat in a retort or flask:—

$$NH_4O,NO_5 = 2NO + 4HO.$$

It has no use in the arts.

BINOXIDE OF NITROGEN.

Symbol, NO_2; *Equivalent*, 30.

Physical Properties.—Permanently gaseous, and without colour. Slightly soluble in water.

Specific Gravity.—1·038.

Chemical Properties.—Composed of one equivalent of nitrogen to two of oxygen. Neutral to test-paper.

In contact with air or oxygen it changes into deep red fumes of hyponitric acid, NO_4.

It is not combustible.

Preparation.—By the decomposition of nitric acid by copper in the cold :—

$$4(HO,NO_5) + 3Cu = 3CuO,NO_5 + NO_2 + 4HO.$$

This gas has no useful application.

NITRIC ACID.

	Symbols.	Equivalents.
Anhydride . .	NO_5.	54
Real acid . .	HO,NO_5; or, (H,NO_6).	63

Sometimes called aquafortis; spirits of nitre.

Physical Properties.—When anhydrous it is a white crystalline solid. In its hydrated state it is a liquid. When concentrated it gives off fumes on exposure to air.

Specific Gravity of the Concentrated Acid.—1·522.

Chemical Properties.—Nitric acid (anhydride) is composed of one equivalent of hydrogen and five of oxygen. In its hydrated state it contains from one to two equivalents of water. Ordinary nitric acid contains from three to four equivalents of water, $NO_5,4HO$.

It is a powerful acid, and attacks the metals with great energy.

Nitric acid stains the skin yellow.

The hydrated acid may be prepared by heating gently in a flask the nitrate of soda with sulphuric acid.

Decomposition :—

$$NaO,NO_5+2(HO,SO_3)=HO,NO_5+\underbrace{NaO,SO_3+HO,SO_3};$$

$$\text{or, }(NaNO_6+2(H,SO_4)=H,NO_6+\underbrace{Na,SO_4+H,SO_4}).$$

Nitric acid is largely used in the arts for etching, dyeing, &c.

AMMONIA.

Symbol, NH_3, HO; or NH_4O; *Dry Gas*, NH_3.

Sometimes called the volatile alkali; spirits of hartshorn.

Physical Properties.—A colourless gas, having a suffocating odour. It liquefies under a pressure of 7 atmospheres.

Specific Gravity.—0·597.

Chemical Properties.—Composed of one of nitrogen to three of hydrogen. Very soluble in water. It is a powerful base, turns vegetable reds blue, and combines with great energy with acids to form salts.

Ammonia is decomposed at a red heat. It is not combustible or a supporter of combustion; and to inhale the gas is very injurious.

Chlorine gas decomposes ammonia into the hydrochlorate of ammonia and nitrogen.

Preparation of Gaseous Ammonia.—Heat a mixture of sal-ammoniac and quicklime in a flask or retort :—

$$NH_4Cl + CaO = CaCl + HO + NH_3.$$

Liquid ammonia is prepared by dissolving the gas in water.

Ammonia is employed in medicine and dyeing. Its uses in nature are very numerous.

PHOSPHORUS.

Symbol, P; *Equivalent*, 32.

This element was discovered by Brandt, and made public in 1737.

Physical Properties.—A solid, generally of a pale straw colour, transparent, soft, and flexible. It is poisonous, and has a strong odour. In the dark it is luminous. There is another kind of phosphorus, called red or amorphous phosphorus, in which all the active properties of ordinary phosphorus are absent.

Specific Gravity.—1·77.

Chemical Properties.—A simple body, insoluble in water, but soluble in the bi-sulphide of carbon.

Phosphorus absorbs oxygen with great avidity. It combines with hydrogen in its nascent state.

Preparation.—By distilling at a red heat a mixture of charcoal and phosphate of lime (bone earth).

Phosphorus is extensively used in the manufacture of lucifer matches.

PHOSPHORIC ACID.

	Symbols.	Equivalents.
Anhydride . .	PO_5	72
Ordinary acid (tribasic) .	$3HO,PO_5$; or, (H_3PO_8).	99

Physical Properties.—In its anhydrous state it is a white solid. There are three modifications of this acid, all soluble in water :—

$$PO_5,HO;\ PO_5,2HO;\ PO_5,3HO;$$
$$\text{or, } (H,PO_6;\ H_2,PO_7;\ H_3,PO_8).$$

Anhydrous phosphoric acid absorbs water with great avidity, passing gradually into its hydrated state.

Chemical Properties.—Phosphoric acid is a very energetic acid. It is decomposed with great difficulty.

Preparation.—By burning phosphorus in dry air or oxygen, white flakes of anhydrous phosphoric acid are formed, which readily dissolve in water.

By distilling phosphorus with nitric acid.

PHOSPHORETTED HYDROGEN.

Symbol PH_3; *Equivalent*, 35.

Physical Properties.—A gas without colour, having a strong, disagreeable odour, like garlic.

Specific Gravity.—1·18.

Chemical Properties.—Composed of one equivalent of phosphorus to three of hydrogen. Has no action on vegetable colours. Phosphoretted hydrogen inflames spontaneously in air, giving rise to phosphoric acid and water.

Preparation.—By heating a mixture of the hydrate of potash with phosphorus.

By decomposing water with the phosphide of calcium.

CARBON.

Symbol, C; *Equivalent*, 6.

Physical Properties.—Solid, very infusible, without taste, sometimes without colour, as the diamond. Generally of a black colour.

Density varies from 2 to 3·55.

It is found in nature crystallized and amorphous.

Carbon crystallized is a hard, dense body; a good conductor of heat and electricity.

Amorphous carbon is less dense, and a bad conductor of heat and electricity.

Varieties.—The diamond, graphite or plumbago, coke, charcoal, bone-black, lampblack, anthracite.

Chemical Properties.—A simple body, insoluble in water and other liquids. It has a great affinity for oxygen at high temperatures, and is very combustible.

Preparation of Carbon.—Charcoal is prepared by burning wood with a limited supply of air.

Lampblack is prepared by burning oil or resinous substances. The lampblack collects in fine powder in the upper part of the chimney.

Coke is prepared by the combustion of coal in closed retorts, as in the manufacture of gas.

Bone charcoal is prepared by calcining bones in closed vessels.

Uses.—For the manufacture of pencils, Indian ink, black crayons; purifying water.

Animal charcoal is used as a decolorising agent in sugar refineries.

CARBONIC ACID.

Symbol, CO_2; *Equivalent*, 22.

Physical Properties.—A heavy, colourless gas with a faint odour and taste. Soluble in its own volume of water. May be liquefied by pressure.

Specific Gravity.—1·53.

Chemical Properties.—Composed of one equivalent of carbon and two of oxygen. It extinguishes flame, and rapidly destroys animal life; it acts as a narcotic poison. Decomposed by passing over red-hot charcoal or iron.

Preparation.—1st. By burning charcoal in O.

2nd. By decomposing carbonate of lime, CaO,CO_2 (Ca,CO_3) with hydrochloric acid:—

$$CaO,CO_2 + HCl = CaCl + HO + CO_2.$$

BISULPHIDE OF CARBON.

Symbol, CS_2; *Equivalent*, 38.

Discovered by Lampadius in 1796.

Physical Properties.—A thin, mobile, colourless liquid with a most disagreeable odour, and very volatile.

Specific Gravity.—1·273.

Chemical Properties.—Composed of one of carbon and two of sulphur. Without action on vegetable colours, insoluble in water, but very soluble in alcohol. It burns, giving rise to sulphurous and carbonic acids. It forms a detonating mixture with oxygen when inflamed.

Bisulphide of carbon is somewhat analogous to carbonic acid.

Uses.—Vulcanizing indiarubber and gutta percha, and killing moths.

LIGHT CARBURETTED HYDROGEN.

Symbol, C_2H_4; *Equivalent*, 16.

Known as inflammable air, marsh gas, fire-damp.

Physical Properties. — Permanently gaseous, without colour, but having a disagreeable odour.

Specific Gravity.—0·556.

Chemical Properties.—It is composed of two equivalents of carbon and four of hydrogen, neutral in its action upon vegetable colours, and insoluble in water.

It is not a supporter of combustion or respiration.

It is combustible; when mixed with common air it detonates violently, forming carbonic acid and water.

Preparation.—Sixty parts of quicklime, forty parts of solid caustic potash, and forty parts of the crystallized acetate of soda, are strongly heated together in a retort. The reaction consists in the conversion of acetic acid by the elements of water into carbonic acid and light carburetted hydrogen. The affinity of potash for carbonic acid is the determining cause. The lime merely prevents the potash from fusing and attacking the glass.

Decomposition:—

$$\left.\begin{array}{r} \text{Acetic acid,}\,C_4H_3O_3 \\ HO \\ \hline C_4H_4O_4 \end{array}\right\} = \begin{array}{l} C_2O_4 = 2CO_2 \\ C_2H_4 \\ \hline C_4H_4O_4. \end{array}$$

This gas has no use in the arts.

HEAVY CARBURETTED HYDROGEN.

Symbol, C_4H_4; *Equivalent*, 28.

Sometimes called olefiant gas.

Physical Properties.—At ordinary temperatures a gas, but not permanently gaseous. It is without colour, and has the odour of ether.

Specific Gravity.—0·978.

Chemical Properties.—It is composed of four equivalents of carbon and four of hydrogen. Neutral to test-paper, and insoluble in water. Mixed with air, it forms an explosive compound. When inflamed, it forms carbonic acid and water.

It is decomposed at a high temperature.

It is not a supporter of combustion or respiration.

When mixed in equal volumes with chlorine an oily liquid is formed—Dutch oil. If the chlorine be in excess and a light applied, there is an abundant deposit of carbon.

Preparation.—By heating in a flask one part by measure of alcohol, $C_4H_6O_2$, and three parts of sulphuric acid, SO_3,HO.

Decomposition:—

$$C_4H_6O_2 + SO_3,HO = SO_3,HO + 2HO + C_4H_4.$$

COAL-GAS.

Obtained by the distillation of coal in iron retorts. It is a mixture of light carburetted hydrogen, heavy carburetted hydrogen, carbonic oxide, hydrogen, and other volatile products. The clearness of the light is owing to the presence of solid particles of carbon raised to white heat.

SILICON.

Silicic Acid.

Symbol, SiO_3; *Equivalent*, 45·3.

Physical Properties.—A solid crystalline body, found in nature as quartz, agate, silex, flint, and sand. It is very refractory, scarcely yielding to the heat of the oxyhydrogen blowpipe.

Chemical Properties.—Composed of one of silicon to three of oxygen. It is insoluble in water and acids. Silicic acid unites with bases, forming silicates. It is decomposed by hydrofluoric acid.

Its uses in the arts are numerous.

Natural Silicates.—Felspar-clays.

Artificial Silicates.—Earthenware, glass, porcelain, &c.

HYDROGEN.

Symbol, H; *Equivalent*, 1.

Discovered by Cavendish in 1777, and called inflammable air.

Physical Properties.—Permanently gaseous, without colour, odour, or taste.

Density, 0·069.

Chemical Properties.—Simple body, slightly soluble in water. Has a great affinity for oxygen, and forms with that gas an explosive mixture, which detonates violently on the application of a light.

It is combustible, but not a supporter of either combustion or respiration. It does not combine directly with nitrogen to form ammonia.

Preparation. — 1st method. — Decomposition of cold water by zinc or iron in the presence of sulphuric acid:—

$$Zn + SO_3,HO = ZnO,SO_3 + H,$$
$$Fe + SO_3,HO = FeO,SO_3 + H;$$
$$\text{or, } (Zn + H,SO_4 = Zn,SO_4 + H,$$
$$Fe + H,SO_4 = Fe,SO_4 + H).$$

2nd method.—By passing steam through a gun barrel, loosely filled with iron turnings, and raised to a red heat—

$$4HO + 3Fe = Fe_3O_4 + 4H.$$

Uses.—Formerly employed to inflate balloons. The oxyhydrogen light of Drummond. The oxyhydrogen blowpipe.

PROTOXIDE OF HYDROGEN—WATER.

Symbol, HO ; *Equivalent*, 9.

Physical Properties.—A compound body, liquid at the common pressure and temperature. At 32° Fahr. it solidifies into ice. At 212° Fahr. it boils. It evaporates at all temperatures.

Water is without colour, but in large masses it appears of a blue colour. It is without taste or smell.

Chemical Properties.—Composed, by weight, of 1 of hydrogen to 8 of oxygen; by volume, 2 of hydrogen to 1 of oxygen. The composition of water was discovered by Cavendish and Nicholson in 1800. The exact quantitative composition of water was determined by Gay-Lussac and Humboldt.

Analysis of Water.—1st method. By the decomposition of steam when passed through a red-hot porcelain tube filled with copper turnings.

2nd method.—The decomposition of water by voltaic electricity.

Water is a solvent for most substances.

The purification of water is effected by distillation; such water is called distilled water, and is extensively used in the laboratory.

The uses of water are almost innumerable.

METALS.

POTASSIUM.

Symbol, K; *Equivalent*, 39.

This metal was discovered by Sir H. Davy, in 1807, by exposing the hydrate of potash, KO,HO (KHO_2), to the action of a powerful voltaic battery between two plates of platinum.

Potassium occurs in nature in combination with silicic acid, SiO_3, in granite, trap, basalt, and other igneous rocks. These rocks crumble down by the action of the atmosphere, and the potash finds its way into the soil. Soils destitute of potash are uniformly barren. It is worthy of remark here, that the leaves and young parts of plants contain the most potash, while the trunks of trees contain scarcely any.

Potassium in combination occurs in most natural waters, and in the sea as the bromide of potassium, KBr, the iodide of potassium, KI, and the chloride of potassium, KCl.

Potassium is prepared in the following manner:—A mixture of the carbonate of potash, KO,CO_2, and powdered charcoal is prepared, and thoroughly dried in a covered iron pot. When quite cold it is reduced to powder, and mixed with $\frac{1}{10}$ of its weight of charcoal, in small lumps. The mixture is then transferred to a well-hammered iron retort, or a bottle in which mercury is imported. The retort is placed in a furnace and the fire supplied with dry wood, so that the retort is enveloped in flame and kept at a uniform heat, approaching to whiteness. A copper receiver is divided by a partition, into one portion of which a tolerably large pipe from the retort dips into naphtha. The other side of the receiver is filled with ice. The fire is now raised in temperature, when decomposition of the carbonate of potash, KO,CO_2, by the charcoal takes place, carbonic oxide, CO, is abundantly disengaged, and large melted drops of potassium fall into the naphtha. Decomposition:—

$$KO,CO_2 + 2C = K + 3CO;$$
$$\text{or, } (KCO_3 + 2C = K + 3CO.)$$

If the potassium is required very pure, it is redistilled with naphtha in an iron retort.

Potassium has a bluish white colour. It can be cut with a knife; at 32°, it is crystalline; at 70°, like treacle; at 150°, it liquefies. It has a sp. gr. of 0·865.

Oxide of Potassium—Potash, KO.—When potassium is exposed to oxygen, it loses its metallic lustre, and becomes potash, KO, which encloses the metal with a white powder. When thrown upon water, HO, the powder is dissolved off, and the pure metal comes in contact with the water, HO, and owing to its powerful affinities for oxygen, the water is decomposed into its elements, H and O. The high temperature caused by this combination is sufficient to ignite the hydrogen. If potassium is heated in dry air we obtain potash, KO, but if the air be moist, then hydrate of potash, KO,HO, or (KHO_2).

The hydrate of potash, KO,HO (KHO_2), is a very important compound. It is always prepared by decomposing the carbonate of potash, KO,CO_2 (KCO_3), with the hydrate of lime, CaO,HO ($CaHO_2$), forming the carbonate of lime, CaO,CO_2 ($CaCO_3$), which is insoluble, the hydrate of potash KO,HO (KHO_2), remaining in solution.

Decomposition—

$$CaO,HO, + KO,CO_2 = KO,HO + CaO,CO_2;$$
$$\text{or, } (CaHO_2 + KCO_3 = KHO_2 + Ca,CO_3).$$

The liquid containing the hydrate is boiled down, when all the water except one equivalent is evaporated. This furnishes the solid hydrate, or caustic potash. It is raised in temperature till it melts, when it is run into cylindrical moulds about the size of lead pencils. It is kept in good stoppered bottles, owing to its powerful affinities for water, HO, and carbonic acid, CO_2.

The hydrate of potash is soluble in water, highly alkaline, and neutralizes most completely the most powerful acids. It dissolves the skin, and is used by medical men for that purpose. If a solution of it be shaken with any of the fixed oils, we have soap.

Carbonate of Potash, $KO,CO_2 + 2HO$ ($K,CO_3 + 2HO$), known as potash or pearlash. It is brought to this country from America and Russia, where it is obtained from the ashes of burned wood (the potash being derived from the soil). The ashes are put into large barrels of water. After some time, the water, saturated with soluble carbonate of potash, KO,CO_2, is drawn off and evaporated to dryness in iron cauldrons. The carbonate is very impure, containing, along with the carbonate, the silicate, sulphate, and chloride of potassium. The pure carbonate is obtained by calcining the tartrate of potash, digesting the mass with water, and evaporating the filtered liquid to dryness. Carbonic acid, CO_2, is too feeble an acid to completely neutralize the alkaline reaction of potash. The carbonate has therefore an alkaline taste and action on vegetable colours. It is used in medicine and the manufacture of glass and soap.

The *Bicarbonate of Potash*, $KO, CO_2 + HO,CO_2$ ($K,CO_3 + H,CO_3$), is obtained by passing carbonic acid, CO_2, through a solution of the carbonate of potash, KO,CO_2, when you obtain a white crystalline precipitate of the bicarbonate of potash, $KO,CO_2 + HO,CO_2$. It is less soluble than the carbonate.

Nitrate of Potash, KO,NO_5 (K,NO_6).—Nitre or saltpetre is brought to this country chiefly from the East Indies, where it appears, in some parts, like snow on the top of the ground. The nitre is collected, dissolved in water, to which a little carbonate of potash is added, to throw down the lime. The liquid is evaporated, and we obtain the nitre of commerce. It is used for the manufacture of nitric acid, HO,NO_5, and gunpowder. It has the property of preserving meat. In France, large quantities of nitrate of potash are manufactured in the following manner:—Heaps of old mortar and earth are collected and protected from the rain in a shed freely exposed to the air by having no sides, merely a roof to keep off the rain. These heaps are constantly turned and watered with putrid urine. When a good quantity of the salt has been obtained it is collected and mixed with water. Besides nitrate of potash, the solution always contains considerable quantities of the nitrates of lime and magnesia. To remove these a solution of pearlash (crude carbonate of potash) is added, which effects the decomposition of the salts alluded to:—

$$CaO,NO_5 + KO,CO_2 = CaO,CO_2 + KO,NO_5;$$

$$\text{or, } (Ca,NO_6 + K,CO_3 = Ca,CO_3 + K,NO_6).$$

The carbonates of lime and magnesia are insoluble. The filtered solution is then evaporated to dryness, and the salts purified by re-solution and crystallization. The nitrate of potash, KO,NO_5, has no action on vegetable colours: it crystallizes in six-sided prisms, and is very soluble in water. When thrown on the metals in a state of fusion, oxidation occurs at the expense of the oxygen of the nitric acid.

Gunpowder is a mixture of nitrate of potash, sulphur, and charcoal, in the following proportions:—

Nitrate of potash (KO,NO_5) .	75	parts by weight.
Carbon (C)	15	”
Sulphur (S)	10	”
	100	

The best charcoal for this purpose is manufactured from dogwood or elder. When this mixture explodes, the following rearrangement of the elements takes place: the oxygen of the nitrate of potash, KO,NO_5, goes to the carbon, forming carbonic acid, CO_2; the sulphur combines with the potassium, forming the sulphide of potassium, KS; and the nitrogen is set free. A volume of gas is formed occupying a space 1600 times greater than the powder, and this accounts for the violence of the explosion.

Sulphate of Potash, KO,SO_3 (K,SO_4).—This salt is formed as the residue of the process for the manufacture of nitric acid, from the nitrate of potash, KO,NO_5, (K,NO_6), in the presence of sulphuric acid, HO,SO_3.

Decomposition—

$$KO,NO_5 + HO,SO_3 = HO,NO_5 + KO,SO_3;$$

$$\text{or, } (K,NO_6 + H,SO_4 = H,NO_6 + K,SO_4).$$

The sulphate of potash is soluble in water, and is used in medicine. The bisulphate of potash, $KO,SO_3 + HO,SO_3$, or $(K,SO_4 + H,SO_4)$, is used in calico printing.

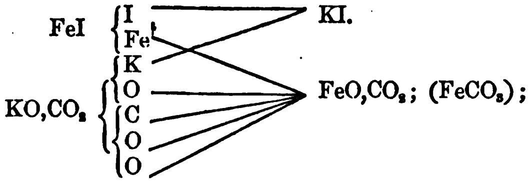
Chlorate of potash, KO,ClO_5; (K,ClO_6), is largely consumed for the manufacture of lucifer matches. It is prepared by passing a current of chlorine through a hot solution of caustic potash. Six equivalents of chlorine act on five equivalents of potash, KO, forming five equivalents of chloride of potassium, KCl, and one of chloric acid, ClO_5, which then unites with a sixth equivalent of potash, KO, forming one equivalent of chlorate of potash, KO,ClO_5:—

$$\begin{array}{l} Cl, \quad + KO \\ Cl,KO \\ Cl,KO \\ Cl,KO \\ Cl,KO \\ Cl,KO \\ \hline 5KCl + KO,ClO_5; \end{array}$$

$$\text{or, } (6\,KO + 6\,Cl = 5KCl + K,ClO_6).$$

When this salt is heated, oxygen is abundantly disengaged both from the acid and the base. The chlorate of potash deflagrates violently with combustible matter, explosion sometimes occurring by slight blows or touches. Sulphur and chlorate of potash explode with a loud report, when rubbed together. It cannot be used for gunpowder for these reasons.

The iodide of potassium, KI, is prepared by adding iron filings to iodine water until the combination is complete and the solution colourless. The resulting iodide of iron, FeI, is filtered and decomposed with a pure solution of carbonate of potash, KO,CO_2 (K,CO_3), when the following decomposition takes place:—



$$\text{or, } FeI + KO,CO_2 = KI + FeO,CO_2.$$

The iodide of potassium, KI, is separated by filtration and evaporation till the liquid is sufficiently concentrated to crystallize in cooling. It is very soluble in water, and crystallizes in cubes. The bromide of potassium, KBr, may be prepared in the same way, substituting bromine for iodine.

All the salts of potash, KO, are without colour when not associated with a coloured metallic oxide or a coloured acid. They are all more or less soluble in water.

Tests for Potash.

If tartaric acid be added to a potash solution, a white crystalline precipitate of cream of tartar is thrown down. Agitation and a few drops of hydrochloric acid, HCl, promote the precipitation. Some time generally elapses before the appearance of the precipitate. Bichloride of platinum, in moderately dilute solutions of potash salts (on the addition of a little hydrochloric acid, gives a crystalline yellow precipitate, which is a double salt of bichloride of platinum and potassium. This compound and the cream of tartar are soluble in 60 parts of cold water. A little alcohol increases the delicacy of both tests.

Salts of potash colour the outer blowpipe flame purple or violet. This is only seen to advantage, when the salts are pure. A little nitrate of potash thrown into a clear coal fire produces the same effect. A salt of potash, placed on a plate with a little alcohol and inflamed, produces a beautiful violet flame.

SODIUM.

Symbol, Na; *Equivalent*, 23.

This metal was obtained a few years after the discovery of potassium, by Davy, by a similar method to that described under that metal. It is usually prepared by decomposing the carbonate of soda, NaO,CO_2, (Na,CO_3), at a high temperature, with charcoal. Six parts of anhydrous soda are dissolved in a little hot water, and mixed with two parts of finely-powdered charcoal and one part of charcoal in lumps. The whole is then evaporated to dryness, transferred to an iron retort, and heated to whiteness. The tube of the retort passes into a receiver, containing rock-oil, and the whole operation is carried on the same as described under potassium. The process is easier than in the preparation of potassium, and more certain in its results. Sodium is a silver-white metal, soft at ordinary temperatures, melts at 195°, and oxidizes in air, passing into soda, NaO. Its sp. gr. is 0·972.

Sodium thrown upon water decomposes it with great energy, the temperature seldom rising high enough to inflame the liberated hydrogen. In hot water it takes fire at once, burning with a yellow flame, and giving rise to a solution of soda, NaO. Sodium resembles potassium in its properties.

Sodium occurs abundantly in nature, but not in its free state. Its most important native compound is the chloride of sodium, NaCl, which, as rock salt, occurs in immense beds in different parts of the earth, and also in sea water. Every natural water contains small quantities of this metal in combination. Small quantities of soda occur in most land plants, but largely in kelp and marine plants.

Soda, NaO, is obtained by exposing thin slices of the metal to dry air.

The *hydrate of soda*, NaO,HO, (Na,HO_2), is obtained by decomposing a dilute solution of carbonate of soda, NaO,CO_2 with hydrate of lime, CaO,HO, the same as described under the hydrate of potassium, KO,HO.

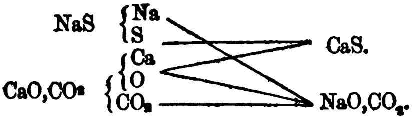
The solid hydrate is white, highly alkaline, and a powerful solvent for animal matter. It is used in large quantities for making soap.

Carbonate of soda, $NaO,CO_2 + 10HO$; or ($Na,CO_3 + 10HO$).—This is one of the most important compounds of soda, from its consumption in glassmaking, bleaching, washing, baking, and medicine. It is obtained from various sources. One is from the ash of a plant which is cultivated on the shores of the Levant, called the salsola soda. It is produced in several places on the coast of Spain. This is the barilla of commerce, and contains about 18 per cent. of the carbonate of soda, NaO,CO_2. That made in Brittany from sea weeds, and called varec, contains about 5 to 6 per cent.

The native *sesquicarbonate of soda*, $2NaO,3CO_2$, is found as a natural product on the margins of certain lakes in Egypt, Africa, Mexico, and Hungary. The natron lakes of Egypt have yielded soda from a very remote period. By far the greater part of the carbonate of soda, NaO,CO_2, is manufactured from the sulphate of soda, NaO,SO_3, (Na,

SO_4), by the following method:— About 600lbs. of chloride of sodium, NaCl, is placed upon the hearth of a well-heated reverberatory furnace, and an equal weight of sulphuric acid, HO,SO_3, sp. gr. 1·6, is poured upon it from an opening in the roof. The sulphuric acid, HO,SO_3, is thoroughly mixed with the salt; hydrochloric acid, HCl, is disengaged, which is either allowed to pass up the chimney or condensed. The salt by this process is converted into the sulphate of soda, NaO,SO_3, and hydrochloric acid, HCl.

This sulphate is reduced to powder and mixed with an equal weight of chalk, CaO,CO_2, ($CaCO_3$), and about half as much ground coal. Two cwt. of this mixture is thrown into a reverberatory furnace and heated to fusion, with constant stirring. When the decomposition is thought complete, the melted mass is raked into an iron trough, and allowed to cool. It is then broken up into little pieces, lixiviated with water, the solution evaporated to dryness, and the salt calcined with a little sawdust in a suitable furnace. This is the soda ash or British alkali of commerce, and, when good, contains about 50 per cent. of pure soda, NaO, partly as the carbonate of soda, NaO, CO_2, and partly as the hydrate of soda, NaO,HO, the remainder being the sulphate of soda, NaO,SO_3, ($NaSO_4$), chloride of sodium, NaCl, and traces of the sulphite and hyposulphite of soda. By dissolving this soda ash in hot water, and filtering the solution, the carbonate of soda, NaO,CO_2, is deposited in large crystals as the solution cools. The reaction which takes place in the calcination of the sulphate with coal dust is as follows:—The combustible matter of the coal converts the sulphate of soda, NaO,SO_3, into the sulphide of sodium, NaS; this sulphide of sodium, NaS, is then decomposed by the chalk, CaO,CO_2.



$$NaS + CaO,CO_2 = CaS + NaO,CO_2;$$

$$\text{or, } (NaS + Ca,CO_3 = CaS + Na,CO_3).$$

This process for the manufacture of the carbonate of soda, was originally proposed by Leblanc, and although other methods have been proposed, this is found the most advantageous.

The ordinary crystals of carbonate of soda, NaO,CO_2, contain 10 equivalents of water, but by careful management they are sometimes obtained with seven, nine, fifteen, and sometimes with only one equivalent of water. When heated, they fuse in their water of crystallization. The salt dissolves in two parts of cold water, and in less than its own weight of boiling water. The solution has a strong alkaline reaction and taste.

Bicarbonate of soda, $NaO,CO_2 + HO,CO_2$, $(Na,CO_3 + H,CO_3)$.—This salt is prepared by passing carbonic acid, CO_2, into a cold solution of the neutral carbonate, or by placing crystals of the carbonate of soda, NaO,CO_2, in an atmosphere of carbonic acid, CO_2. A solution of the bicarbonate of soda, NaO,CO_2+HO,CO_2, is feebly alkaline, and much milder than the carbonate.

The *sesquicarbonate of soda*, $2NaO,3CO_2 + 4HO$, occurs on the banks of the soda lakes of Sokena in Africa. It is exported to this country under the name of trona. This is a by-product in some chemical manufactures.

Sulphate of soda, *Glauber's Salt*, $NaO,SO_3 + 10HO$; or, $(Na,SO_4 + 10HO)$.—It may be prepared by adding dilute sulphuric acid, HO,SO_3, to a solution of the carbonate of soda, NaO,CO_2. This salt is slightly bitter and purgative. The mineral springs of Cheltenham contain this salt.

The *bisulphate*, $NaO,SO_3 + HO,SO_3 + 3HO$; or, $(Na, SO_4 + H,SO_4 + 3HO)$.—This salt is very soluble in water; has no acid reaction. When strongly heated it gives up an equivalent of sulphuric acid, and becomes the sulphate of soda, NaO,SO_3, (Na,SO_4).

The hyposulphite, NaO,S_2O_2 (or Na,S_2O_3), is largely used in photography. It is prepared by forming the neutral sulphite of soda by passing a stream of well-washed sulphurous acid, SO_2, into a strong solution of the carbonate of soda, NaO,CO_2; and digesting the solution with sulphur at a gentle heat for several days; then, by a careful evaporation at a moderate temperature, large and regular crystals of hyposulphite of soda, NaO,S_2O_2, are obtained, which are very soluble in water.

Nitrate of soda, NaO,NO_5; (Na,NO_6).—*Cubic Nitre* occurs native in enormous quantities in Peru, where it forms regular beds to a very great extent. The nitrate of soda, NaO,NO_5, is used for making nitric acid; but it cannot be used for gunpowder, as the salt burns too slowly and absorbs moisture.

The *tribasic phosphate* of soda, $2NaO,HO,PO_5 + 24HO$ $(Na_2H,PO_8 + 24HO)$, is prepared by precipitating the phosphate of lime with the carbonate of soda.

A second salt, sometimes called the sub-phosphate, has the formula $3NaO,PO_5 + 24HO$; or $(Na_3,PO_8 + 24 HO)$. This is obtained by adding caustic soda to the preceding salt.

The *tribasic phosphate of soda and ammonia*, or *Microcosmic Salt*, $NaO,NH_4O,HO,PO_5 + 8HO$; or, $(Na,NH_4, H,PO_8 + 8HO)$. This salt is prepared by dissolving six parts of the common phosphate of soda in two of water, when the whole is liquefied. One part of chloride of ammonium, NH_4Cl, is added; chloride of sodium, $NaCl$, separates, and is removed by the filter; and from the properly concentrated solution, the microcosmic salt crystallizes out.

Bibasic phosphate of soda, $2NaO,PO_5 + 10HO$; or $(Na_2PO_7 + 10HO)$.

Monobasic phosphate of soda, NaO,PO_5 $(NaPO_6)$.

All the tribasic phosphates give a bright yellow precipitate with nitrate of silver, AgO,NO_5; the bibasic and monobasic give white precipitates with the same salt.

Biborate of Soda, *Borax*, $NaO,2BO_3 + 10HO$.—This salt occurs in the waters of certain lakes in Persia. The impure salt is imported to this country from the East Indies, under the name of tincal. A great quantity of this salt is now prepared from the native boracic acid, BO_3, of Tuscany, where it occurs, uncombined in the hot springs of that country. Although this is an acid salt, it has an alkaline reaction to test-paper. It is used in the arts for soldering metals; its action consists in rendering the surfaces of the metals to be joined metallic, by dissolving off the oxide. It sometimes enters into the composition of the glaze with which stoneware is covered. The refined borax is used by the washerwomen of Holland and Belgium, to whiten linen and to soften water. About a handful is mixed with ten gallons of boiling water. The saving of soap is nearly one half, and the borax does not injure the linen.

Common Salt, *chloride of sodium*, $NaCl$.—A large quantity of salt is obtained from the brine springs of Droitwich; but our chief supply is from the mines of Cheshire. Rock salt is always too impure for use; and if no natural brine-spring exists, one is formed by sinking a shaft into the rock salt, and introducing water. This water, when saturated, is pumped up, and evaporated in large iron pans. As the salt separates, it is removed by means of a wooden scoop, and pressed, while moist, into moulds, and then transferred to the drying stove. When large crystals are required, as for bay salt, the evaporation is conducted very slowly. Common salt generally contains a little chloride of magnesium, $MgCl$. The *iodide* of sodium, NaI, and the *bromide* of sodium, $NaBr$, are similar to the corresponding potassium compounds.

Tungstate of Soda, $NaO,Wo_2,2HO$, or $NaWo_3,2HO$. Every year considerable loss of life occurs from the inflammable nature of materials used for dress. Solutions of several salts have been proposed with a view of rendering fabrics non-inflammable. From numerous experiments, it appears that a solution of the tungstate of soda is greatly to be preferred. A concentrated neutral solution of the salt is diluted with about one-third of water, and then mixed with three per cent. of phosphate of soda.

This solution is found to keep well, and is used in the Royal laundries. The lightest muslin washed in this solution and dried, becomes non-inflammable.

Analysis of Droitwich Brine.

	Grains.
Chloride of sodium	21761·872
Chloride of magnesium . . .	2·56
Sulphate of lime	91·12
Sulphate of alumina	14·4
Sulphate of soda	342·72
Iodide of sodium	·208
Total *salts* to an imperial gallon . .	22212·88

Tests.

All the soda salts are very soluble in watrr, except the antimoniate. The presence of soda is generally determined by purely negative evidence in the following manner. To the solution of the unknown substance caustic potash is added, and heat applied: no odour of ammonia indicates the absence of that body. The salt then must contain either potassium or sodium. A solution of bichloride of platinum or tartaric acid is added. If no precipitate is produced, the salt contains sodium. This may be confirmed by adding solution of antimoniate of potash, which produces a white precipitate.

In the blowpipe flame sodium compounds burn with a yellow colour.

AMMONIUM.

Symbol, NH_4; *Equivalent*, 18.

In connection with potassium and sodium, ammonium may be conveniently studied.

Ammonia, NH_4O, is the oxide of a hypothetical body called ammonium. This oxide however, when liberated from any combination, always splits up into gaseous ammonia, NH_3, and water, HO:—$NH_4O = NH_3 + HO$. The salts of ammonia correspond exactly with those of potash and soda. All attempts to isolate ammonium have failed, from its tendency to separate into ammonia, NH_3, and hydrogen. Ammonium plays the part of a metal in the composition of salts; it is capable of replacing sodium and potassium. The sal-ammoniacum of the Romans was the hydrochlorate or muriate of ammonia, and was obtained as a natural product from a district in Libya. Ammonia was called by the alchemists spirits of hartshorn, because it was obtained by the distillation of deers' horns, in close vessels. It is sometimes called the volatile alkali.

Although ammonium has never been isolated, it is possible that it is not behind potassium and sodium in metallic lustre, and the ordinary physical characters of a metal. An amalgam of NH_4 with mercury may be prepared.

If a small quantity of mercury be heated in a test-tube with a grain or two of potassium or sodium, and a gentle heat applied, violent combination ensues, attended with the production of light and heat. When cold, the fluid amalgam is put into a saucer, and covered with a strong solution of the hydrochlorate of ammonia. The mercury increases enormously in volume, and a pasty ammoniacal amalgam is formed. Left to itself, it quickly decomposes into mercury, ammonia, and water.

The great resemblance the ammoniacal salts bear to those of potash and soda, is the chief argument in favour of the existence of ammonium.

In the manufacture of coal gas, much ammonia is formed by the combination of the nitrogen and hydrogen of the coal. In the manufacture of bone charcoal ammonia is also produced; but in both cases it is very impure.

If sal-ammoniac be reduced to powder, mixed with quick-lime, and heated in a retort, gaseous ammonia is abundantly evolved; and as water dissolves about 500 volumes of the gas, it must be collected over mercury, or by displacement. The vessel in which the gas is collected must be held above the orifice from which it is issuing, the same as if it were hydrogen. The following decomposition takes place:—

$$NH_4Cl + CaO = CaCl + HO + NH_3.$$

Ammonia compounds are now largely manufactured from gas liquor, and the condensed products obtained from the distillation of bones and other animal refuse, as mentioned under bone charcoal. It was formerly obtained in Egypt from the soot of camel's dung.

The alkaline action of ammonia may be tried with turmeric, litmus paper, or purple cabbage water.

When ammonia comes in contact with one of the volatile acids, dense white fumes are formed. A test-tube being rinsed out with hydrochloric acid, HCl, and held over a vessel from which ammonia is issuing, white fumes and solid crystals of sal-ammoniac are formed by the union of the ammonia with the hydrochloric acid.

If gaseous ammonia, or ammonia water, be added to a solution of sulphate of copper, a splendid azure blue compound is formed. This colour is formed by a bluish precipitate, which is redissolved by excess of ammonia, and the blue colour produced afterwards.

Ammonia, when uncombined, may be readily recognized by any of the characters above described.

Chloride of Ammonium, Sal-ammoniac, NH_4,Cl, is prepared from gas- or bone-liquor. The solution is treated with a slight excess of hydrochloric acid, HCl, by which the ammonia is neutralized, and the carbonate of ammonia, NH_4O,CO_2, and sulphide of ammonium, NH_4S, are decomposed, with evolution of carbonic acid, CO_2, and sulphuretted hydrogen, HS. The liquid is evaporated to dryness. and the salt carefully heated to expel the tarry matter with with it is associated. It is then purified by sublimation in iron vessels lined with clay and covered with domes of lead. This salt is sublimed without decomposition. Chloride of ammonium, NH_4Cl, forms double salts, with the chloride of magnesium, $MgCl$, chloride of nickel, $NiCl$, chloride of cobalt, $CoCl$, chloride of zinc, $ZnCl$, chloride of copper, $CuCl$, chloride of manganese, $MnCl$, and bichloride of platinum, $PtCl_2$.

Sulphate of Ammonia, $NH_4O,SO_3 + HO$; ($NH_4,SO_4 + HO$).—This salt may be prepared by neutralizing the carbonate of ammonia with sulphuric acid; but it is largely manufactured for agricultural and commercial purposes by neutralizing gas liquor with sulphuric acid, evaporating the solution, and purifying the product, which is impure sulphate of ammonia, $NH_4O,SO_3 + HO$. It is soluble in two parts of cold water, and crystallizes in long six-sided prisms, which lose an equivalent of water when heated. The sulphate of ammonia is entirely decomposed by ignition, and by long boiling with water the ammonia is volatilized, and the liquid rendered acid.

Carbonate of Ammonia, NH_4O,CO_2.—The carbonates of ammonia are very numerous. The neutral anhydrous carbonate (which is, however, not a true salt), NH_3,CO_2, is prepared by the direct union of well dried and cooled carbonic acid, CO_2, and ammoniacal gas. These gases combine in the proportion of one volume of carbonic acid, CO_2, with two volumes of ammonia, NH_3. The compound is very volatile, has a pungent odour, is soluble in water, and condenses in white flakes. The carbonate of ammonia used in medicine is prepared by subliming a mixture of chloride of ammonium, NH_4Cl, and carbonate of lime, CaO,CO_2. This compound contains a less quantity of base than the neutral carbonate. Its composition varies, but in freshly-prepared specimens it approaches the sesquicarbonate, $2(NH_4O),3CO_2$.

Sesquicarbonate of ammonia $2(NH_4O),3CO_2$.—This salt is very unstable; it is decomposed by a slight increase of temperature. By exposure to air, it loses its odour, and becomes a soft, white powder, which is the bicarbonate of ammonia.

Bicarbonate of ammonia, $NH_4O,CO_2 + HO,CO_2$; $(NH_4,CO_3 + H,CO_3)$.—This salt, although volatile, is more permanent than the sesquicarbonate. The bicarbonate of ammonia, $NH_4O,CO_2 + HO,CO_2$, resembles, in every respect, the bicarbonate of potash, $KO,CO_2 + HO,CO_2$.

Nitrate of ammonia, NH_4O,NO_5, (NH_4,NO_6), is prepared by adding dilute nitric acid, HO,NO_5, to the carbonate of ammonia, NH_4O,CO_2, until the solution is neutralized, and slowly evaporating at a moderate temperature. The salt crystallizes out in six-sided prisms, like the nitrate of potash, KO,NO_5, and deflagrates like that salt. When heated it is decomposed into nitrous oxide and water.

The acetate and citrate of ammonia are prepared by neutralizing the carbonate with acetic or citric acid.

When ammonia, or its carbonate, is neutralized by hydrochloric acid, HCl, hydrosulphuric acid, HS, hydriodic acid, HI, and hydrobromic acid, HBr, the resulting salts are anhydrous, and contain ammonium, NH_4, united to the radical of the respective acids:—NH_4,Cl; NH_4,S; NH_4,I; NH_4,Br.

The sulphide of ammonium may be prepared by distilling the chloride of ammonium, NH_4Cl, with the corresponding sulphide of potassium or sodium. The bisulphide of ammonium is a compound of great practical utility, $NH_4S + HS$. It is obtained by saturating a solution of ammonia with well-washed sulphuretted hydrogen until the solution is saturated. By exposure to light, it becomes slightly yellow. It gives precipitates with most metallic solutions, several of which are very characteristic.

Tests.

All the salts of ammonia are more or less soluble. The acid tartrate, and the double chloride of ammonium and platinum being the least so. Ammonia, therefore, gives with the bichloride of platinum, $PtCl_2$, and with tartaric acid, $2HO,T$, the same reactions as potash. The precipitates are similar in composition.

Ammonia may be recognized with certainty in the presence of all other bases by its being set free in a gaseous state, by the action of any of the caustic alkalies or alkaline earths (the hydrate of lime, CaO,HO, is the best for this purpose); and by its odour, and the white fumes produced when a glass rod moistened with hydrochloric acid, HCl (not fuming acid), is brought near the compound acted on as above. There is no blowpipe test for ammonia.

BARIUM.

Symbol, Ba; *Equivalent*, 68.

Barium was obtained by Sir H. Davy: it is a white, silver-looking metal, malleable, melts below a red heat, and slowly oxidizes in air. It may be prepared by heating baryta, BaO, in a porcelain tube, through which the vapour of potassium is passed. The barium is then extracted by mercury, and the amalgam distilled in a small Bohemian glass retort.

There are two oxides of barium,—the protoxide, BaO, and the binoxide, BaO_2.

Baryta, or barytes, BaO.—The carbonate and sulphate of baryta occur in considerable quantities in some lead mines. From these salts baryta is obtained. If pure baryta is required, the crystallized nitrate of baryta is decomposed by heat in a porcelain crucible until red vapours of NO_4 cease to be disengaged. Baryta remains in the crucible as a grey spongy mass. If this residue be mixed with water, it forms a hydrate of baryta, BaO,HO; ($BaHO_2$). The hydrate of baryta has a great affinity for carbonic acid, and is soluble in water. A hot saturated solution deposits, on cooling, crystals having the following composition:—BaO,HO + 9HO.

The hydrate of baryta is highly alkaline; and a valuable reagent. The smallest trace of carbonic acid instantly renders the solution turbid.

The binoxide, BaO_2, is prepared by raising to a red heat baryta, BaO, in a porcelain tube, through which a current of oxygen is passed.

Chloride of barium, BaCl + HO.—If the carbonate of baryta be dissolved in hydrochloric acid, and the solution filtered and slowly evaporated, crystals of the chloride of barium are deposited in the water of crystallization.

Sulphate of baryta, BaO,SO_3 (Ba,SO_4), called heavy spar, is often found beautifully crystallized in lead mines. It may be produced when any soluble sulphate is added to a solution of any baryta salt. It is insoluble in water, or in acids. Boiling nitric acid dissolves a little, but the salt separates on cooling. Sulphate of baryta is used for painting, and adulterating white lead.

The *nitrate of baryta,* BaO,NO_5 (Ba,NO_6), is prepared in the same way as the chloride, substituting nitric acid for hydrochloric, but the crystals are anhydrous. This salt is more soluble in pure water than in dilute nitric acid. A precipitate of the nitrate of baryta may be distinguished from the sulphate by its disappearance on the application of heat, or the addition of a large quantity of water.

Nitrate of baryta is used in the manufacture of green fire :—

Nitrate of baryta	450	grains
Sulphur	150	„
Chlorate of potass	100	„
Lamp-black	25	„

The same precautions to be observed as are given under "red fire," p. 117.

Carbonate of baryta, Witherite, BaO,CO_2 (Ba,CO_3).—This, like the sulphate, is found in some lead mines. It may be prepared by dissolving the chloride of barium, $BaCl$, in water, and precipitating with the carbonate of soda or any other alkaline carbonate.

The hydrate of baryta is used in the laboratory to separate carbonic acid from gaseous mixtures; the nitrate and chloride to precipitate sulphuric acid.

Tests.

Sulphuric acid and the soluble sulphates precipitate immediately the sulphate of baryta, BaO,SO_3; insoluble in acids and alkalies.

Chromate of potash precipitates in neutral and alkaline solutions a pale yellow precipitate of the chromate of baryta; insoluble in acetic acid and alkalies.

Oxalic acid precipitates in concentrated solutions the white oxalate of baryta; soluble in acids. A few drops of ammonia promote the formation of the precipitate.

STRONTIUM.

Symbol, Sr; *Equivalent*, 43·8.

This metal may be prepared by means like those described for barium, and is analogous to that metal in its physical and chemical properties. All the soluble salts of baryta are poisonous; this is not the case with strontia salts. It is found as a carbonate and sulphate in lead mines.

Protoxide of strontium, strontia, SrO.—This oxide may be prepared by decomposing the nitrate with heat, as under baryta. The hydrate has a great affinity for carbonic acid.

The *binoxide of strontia*, SrO_2, may be prepared like the binoxide of barium.

All the salts of strontia may be prepared by methods similar to those described under the salts of baryta.

The *chloride of strontium*, SrCl, is soluble in two parts of cold water. It is also soluble in alcohol, and when kindled, burns with a beautiful crimson flame.

The *nitrate of strontia*, SrO,NO_5 (Sr,NO_6), is used for the manufacture of red fire:—

Nitrate of strontia (dry)	800	grains
Sulphur	225	„
Chlorate of potass	200	„
Lamp-black	50	„

The strontia, sulphur, and lamp-black must be well powdered and well mingled together, after which the chlorate of potass may be added in rather a coarse powder, and mixed with the other ingredients, without much friction. The red fire has sometimes ignited spontaneously.

Tests.

The soluble sulphates and sulphuric acid precipitate completely from its solutions the sulphate of strontia; the precipitate does not appear for some time, unless the solution is concentrated. The precipitate is insoluble in acids and alkalies.

Chromate of potash produces no precipitate, the chromate of strontia being soluble.

Oxalic acid precipitates the white oxalate of strontia in very dilute solutions; a little ammonia promotes the formation of the precipitate.

Salts of strontia, heated on a platinum wire in the inner flame of the blowpipe, give a crimson colour.

CALCIUM.

Symbol, Ca; *Equivalent*, 20.

This metal is obtained with difficulty. It has been seen by very few persons. It is described as a silver white metal, and has a great affinity for oxygen, passing into lime, CaO. The term calcareous is derived from the Latin word *calx*.

Calcium may be obtained by electrolysing in contact with mercury the hydrate of lime; the amalgam is then decomposed by distillation. It may also be obtained by strongly heating lime in an iron tube through which the vapour of potassium is passed. The reduced calcium is extracted by mercury, and the mixture distilled in a small green glass retort.

The *protoxide of calcium*, or *quicklime*, is obtained by heating the carbonate of lime, CaO,CO_2. The black limestone of Kilkenny and Derbyshire is the best for this purpose. If lime be required absolutely pure, the nitrate of lime is precipitated with the carbonate of ammonia, and ignited to whiteness in a platinum crucible. Lime for ordinary purposes is obtained by calcining chalk in a kiln; a red heat continued for some hours is sufficient to liberate the whole of the carbonic acid. In the best lime kilns the process is carried on continuously by broken limestone and fuel being thrown in from the top and the lime raked out at intervals at the bottom. When the limestone contains silica, and the temperature is very great, the lime refuses to slake; silicate of lime is formed, and the lime is said to have been overburned. Pure lime is quite infusible; at a high temperature it emits a pale light. When sprinkled with water it slakes and crumbles into a soft white powder, which is the hydrate of lime, CaO,HO ($CaHO_2$). When quicklime is mixed with water the temperature is raised sufficiently to explode gunpowder and char wood; carts laden with lime have sometimes been set on fire by a shower of rain. The equivalent of water may be expelled by raising it to a red heat, the hydrate returning to its original protoxide of calcium. This hydrate of lime is soluble in water; and the colder the water, which is an exception to the general rule, the more lime is taken up. A pint of water at 60° will dissolve about eleven grains; at 212°, about six.

Lime-water may be prepared by agitating with water in a stoppered bottle the hydrate of lime: when the excess of lime falls to the bottom, the water may be poured off and a fresh quantity added. If lime-water be exposed to the atmosphere in an open vessel, it becomes covered with a thin film of the carbonate, which is due to the absorption of carbonic acid from the atmosphere. The hardening of mortar is partly owing to the gradual absorption of carbonic acid. Some cements, which are used for cementing the piers of bridges and other structures under water, contain alumina, silica, and oxide of iron. The cement has to be prepared with great care. When these cements are ground to powder and mixed with water, solidification ensues, and in this state they effectually resist the action of water. It is thought that a chemical combination takes place between the silica, oxide of iron, and alumina. In ordinary mortar no such combination takes place between the silica (sand) and the lime.

The beneficial use of lime for agricultural purposes is a matter of dispute; it is however, useful in causing the rapid decomposition of organic matter in peaty soils. These soils are barren from the excess of vegetable matter. In clayey soils it acts beneficially by liberating the potass, which exists as an insoluble silicate in combination with alumina.

The *binoxide of calcium*, CaO_2, may be prepared by exposing lime to a red heat in a porcelain tube, and a current of oxygen passed through the tube.

Carbonate of lime, CaO,CO_2 (Ca,CO_3).—This is one of the most widely distributed compounds in the mineral world, forming whole mountains of immense thickness in almost every part of the world. It is the chief constituent of chalk, shells, and coral reefs. In nature the carbonate of lime is seldom pure; it generally contains clay, iron, and organic matter. But the most ancient of the crystalline limestones are destitute of organic remains, while the chalk is almost entirely made up of the exuviæ of once living beings. Limestone caverns are often lined with magnificent crystals of calcareous spar, which have been slowly deposited from a watery solution. These stalactite incrustations are often of great beauty. In pure water the carbonate of lime is scarcely soluble, but when the water contains free carbonic acid it is taken up. If lime-water be added to a vessel containing carbonic acid, a milky-white precipitate is formed; but on agitation, if the carbonic acid be in excess, the turbidity disappears, and a clear, transparent solution of carbonate of lime is obtained. If the clear solution is boiled, the carbonic acid is expelled, and the carbonate of lime again precipitated. All natural waters contain carbonic acid. Lime in the condition of a carbonate is almost always present in river and spring water. In limestone districts it is often present to a very great extent. The hardness of water is usually owing to the presence of carbonate of lime. This hardness may be considerably removed by boiling; it is, however, most completely removed by the addition of lime-water, when the carbonate of lime is precipitated. It is upon this principle that Professor Clark's process for softening water is based. The hardness of water containing carbonate of lime is called temporary, because it may be removed in a great degree by the methods mentioned. The use of such water for domestic purposes is attended with a great loss of soap; soap cannot act as a detergent until the whole of the lime is precipitated. Tea-kettles and boilers in which such water is heated become lined with a thick incrustation. Dr. Ritterband has proposed a method of removing this deposit, which consists of throwing into the boiler a small quantity of sal ammoniac, when carbonate of ammonia is formed, which is volatilized by the high temperature, chloride of calcium remaining in solution. This method is not applicable to permanently hard waters, which generally contain sulphate of lime, CaO,SO_3. All the other salts of lime may be obtained from the carbonate by dissolving it in the various acids.

Sulphate of lime, gypsum, alabaster, CaO,SO_3 (Ca,SO_4).—This salt is occasionally found, associated with rock salt, in an anhydrous and crystalline state. When regularly crystallized it is called selenite. If the chloride of calcium in solution be mixed with sulphuric acid, the sulphate of lime is precipitated; it is soluble in about 500 parts of cold water. The sulphate of lime is more soluble in water containing the nitrate of potash. Gypsum, or the hydrated sulphate, is largely employed in the arts for moulds, statues, and casts. The native hydrated sulphate, $CaO,SO_3 + 2HO$, if exposed to a temperature of 250° F. in an oven, is deprived of its water of crystallization, and afterwards reduced to powder. When mixed with water it solidifies into a hard hydrate; but this solidification does not take place if the gypsum has been overheated. It is called by workmen plaster of Paris. Scagliola work, or artificial-coloured marble, is often prepared by inserting pieces of natural marble in the sulphate when the plaster is soft, and polishing the surface when the plaster has hardened. It is not possible to burn plaster when it has been once used, so as to make it set again. Sulphate of lime is a common impurity of spring water. When water contains this salt in solution it is said to be permanently hard, since this hardness cannot be removed.

Phosphates of Lime.—There are several compounds of lime with phosphoric acid, two of these are tribasic—

$$2CaO,HO,PO_5;\ \text{or},\ (Ca_2H,PO_8).$$
$$3CaO,PO_5;\ \text{or},\ (Ca_3,PO_8).$$

These two salts are produced when the corresponding soda salts are added to a solution of the chloride of calcium. Bone earth consists of a combination of these salts. Bibasic and monobasic phosphates of lime exist. These are insoluble in water, but soluble in dilute acids.

Chloride of lime, CaO,Cl. — This salt is extensively manufactured as *bleaching powder*. If moistened hydrate of lime be exposed to chlorine gas this compound is produced. The chlorine is at first supplied slowly, so as to prevent any great increase of temperature. It is soluble in 10 parts of water. The solution is highly alkaline and its bleaching properties are weak. When the hydrate of lime is suspended in cold water, and chlorine passed through the solution, the lime is gradually dissolved and the same bleaching compound is also produced. The hydrates and carbonates of the alkalies may, by similar methods, be made to absorb chlorine and form corresponding compounds, which are sold as disinfecting fluids.

The exact constitution of the bleaching compounds of lime is involved in some obscurity. It is thought that these salts contain hypochlorous acid, a body as remarkable for its bleaching properties as chlorine.

If good bleaching powder be distilled with dilute sulphuric or nitric acid, but not in sufficient quantity to entirely decompose the salt, hypochlorous acid is obtained, sulphate or nitrate of lime being formed simultaneously with the chloride of calcium. The following equation will illustrate the decomposition—

$$(2CaO,Cl) + H,SO_4 = Ca,SO_4 + CaCl + ClO + HO.$$

This has led some chemists to view bleaching powder as a combination of chloride of calcium and hypochlorite of lime:—

$$2(CaO,Cl) = CaCl + CaO,ClO.$$

If these salts be exposed to heat, their bleaching properties are entirely destroyed. If exposed in an open vessel they absorb carbonic acid and evolve chlorine.

Bleaching powder thus varies in value with its age, the care bestowed in its preparation, and its subsequent exclusion from the atmosphere.

The general outline of the method followed in bleaching is as follows:—The goods are soaked in a dilute solution of chloride of lime; they are then removed to a vat containing dilute sulphuric acid. The bleaching powder now undergoes decomposition by the acid.

$$CaO,Cl + H,SO_4 = Ca,SO_4 + HO + Cl;$$

$$\text{or, } CaCl + CaO,Cl + 2(H,SO_4) = 2(CaSO_4) + 2HO + 2Cl.$$

The free chlorine disengaged in contact with the article to be bleached causes a complete destruction of the colouring matter. As the solutions of lime and sulphuric acid are very dilute, this process has to be repeated several times. If a pattern be stamped upon a coloured cloth with tartaric acid thickened with gum-water, and the cloth afterwards immersed in the lime-bath, no alteration of colour will take place except in those parts stamped with the acid. The acid acts, in this case, precisely the same as in the dilute sulphuric acid bath. The chlorine is disengaged, and the colour completely removed, in those parts stamped with the acid. We then obtain white patterns upon a coloured ground. If chloride of lime be mixed with water and placed in a saucer or basin, and a little acid added to the mixture, free chlorine is liberated; or cloths may be dipped in the solution and hung in rooms. One of these methods is usually adopted when chlorine is used as a disinfectant.

Chloride of Calcium, CaCl.—This salt is prepared by dissolving marble in hydrochloric acid. It is also a product in many chemical operations, and is used for drying gases. When mixed with snow or powdered ice it produces artificial cold. It is also used to determine the amount of hydrogen in organic bodies.

Sulphide of calcium, CaS, is obtained by reducing the sulphate of lime at a very high temperature with charcoal. This salt is slightly soluble in water. It is produced in the manufacture of carbonate of soda.

Fluoride of calcium, or *fluor-spar*, CaF, is the source from which hydrofluoric acid is prepared, and the other fluorides. It is found in beautiful cubic crystals, of various colours, in lead mines. It is decomposed by sulphuric acid.

Phosphide of calcium, CaP, is a brown-looking substance; it yields inflammable phosphuretted hydrogen when thrown on water. It is prepared by passing the vapour of phosphorus over fragments of lime heated to redness in a porcelain tube.

Tests.

Sulphuric acid and the alkaline sulphates give no precipitate, except in concentrated solutions of lime, and then only a partial precipitate, which completely dissolves in excess of water.

Chromate of potash gives no precipitate.

Oxalic acid throws down in dilute neutral solutions of lime the white oxalate of lime; a little ammonia renders the reaction more delicate.

MAGNESIUM.

Symbol, Mg; *Equivalent*, 12.

Magnesium is a white malleable metal, resembling zinc. It has no use in the arts, and is prepared by heating the chloride of magnesium with sodium in a test tube of German glass. Combination takes place with great energy and elevation of temperature, the chloride of sodium being formed, and metallic magnesium set free. When the tube is cold it is broken up and the contents put into cold water, by which the metal is separated from the salt.

Magnesia, MgO, is the only oxide of this metal. It may be obtained pure by raising to a red heat, in a porcelain crucible, the magnesia of the shops. It is a soft white powder, attracts moisture and carbonic acid from the atmosphere; unites slowly with water, forming a hydrate MgO,HO, or (Mg,HO_2); it is slightly soluble in water. Its alkaline reaction is best observed by moistening a little, and placing it upon litmus paper. The most powerful acids are completely neutralized by magnesia.

Carbonate of magnesia; MgO,CO_2, or ($MgCO_3$), occurs native, very much resembling calcareous spar.

If magnesia, MgO, be dissolved in water saturated with carbonic acid, and the solution left to evaporate, small prismatic crystals of the carbonate of magnesia are deposited with three equivalents of water. The carbonate sold in the shops is usually prepared by the addition of a hot solution of the carbonate of potass or soda to any soluble salt of magnesia. The precipitate, although usually called, and sold as, the carbonate, is not so. The salt is not constant in its composition. It usually has the following formula:—

$$4(MgO,CO_2) + MgO,HO + 6HO;$$

Sulphate of Magnesia, Epsom Salts, $MgO, SO_3 + 7HO$; ($Mg,SO_4 + 7HO$).—This salt is now largely manufactured by the action of sulphuric acid on magnesian limestone. It occurs in sea water and many mineral springs, is very soluble in water, has a disagreeably bitter taste, and purgative action. When heated, six equivalents of water pass off, the other equivalent being retained. Sulphate of magnesia forms double salts with the sulphates of potash and ammonia.

Phosphate of Magnesia, $2MgO,HO,PO_5 + 14HO$; or, ($Mg_2H,PO_8 + 14HO$).—If a solution of the sulphate of magnesia and the phosphate of soda are mixed together, and allowed to remain some time, small prismatic crystals of the phosphate of magnesia separate. This salt exists in the grain of wheat, barley, and oats, and is found in considerable quantities in beer.

Phosphate of Magnesia and Ammonia.—$2MgO,NH_4O,PO_5 + 12HO$; or, $(Mg_2NH_4,PO_8 + 12HO)$.—When any soluble phosphate is added to the sulphate of magnesia, a white precipitate slowly separates. If ammonia be added with the phosphate, and the solution concentrated, a crystalline precipitate appears at once; if the solution be dilute, the precipitate only appears after some time. This precipitate is the phosphate of magnesia and ammonia, $2MgO,NH_4O,PO_5 + 12HO$. This salt is slightly soluble in pure water. When heated, it is resolved into the bibasic phosphate of magnesia, $2MgO,PO_5$, or $(Mg_2 PO_7)$. At a red heat, it fuses into a white enamel-looking mass, and is present in some forms of urinary calculus.

Silicates of Magnesia, MgO,SiO_2. *Steatite*, or *Soapstone*, chiefly found in Cornwall. *Meerschaum*, $MgO,SiO_2 + HO$, used for making pipe bowls. Chrysolite, $3MgO, SiO_2$; augite, hornblende, and jade, are composed of variable quantities of magnesia, silicic acid, alumina, and iron.

Chloride of Magnesium, Mg,Cl.—This salt is obtained with some difficulty, owing to the obstinacy with which the last portions of water are retained. When magnesia is dissolved in hydrochloric acid, water and the chloride of magnesium are formed—

$$MgO + HCl = MgCl + HO.$$

This salt is very soluble, and the last portions of water are not evaporated from the solution without decomposition, the magnesia uniting with the oxygen and the hydrogen with the chlorine.

Tests.

The salts of magnesia are colourless.

The caustic alkalies and ammonia give a white gelatinous precipitate with magnesium salts, insoluble in excess, but soluble in chloride of ammonium.

Carbonates of soda and potash give a white precipitate. Carbonate of ammonia gives no precipitate in the cold.

Soluble phosphates give a white crystalline precipitate on the addition of a little ammonia.

MANGANESE.

Symbol, Mn; *Equivalent*, 27·6.

The metal manganese, or more properly the carbide, is prepared by heating the oxide with charcoal and borax. The process is rather difficult.

Manganese is only known as a chemical curiosity; it is put to no use in the arts. The metal is described as resembling cast iron, hard, brittle, and destitute of magnetic properties.

There are seven oxides of manganese, but two of these oxides are probably secondary compounds—

MnO, the protoxide.
Mn_2O_3, sesquioxide,
MnO_2, binoxide,
Mn_3O_4, red oxide $= MnO,Mn_2O_3$,
Mn_4O_7, varvicite $= Mn_2O_3,2MnO_2$,
MnO_3, manganic acid,
Mn_2O_7, permanganic acid.

The *protoxide*, MnO, may be obtained in its anhydrous state by heating the carbonate of manganese in an atmosphere of hydrogen or steam; the carbonic acid is disengaged, and a green-coloured powder left behind, which is the anhydrous protoxide of manganese. It is best known in the form of hydrate, and may be easily prepared by precipitating a salt of manganese with an alkali. The precipitate first appears white, then buff, then brown, by rapidly passing into higher states of oxidation. The protoxide is a powerful base, neutralizing most completely the strongest acids.

Sesquioxide, Mn_2O_3.—This compound in union with water is found crystallized in some parts of Germany. If the hydrated protoxide be exposed to the air, it passes into the sesquioxide. When oxygen is prepared by heating the binoxide in an iron retort, the residue is chiefly the sesquioxide, and is either of a black or brown colour, according to the circumstances under which it is produced. It is feebly basic. When heated with dilute sulphuric acid it dissolves; and on the addition of sulphate of potash deposits eight-sided crystals, having a similar constitution to common alum:—

$$Mn_2O_3,3SO_3 + KO,SO_3 + 24HO;$$
$$\text{or, } (2Mn,3SO_4 + K,SO_4 + 24HO).$$

Binoxide, MnO_2.—This is the most common native compound of manganese. It is found both amorphous and crystallized. It may be prepared artificially by adding a solution of bleaching powder to a salt of the protoxide of manganese. It is not obtained anhydrous by this method.

The binoxide is insoluble in water, of a black colour, and does not unite with acids to form salts. It is decomposed by hot hydrochloric and sulphuric acids, as in the preparation of chlorine and oxygen. This substance is largely used in the manufacture of bleaching powder.

Red oxide, Mn_3O_4, is found native. It is a reddish-looking substance; does not form salts. Borax and glass in a melted state dissolve this substance, and the glass acquires a beautiful amethyst colour.

Varvicite, $Mn_4O_7 = MnO + 2MnO_3$.—This compound resembles the binoxide, but is harder and more brilliant. It was first obtained from some manganese ores of Warwickshire.

The *chloride of manganese*, MnCl, may be prepared from the residue left in the retort after the preparation of chlorine from the binoxide of Mn and HCl. A solution of the chloride of manganese, has usually a pink colour, which deposits crystals of the same colour on evaporation. It dissolves readily in alcohol, with which it forms a crystallizable compound.

Sulphate of the protoxide of manganese, MnO,SO_3, $+ 7Ho$; or, $(Mn,SO_4 + 7HO)$, is largely used in dyeing. It has a beautiful rose colour, and is prepared by raising to a red heat in a closed vessel the binoxide of manganese and coal, and dissolving out the impure oxide thus obtained with sulphuric acid. Towards the end of the process a little hydrochloric acid is added; the solution is evaporated to dryness, and again exposed to a red heat. Water is then added, which dissolves out the pure sulphate of manganese. The salt is used to produce a permanent brown dye. It forms, like magnesia, a double salt with potash.

The carbonate, MnO,CO_2, is prepared by precipitating the protochloride of manganese with any alkaline carbonate. It is not soluble. When exposed to heat, it parts with its carbonic acid, and absorbs oxygen.

Manganic Acid, MnO_3.—If the binoxide of manganese be fused with an alkali, an additional quantity of oxygen is absorbed from the air, and a deep green mass is obtained, which is a salt of manganic acid. The presence of nitrate of potash facilitates the production of this acid. MnO_2, by combining with another equivalent of O, becomes MnO_3, which unites with the alkali, forming KO,MnO_3. It is of a green colour, soluble in water, and if evaporated in vacuo, yields green crystals.

Permanganic Acid, Mn_2O_7.—If the manganate of potash, free from excess of the alkali, is put into a large quantity of water, it is decomposed into the hydrated binoxide of manganese, which falls to the bottom, and forms a deep purple liquid, which contains the permanganate of potash. The changes of colour which accompany this decomposition have procured for the substance the name of *mineral chame-*

leon. When the potash is in excess the manganate is more stable, and the reaction is very slow.

The manganates and permanganates are decomposed by contact with organic matter: this is the basis of their use for all those purposes of disinfection and purification for which Condy's fluid is so remarkable. Having no odour, they admit of very extensive application for domestic purposes, and are less objectionable than ordinary disinfectants.

Nearly all the salts of this metal are soluble in water, giving rose-coloured solutions. All the salts are soluble in hydrochloric acid.

Tests.

Hydrosulphuric acid, HS, produces no precipitate with solutions of the salts of manganese. NH_4S precipitates the metal completely as a flesh-coloured sulphide, soluble in strong acetic acid, and passing, by the absorption of oxygen, into a brownish black colour.

The fixed caustic alkalies precipitate the metal as a white hydrate, which becomes brown.

The alkaline carbonates precipitate the white carbonate of the oxide of manganese.

The carbonates of the alkaline earths do not decompose manganese salts, except by an increase of temperature.

If any compound of manganese be fused with carbonate of soda on a platinum foil, the green manganate of soda is produced. This is a very delicate test.

An amethyst bead is produced in the oxidizing flame of the blowpipe when any compound of manganese is fused with borax; the colour disappears in the reducing flame.

ALUMINIUM.

Symbol, Al; *Equivalent,* 13·69.

There is but one oxide of this metal, namely, alumina, Al_2O_3.

In combination with silica and potash, as the silicate of alumina and potash, $Al_2O_3,3SiO_3 + KOSiO_3$, it is very abundant in nature, forming felspar, one of the constituents of granite, from which our various modifications of clay are derived. Clay consists chiefly of a silicate of alumina, $Al_2O_3,3SiO_2$. Pure clay, such as that employed for porcelain manufacture, is quite white: it is derived from the colourless felspar occurring in the granite rocks of Cornwall and Devonshire. Clay is sometimes red, owing to the presence of oxide of iron; sometimes blue or black, from the excess of vegetable matter.

The pure metal aluminium may be obtained by introducing into a platinum tube closed at one end the sesquichloride of alumina, Al_2Cl_3, with an equal quantity of potassium loosely wrapped up in platinum foil. The closed end of the tube is then heated, so as to sublime the chloride, bringing its vapours into contact with the melted potassium. The sesquichloride of aluminium is then reduced. The metal is separated from the chloride of potassium by cold water. It has metallic lustre, it is malleable, and of a tin-white colour. When heated in air or oxygen, it takes fire and returns to its original oxide.

Alumina, Al_2O_3.—If a solution of common alum be mixed with an excess of ammonia, a white gelatinous precipitate of the hydrate of alumina is thrown down. This precipitate is washed, dried, and ignited to whiteness. The alumina so obtained is a white, tasteless mass, scarcely affected by acids. It is fusible before the oxyhydrogen blowpipe. The ruby and sapphire are transparent varieties of crystallized alumina, and emery powder is a coarse variety of this oxide. Alumina is a very weak base, and its salts have frequently an acid reaction.

The *sesquichloride of alumina*, Al_2Cl_3, may be prepared from pure alumina (obtained by the method just described) mixed with charcoal and strongly calcined in a covered crucible. The contents of the crucible are then transferred to a porcelain tube heated to redness in a furnace; a current of chlorine passes through the tube, when the sesquichloride of alumina condenses in the cold part of the apparatus. It is a crystalline substance and very soluble in water.

Sulphate of alumina, $Al_2O_3,3SO_3+18HO$; or, $(2Al,3SO_4 + 18\,HO)$, is obtained by boiling sulphuric acid with clay, or by saturating dilute sulphuric acid with the hydrate of alumina $(Al_2O_3,3HO)$, and evaporating, when thin crystallized scales of the sulphate of alumina are obtained; it has an acid reaction, and when heated to redness is decomposed. The sulphate of alumina, combined with soda, potash, and ammonia, forms double salts. All the preparations of alumina are obtained from common alum, which is a double salt of potash and alumina, $Al_2O_3,3SO_3 + KO,SO_3 + 24HO$; $(2Al,3SO_4 + K,SO_4 + 24HO)$. Alum is chiefly manufactured from a slaty kind of clay which contains the bisulphide of iron. This alum schist, as it is called, is first gently roasted, then moistened with water and exposed to the atmosphere. The sulphur of the iron is converted into sulphuric acid by the absorption of oxygen from the air. The sulphate of the protoxide of iron and the sulphate of alumina are produced. These salts are separated by lixiviation with water and subsequent crystallization; or the solution is concentrated by evaporation, and mixed with the chloride of potassium, which decomposes the iron salt forming the protochloride of iron and sulphate of potash. This sulphate of potash then combines with the sulphate of alumina to form alum. The chloride of iron, which is very soluble in water, is separated from the alum by crystallization.

By a repetition of this process the alum is purified. It is largely used in preparing skins, dyeing, and pigments. Alumina has a great attraction for the fibre of cloth and organic colouring matter. If a solution of common alum be filtered through a piece of calico, part of the alumina will be abstracted and retained by the cloth. If this cloth be then rinsed in an infusion of madder, a colour is produced, which is insoluble in water, and unaffected by soap.

Substances which act in this manner are called ***mordants.*** The attraction of alumina for organic colouring matter may be also shown by adding to an infusion of litmus a solution of alum, and then a solution of soda, so as to precipitate the alumina. As the precipitate falls, the colouring matter will be carried with it. Such precipitates are called *lakes.* Ammonia and soda may be made to replace the potash in common alum.

The *Silicate of Alumina.*—$Al_2O_3,3SiO_2$; $(2Al,3SiO_4)$—is the next most important salt, since it enters into the composition of many crystallized substances. In the primary unstratified rocks, the silicate of alumina is very abundant. All clays are silicates of alumina, but their composition is very variable. By boiling with sulphuric acid the alumina is dissolved out, and the silica left behind. Clay containing the carbonate of lime is called marl. These may be easily recognized by their effervescing with hydrochloric acid.

Analysis of Best Stourbridge Fire Clay.

Water	6·70
Silica	72·04
Alumina	18·79
Peroxide of iron	1·68
Lime	·69
Magnesia	trace
And traces of alkalies.	
	99·90

Tests.

Alumina in solution can be distinguished without difficulty.

Caustic potash, or soda, added drop by drop, will occasion a white gelatinous precipitate, which is the hydrate of alumina, soluble in an excess of the reagent.

$$Al_2O_3,3SO_3 + 3(NaO,HO) = \\ Al_2O_3,3HO + 3(NaO,SO_3).$$

Ammonia produces a similar precipitate.

The alkaline carbonates and the carbonate of ammonia precipitate the hydrate of alumina, with the escape of carbonic acid. The precipitates are insoluble in excess.

IRON.

Symbol, Fe; *Equivalent*, 28.

This is one of the most important of the metals. The progress of commerce and civilization is intimately connected with the uses and properties of iron. The pure metal is exceedingly rare, but its compounds are widely distributed throughout nature. It forms a great part of the colouring matter of many rocks; it is contained in plants, and forms an essential component of the blood. Pure iron is of a white colour, and crystallizes in the form of a cube. The fibrous texture of the iron (upon which its strength mainly depends) is not natural to the metal; it is produced by hammering and rolling, and may always be observed when the metal has been attacked by rust or the application of an acid. Iron is the most tenacious of all metals; a wire $\frac{1}{36}$ of an inch in diameter will sustain a weight of 60 pounds.

The chief source of this metal is the carbonate of the protoxide. The most important and abundant source is the clay ironstone, which consists of the carbonate of the protoxide, usually associated with clay, the carbonates of lime and magnesia, silica, manganese, and coaly matter. This ore contains from 20 to 45 per cent. of iron. The iron ore is first roasted, to get rid of the water, carbonic acid, and carbonaceous matter. The roasted ore is then introduced into a blast-furnace. The ore is introduced with lime and coal at the top of the furnace, the two latter substances being employed as reducing agents. A very high temperature is kept up by forcing air in at the bottom of the furnace. Formerly cold air was employed, but now the air is heated by passing it through hot pipes; the air by this means is raised to a temperature of 600° or 800°, and this is called the hot blast. This effects a great saving of fuel, and coal may be used instead of coke. The limestone introduced into the furnace separates the silicate of alumina from the iron, so that the coal can act upon the metal. The clay combining with the lime forms a glassy compound of the silicate of alumina and lime, which is called slag. The oxide of iron parts with its oxygen to the coal. As the iron melts it falls to the bottom of the furnace, uniting in its descent with the white hot carbon of the coal, forming a carbide of iron, which accumulates in a stratum at the bottom. The silicate of lime and alumina, having less specific gravity than the melted iron, floats like cream upon the surface. The iron and slag are drawn off at different openings at the bottom of the furnace. The iron is run into grooves made of sand, and this constitutes the pig iron of commerce. This iron is associated with carbon, and generally with silicon, manganese, magnesium, calcium, sulphur, and phosphorus. Cast iron is crystalline, brittle, and much more fusible than wrought iron.

Malleable wrought iron is prepared from cast iron, by a process called puddling. The object of this process is to deprive the iron of its carbon. The cast iron is laid on the floor of a reverberating furnace. The stream of burning gases is bent over upon the iron lying on the bed of the furnace. The workman, by means of a large iron rod like a boat-oar, stirs up the melted mass, so as to bring every part successively to the surface. The object of this is to oxidise the carbon of the iron by contact at a high temperature with oxygen. The carbonic oxide thus formed burns with its beautiful blue flame on the surface of the melted iron. Water is sometimes sprinkled on the melted mass, so as to supply oxygen to the carbon. The withdrawal of the carbon diminishes the fusibility of the metal. It first becomes thick and treacly, then like coarse sand. The temperature of the furnace is now increased; the iron again becomes about the consistence of treacle. The workman gathers the iron round his puddling-rod into large balls, which are exposed to hammering by steam hammers. All the liquid cast iron contained in these balls is squeezed out, and the particles of iron welded together in one mass. This mass is heated a second time and passed between large iron rollers. The fibrous structure of iron is mainly due to this process. Iron which is to sustain great tension is submitted, with slight modification, to this latter process several times. This iron is no longer crystalline; it is tough and fibrous in its texture, which may be seen by dissolving a piece of iron wire in dilute acid. By violent concussion this toughness and fibrous character disappears, and it again becomes crystalline and brittle. This malleable iron, although much purer than the cast iron, is never free from carbon and other elements.

The purest and best malleable iron is made into steel by placing alternately in an iron box bars of iron with charcoal powder, and exposing the whole to high red heat for forty or fifty hours. The iron combines with a small portion of the carbon. It is now more fusible than malleable iron, but not so fusible as cast. It has also lost in a great degree its malleability. If fine steel is required, several bars are welded together under the steam hammer, which gives greater uniformity of composition.

Steel possesses great hardness and elasticity. If raised to a high temperature and plunged into water it becomes elastic, hard, and brittle; if it be slowly cooled it softens, loses its hardness and elasticity, and becomes very much like malleable iron. By altering the temperature and the rapidity of cooling every degree of hardness can be obtained. This is called tempering.

Iron does not rust or oxidize in dry air at the common temperature; heated to redness it becomes covered with the black oxide, the same compound as is produced when iron wire is burnt in oxygen. Iron at a red heat decomposes water, like potassium at the common temperature, liberating the hydrogen. Dilute sulphuric and hydrochloric acids dissolve iron freely. If magnetic iron be raised to a red heat it loses its magnetic properties.

There are four oxides of iron :—

FeO, protoxide,
Fe_2O_3, sesquioxide, or peroxide,
Fe_3O_4, protosesquioxide, or black oxide,
FeO_3, ferric acid.

The *Protoxide*, FeO, has strongly marked basic properties; it neutralizes acids most completely. It is scarcely known in its separate state, owing to its proneness to pass into the sesquioxide. If any salt of the protoxide be mixed with caustic alkali a bulky white precipitate of the hydrated protoxide is thrown down; this hydrate when exposed to air becomes first green then reddish brown. The soluble salts of the protoxide of iron have usually a faint green colour.

The *Sesquioxide*, Fe_2O_3, occurs as a native compound, both amorphous and crystalline. The island of Elba has furnished some very fine specimens. The sesquioxide of iron in powder has a deep red colour, and is used for painting. It is prepared by calcining the sulphate of the protoxide. The colour varies according to the temperature to which the sulphate has been exposed. The sesquioxide is reduced at a high temperature in the presence of carbon or hydrogen, but in ordinary fires it undergoes no change. It is a feeble base, dissolves with difficulty in acids, and its salts have an acid reaction.

The *Black Oxide*, $FeO + Fe_2O_3$, is a natural product, known as the magnetic iron ore; it is one of the most important iron ores. This compound is the chief product of burning iron in oxygen. It does not form salts by union with any acid.

Ferric Acid, FeO_3.—This compound is unimportant.

Protochloride of Iron is prepared by passing well-dried HCl gas over red-hot malleable iron, or by dissolving iron in HCl. The solution when duly concentrated yields green crystals of the protochloride with four equivalents of water.

Sesquichloride, Fe_2Cl_3, is generally prepared by dissolving the sesquioxide in HCl. The solution, concentrated and evaporated, yields crystals of a red colour, which are very soluble in water and alcohol.

Protoiodide of Iron, FeI, is an important medicine. It is easily prepared by digesting iodine with water and metallic iron. The solution, which is of a green colour, yields on evaporation crystals of the protoiodide of iron.

There are several compounds of iron with sulphur, but only two are important. The *protosulphide*, which is a black, brittle substance, from which hydrosulphuric acid is prepared. The protosulphide of iron, FeS, is obtained by throwing into a red-hot crucible two and a half parts of sulphur and four parts of iron filings. The *bisulphide*, FeS_2, is a natural product; it is formed in some cases by the gradual deoxidation of the sulphate of iron by organic matter. When heated, an intermediate sulphide is produced analogous to the black oxide.

Sulphate of the Protoxide, $FeO,SO_3 + 7HO$ $(Fe,SO_4 + 7HO)$.—This salt is prepared by dissolving iron in dilute sulphuric acid. It is often prepared on a very large scale by contact of air and moisture with iron pyrites, which, by absorption of oxygen, furnishes the sulphate of iron. The salt is dissolved out by water, and the solution crystallized. It forms large green crystals.

Sulphate of the Sesquioxide, $Fe_2O_3,3SO_3$; $(2Fe,3SO_4)$.—This salt yields compounds with the sulphate of potash and ammonia analogous in their constitution to common alum.

Carbonate of Protoxide of Iron.—If any of the protosalts of iron be mixed with the alkaline carbonates a white precipitate is obtained. This compound loses its carbonic acid when washed and dried, passing into a higher oxide. This compound occurs in nature as the clay ironstone, from which nearly all the British iron is manufactured. It is found in many mineral waters, being soluble in excess of carbonic acid. Water containing iron is known by the rusty matter it deposits by exposure to the air.

The salts of the protoxide of iron are colourless, but if concentrated have a pale greenish colour; when anhydrous they are white. They absorb oxygen, passing into the sesquioxide, which from neutral solutions is deposited as a yellow salt.

Tests.

In solutions of the *protosalts*,

HS forms only a slight precipitate of the black sulphide in neutral solutions; in acid solutions no precipitate is produced except in the acetate.

NH_4S throws down iron as a black precipitate, insoluble in excess. When any non-volatile organic acid is present, a small quantity of the sulphide remains suspended, imparting to the solution a light green colour; on heating, this is deposited, and the liquid becomes of a bright yellow colour.

The caustic alkalies precipitate from salts of the protoxide of iron a white hydrate of the oxide, which on exposure to the air becomes a dirty green, then a reddish brown.

Ammonia will completely precipitate the salts of the protoxide of iron when no ammoniacal salts are present.

Alkaline carbonates precipitate the white carbonate of the oxide, soluble in the chloride of ammonium.

Carbonates of the alkaline earths give no precipitate.

Ferrocyanide of potassium gives, in solution, quite free from the sesquioxide and common air, a white precipitate of the ferrocyanide of iron and potassium, $K + 2Fe + 3Cy$. This precipitate assumes a bluish colour when the sesquioxide and air are present, and passes into Prussian blue.

Ferricyanide of potassium produces with protosalts of iron a blue precipitate. These precipitates are insoluble in acids, but are easily decomposed by the alkalies.

All the soluble *persalts* of iron form, with water or hydrochloric acid, brownish-yellow solutions.

Metallic iron, sulphurous acid, and hydrosulphuric acid, reduce salts of the sesquioxide to salts of the protoxides.

The hydrosulphuric acid renders the solution milky by the precipitation of sulphur —

$Fe_2O_3,3SO_3 + Fe = 3(FeO,SO_3).$
$Fe_2O_3,3SO_3 + SO_2 = 2(FeO,SO_3) + 2SO_3.$
$Fe_2O_3,3SO_3 + HS = 2(FeO,SO_3) + HO,SO_3 + S.$

Potash, soda, and ammonia throw down the hydrate of the sesquioxide, insoluble in excess; the presence of fixed organic acids prevents the precipitation altogether.

The carbonates of the alkaline earths throw down completely the salts of the sesquioxide.

Phosphate of soda precipitates the white phosphate of iron, which dissolves in ammonia.

The phosphate of the sesquioxide of iron is insoluble in acetic acid.

Ferrocyanide of potass, blue precipitate, ferrocyanide of iron insoluble in HCl.

Ferricyanide, no precipitate, but a darker colour.

The sulphocyanide of potassium gives a deep red colour to the most dilute solutions of the sesquioxide of iron.

The most dilute solutions of salts of the peroxide give a bluish-black colour with an infusion of gall-nuts.

NICKEL.

Symbol, Ni; *Equivalent*, 29·6.

This metal is found in meteoric stones, and in certain ores associated with cobalt, and may be prepared from the oxalate. The oxalate is obtained by precipitating the sulphate with oxalic acid. It is white, malleable, and strongly magnetic. It is prepared like cobalt.

The *Protoxide of Nickel*, NiO, has basic properties. It may be prepared by precipitating any soluble salt of the metal with caustic potash, and washing, drying, and igniting the precipitate, which is of a green colour. This oxide completely neutralizes acids.

The *Sesquioxide of Nickel*, Ni_2O_3, is a neutral compound. It is produced when any salt of nickel is mixed with a solution of bleaching powder. This oxide is decomposed by heat, and evolves chlorine when heated with hydrochloric acid.

The *Chloride of Nickel*, NiCl, is prepared by dissolving the carbonate of nickel with hydrochloric acid, $NiO,CO_2 + HCl = NiCl + HO + CO_2$. From the green solution crystals of the chloride of nickel are obtained. When the crystals are anhydrous they have a yellow colour.

Sulphate of Nickel, $NiO,SO_3+7HO,(Ni,SO_4+7HO)$.—This is the most important of the nickel salts. The crystals are of a green colour, and usually contain seven equivalents of water. With potash and ammonia it forms double salts: $NiO,SO_3 + KO,SO_3 + 6HO$; $(Ni,SO_4 + K,SO_4 + 6HO)$; $NiO,SO_3 + NH_4O,SO_3 + 6HO$; $(Ni,SO_4 + NH_4,SO_4 + 6HO)$.

Carbonate of Nickel, NiO,CO_2—When a solution of the carbonate of soda is added to the chloride of nickel, a pale green precipitate of the carbonate and hydrate is obtained: $NaO,CO_2 + NiCl = NiO,CO_2 + NaCl$.

The anhydrous salts of nickel are of a straw colour. When associated with water they are of a green colour. Nickel is chiefly employed in the manufacture of German silver. The best German silver contains 8 parts of copper, 3 of nickel, and 3½ of zinc; inferior kinds contain 8 of copper, 2 of nickel, and 4 of zinc.

Tests.

Hydrosulphuric acid and sulphide of ammonium have a similar reaction on nickel solutions to those described under cobalt.

Ammonia in neutral solutions precipitates a portion of hydrate of the protoxide, which is dissolved in excess, giving rise to a blue solution. If caustic potash be added to this solution, a green hydrate is thrown down.

The alkaline carbonates give a pale-green precipitate, easily soluble in carbonate of ammonia, the solution becoming of a bluish-green colour. Carbonates of the alkaline earths, same as cobalt.

Cyanide of potassium, a greenish-white precipitate, soluble in excess. From this solution the cyanide may be thrown down by hydrochloric or sulphuric acid, but the cyanide of cobalt cannot be again precipitated with acids.

In the outer flame of the blowpipe, compounds of nickel, when heated with borax and phosphorus salt, form a reddish-coloured globule; in the inner flame the metal is reduced, and the bead becomes of a grey colour.

COBALT.

Symbol, Co; *Equivalent*, 29·55.

Cobalt and *Nickel* are very similar in their general and chemical properties, and are often found in nature associated together. Cobalt is most conveniently obtained by heating the oxalate to a white heat, in a crucible lined with plumbago. It is a white brittle metal, and melts at a temperature of about 2000°. It does not oxidize in air at common temperatures. Dilute hydrochloric and sulphuric acids have but slight influence on the metal. It has strong magnetic properties.

The *Protoxide*, CoO, is a grey powder, very soluble in acids, and has strong basic properties, giving a fine red tint to its salts. It may be prepared by precipitating the chloride of cobalt with carbonate of soda, and igniting the precipitate; $CoCl + NaO,CO_2 = NaCl + CoO,CO_2$; or, $(CoCl + NaCO_3 = CoCO_3 + NaCl)$. When a solution of cobalt is mixed with caustic potash, a blue precipitate falls, which becomes violet on the application of heat; and then a dirty red colour from the absorption of oxygen, and its subsequent change into a hydrate.

The *Sesquioxide of Cobalt*, Co_2O_3, is a neutral oxide of a black colour, and insoluble. If the chloride of lime be mixed with a solution of protochloride of cobalt, the sesquioxide is obtained.

The *Chloride of Cobalt*, CoCl, is prepared by dissolving the protoxide in hydrochloric acid, $CoO + HCl = CoCl + HO$. It gives a deep red colour to the solution, which deposits, when concentrated, beautiful red crystals. When the solution is evaporated, it deposits blue anhydrous crystals, which again become red in contact with water. A dilute solution of this salt forms a sympathetic ink. Letters written with this ink are invisible, until the salt has been rendered anhydrous by exposure to heat, when the letters appear blue. When laid aside, moisture is absorbed, and the writing again becomes invisible.

Sulphate of Cobalt, $CoO,SO_3 + 7HO,(Co,SO_4 + 7HO)$. This salt is similar in constitution to the sulphate of magnesia. It combines with the sulphates of potash and soda, forming double salts. If a solution of the sulphate of cobalt be mixed with oxalic acid, the oxalate of cobalt separates after some time.

Carbonate of Cobalt.—The alkaline carbonates precipitate a pale peach-coloured carbonate and hydrate of cobalt, $3(CoO,HO) + 2(CoO,CO_2)$.

The oxide of cobalt gives a beautiful blue colour to glass. Smalt, or azure blue, a substance much used by painters, paper-stainers, and for giving a blue colour to pottery, is prepared in Saxony and Bohemia by a complicated process from the native arsenio-sulphide of cobalt.

Tests.

All the cobalt salts, when anhydrous, are blue; when in dilute solutions in water, they are of a crimson colour. Hydrosulphuric acid, in the presence of an alkaline acetate, and the sulphide of ammonium, precipitate the black sulphide of cobalt, insoluble in excess of the sulphide of ammonium. In neutral solutions, free from the salts of ammonia, ammonia gives a blue precipitate, which after some time changes to red, then green, then brown. If the solution contain any salts of ammonia, the solution becomes red, absorbs oxygen from the air, and the brown hydrate of cobalt is precipitated. Cobalt solutions, free from common air, give a blue precipitate with the fixed caustic alkalies, which on boiling becomes of a rose colour.

The alkaline carbonates precipitate a pink carbonate of cobalt.

The carbonates of the alkaline earths give no precipitate unless heated. Cyanide of potassium gives, with acid solutions of cobalt, a brown-white precipitate of the cyanide of cobalt, soluble in excess. Borax and phosphorus salts, when heated with any cobalt salt, give a blue glass globule in both blowpipe flames.

CHROMIUM.

Symbol, Cr; *Equivalent*, 28.

This metal is found as an oxide in combination with the oxide of iron. It is rather abundant in the Shetland Isles. Chromium was first obtained in a state of purity from the chromate of lead. The metal is procured by mixing the oxide of chromium with $\frac{1}{5}$ its weight of charcoal powder, and placing the mixture in a crucible lined with charcoal. The crucible is then placed in a furnace, and raised to the very highest heat. The oxide of chromium is reduced by the charcoal, and the metal is obtained as a greyish-white brittle substance, and very difficult of fusion. The most powerful acids have little effect on chromium, and the pure metal has been seen by very few persons.

There are seven oxides of chromium, resembling very much the corresponding oxides of iron.

Protoxide of Chromium, CrO.—If potash be added to a solution of the protochloride of chromium, a brown precipitate is obtained, which passes into a deep red colour. If ammonia be substituted for potash, a pale green precipitate is obtained, which is the protoxide of chromium, CrO, This oxide is a powerful base, forming pale blue salts, which have a great affinity for oxygen.

Protosesquioxide, CrO,Cr_2O_3.—This compound is produced by the action of water upon the protoxide, with the evolution of hydrogen. It corresponds to the magnetic oxide of iron.

Sesquioxide, Cr_2O_3.—If the nitrate of the suboxide of mercury be mixed with the chromate of potash, and raised to red heat in a crucible, the chromate of mercury is formed, and the sesquioxide of chromium remains behind. It is used in painting and glass-staining. It has a fine green colour. After ignition, it is insoluble in acids. The hydrate of the sesquioxide, $Cr_2O_3,3HO$, may be prepared by boiling a dilute solution of the bichromate of potash with excess of hydrochloric acid, to which a small quantity of alcohol is every two or three minutes added, till the liquid changes from orange to green. From this solution the hydrate is precipitated by ammonia as a pale green colour.

The anhydrous sesquioxide may also be prepared by heating to redness the bichromate of potash in an earthen crucible. The sesquioxide of chromium is a feeble base, resembling the sesquioxide of aluminum. It communicates a beautiful green tint to glass, and is used in enamel painting.

Protochloride, $CrCl$, is obtained by heating to redness the sesquichloride in a porcelain tube through which a current of pure dry hydrogen is passing. HCl is liberated; a white mass is obtained which dissolves in water with a great increase of temperature, yielding a blue solution which absorbs oxygen by exposure, and passes into a deep green colour. The protochloride of chromium is one of the most powerful reducing or deoxidizing agents known.

Sesquichloride, Cr_2Cl_3.—This compound may be obtained anhydrous by heating to redness in a porcelain tube the sesquioxide of chromium intimately mixed with charcoal, and passing dry chlorine gas through the tube. The sesquichloride condenses in the cold part of the tube in plates of a beautiful violet colour. This compound is quite insoluble in water.

Chromic acid, CrO_3.—When the sesquioxide of chromium is heated with potash in contact with air, oxygen is absorbed, and CrO_3 generated. If 100 measures of a saturated solution of the bichromate of potash be mixed with 150 measures of sulphuric acid, and allowed to cool, the chromic acid crystallizes out in brilliant red prisms. Chromic acid is very soluble in water, and is instantly reduced by contact with organic matter.

The *Chromate of Potash*, KO,CrO_3 (K,CrO_4), is the source from which all the chromic preparations are obtained. It is made from chrome iron ore, which is a compound of iron, $FeO + Cr_2O_3$. It is reduced to powder, calcined and heated for a long time in a furnace with the nitrate of potash, KO,NO_5. The product is treated with water, which yields a yellow solution, which deposits, on evaporation, crystals of the same colour. If sulphuric acid be added to the chromate of potash, KO,CrO_3, one half of the base is removed, and the neutral chromate converted into the bichromate, $KO,2CrO_3$. This salt is extensively used in the arts, and, by slow evaporation, crystallizes in beautiful red crystals. It melts when heated, is soluble in ten parts of water, and has an acid reaction.

Chromate of Lead, PbO,CrO_3, (Pb,CrO_4).—When a solution of the nitrate of lead is mixed with a solution of chromate of potash, a yellow precipitate falls, which is the chromate of lead or the chrome yellow of the painter. If this compound be boiled with lime-water, half of the acid is removed, and a sub-chromate formed, which has an orange-red colour. The colour of this subchromate rivals vermilion in the beautiful delicacy of the tint. The yellow and orange chrome colours are fixed on cloth by the alternate application of the two solutions.

Chromate of Silver, AgO,CrO_3, (Ag,CrO_4).—When the nitrate of silver is mixed with the chromate of potash, the chromate of silver is precipitated as a reddish-brown powder.

Tests.

With Chromates:—

Lead salts give a lemon-yellow precipitate of the chromate of lead.

Baryta salts give a pale yellow precipitate insoluble in acetic acid.

Silver salts, purple-red precipitate.

The chromium compounds impart a beautiful green colour to borax in the inner and outer flame of the blowpipe.

Salts of the sesquioxide of chromium may be easily recognized:—

The caustic akalies precipitate the hydrated oxide, easily soluble in excess of the reagent.

Ammonia precipitates a bluish-green hydrate, which is slightly soluble in strong ammonia, forming a pink solution.

The alkaline carbonates throw down a green precipitate, soluble in excess.

Sulphuretted hydrogen, no change.

Sulphide of ammonium precipitates the hydrate of the sesquioxide.

ZINC.

Symbol, Zn; *Equivalent*, 32·55.

This metal is generally found in nature associated with lead ores, as the carbonate ZnO,CO_2, and sulphide ZnS. Our chief supplies of this metal are from Silesia. The ore is first roasted to expel the water and carbonic acid. It is then mixed with charcoal or coke, and distilled at a red heat in a large earthenware retort. The reduced metal rises in vapour. This vapour is condensed in a suitable apparatus. The metal generally contains small portions of arsenic.

The metal is of a bluish colour, and is much used instead of lead, in consequence of its cheapness. It is crystalline, and under ordinary circumstances brittle; but raised to a temperature of 270° it is malleable, and may be rolled out in sheets without fracture. By such treatment the metal retains its malleability. At a temperature of 400° it may be reduced to powder. At a red heat it boils and volatilizes, and in the presence of air burns with a green light, giving rise to the protoxide of the metal, ZnO.

Zinc is readily dissolved by dilute acids. It is generally used in the preparation of hydrogen.

The only known *oxide* of this metal is the protoxide ZnO. It is a strong base and may be prepared by raising the carbonate to a red heat.

The neutral *carbonate* of zinc, ZnO,CO_2, or (Zn,CO_3), is found native; it is the chief ore of the metal and is called calamine stone. It may be prepared by adding the carbonate of potash to a solution of the chloride of zinc, $ZnCl + KO,CO_2 = ZnO,CO_2 + KCl$; or $(ZnCl + K,CO_3 = ZnCO_3 + KCl)$.

The *Sulphate of Zinc*, $ZnO,SO_3 + 7HO$, or $(Zn,SO_4 + 7HO)$, *White Vitriol*, resembles in appearance the sulphate of magnesia. This salt forms the chief part of the residue left when hydrogen is evolved by the action of dilute sulphuric acid on zinc. The sulphate of zinc is used in medicine as an emetic. It forms double salts with potash and ammonia, $ZnO,SO_3 + KO,SO_3 + 6HO$; or $(Zn,SO_4 + K,SO_4 + 6HO)$.

The *Chloride of Zinc*, $ZnCl$, is prepared by dissolving the oxide of zinc in hydrochloric acid, $ZnO + HCl = ZnCl + HO$. It is very fusible and soluble in water and alcohol. The chloride of zinc combines with the chloride of ammonium to form a salt used in tinning and soft-soldering the metals. The chloride of zinc is used as a disinfectant, and wood saturated with a solution of the chloride is not affected with dry rot. The salts of zinc are all soluble in hydrochloric acid, and most of them are soluble in water. In acid solutions hydrosulphuric acid produces no precipitate, but if alkaline acetate be added, the white sulphide of zinc is thrown down. This precipitate is insoluble in potash.

Tests.

Sulphide of ammonium precipitates the sulphide of zinc. Caustic alkalies give a white precipitate of the hydrate, soluble in excess.

The carbonates of the fixed alkalies precipitate the basic carbonate, insoluble in excess, but soluble in any salt of ammonia.

Carbonate of ammonia precipitates the carbonate, soluble in excess.

The carbonates of the alkaline earths give no precipitate with zinc salts.

Zinc compounds when fused with carbonate of soda on charcoal in the inner flame of a blowpipe, give a white incrustation of the oxide of zinc. When moistened with cobalt and solution strongly heated, the colour becomes green.

CADMIUM.

Symbol, Cd; *Equivalent*, 56.

This metal is more volatile than zinc. It is generally associated with its ores. When the calamine stone is subjected to distillation with charcoal the cadmium first comes off in vapour. It has a white tin colour, very malleable; burns when strongly heated. Hydrochloric and sulphuric acid have no effect on the metal at ordinary temperatures. It dissolves easily in nitric acid, and slightly oxidizes in air.

The *Protoxide of Cadmium*, CdO, may be prepared by igniting the carbonate in a crucible; it is infusible, dissolves in acids, and its salts are without colour.

The *Sulphate of Cadmium*, $CdO,SO_3 + 4HO$; or $(Cd,SO_4 + 4HO)$. — This salt is obtained by dissolving the oxide in dilute sulphuric acid, $CdO + HO,SO_3 = HO + CdO,SO_3$. It is very soluble in water and forms double salts with ammonia and potash.

The *Sulphide of Cadmium*, CdS, is obtained by passing sulphuretted hydrogen gas through a solution of the chloride or any salt of cadmium, $CdCl + HS = CdS + CHl$.

Tests.

Hydrosulphuric acid and the sulphide of ammonium precipitate the yellow sulphide of cadmium; insoluble in alkalies.

Ammonia precipitates a white hydrate, soluble in excess, from which hydrosulphuric acid precipitates the sulphide.

The fixed caustic alkalies produce a precipitate of the white hydrate, insoluble in excess.

The alkaline carbonates precipitate the carbonate of cadmium; insoluble in excess.

Cadmium compounds, when exposed to the inner blowpipe flame with carbonate of soda or charcoal, give a yellow-brown incrustation of the oxide of cadmium.

TIN.

Symbol, Sn; *Equivalent*, 59.

Tin occurs in nature as the oxide and sulphide, but chiefly as the former. The chief sources of this metal are the mines of Cornwall, Saxony, Bohemia, and Malacca. It is also found as an alluvial deposit in the beds of many small rivers. These two varieties are called mine tin and stream tin. The latter has a much higher commercial value.

Tin ore is reduced by first grinding it to powder, then roasting to get rid of the water, sulphur, and arsenic. It is then heated with coal, and the reduced metal is cast into large square blocks; a specimen of the metal is assayed, and the stamp of the Duchy of Cornwall placed on each block. Tin, when quite pure, is of a white colour, soft and malleable. It is not much affected at ordinary temperatures by exposure to water or air: when raised to its melting point it rapidly oxidizes, passing into a white powder, a substance which is used in the arts for polishing, and known as putty powder. The metal is easily dissolved by hydrochloric and nitric acid.

The *Protoxide of Tin*, SnO, is prepared by adding a solution of carbonate of potash to a solution of the protochloride of tin:—

$$SnCl + KO,CO_2 + HO = KCl + CO_2 + SnO,HO.$$

When the precipitated hydrate is dried and heated in an atmosphere of carbonic acid, it becomes a heavy black powder, unalterable in air, but burns like tinder in contact with a hot body, and changes into binoxide. The protoxide has no basic properties.

Sn_2O_3.—If a solution of the protochloride of tin be mixed with a solution of the hydrated sesquioxide of iron, a grey substance falls, which is the sesquioxide of tin.

The *Binoxide of Tin*, SnO_2, has different properties according to the method by which it is prepared. If nitric acid be made to act on pure metallic tin, a white hydrate of the binoxide of tin is produced, insoluble in acids. If the bichloride of tin be precipitated by an alkali, a white precipitate is obtained of the hydrate of the binoxide, freely soluble in acids; SnO_2,HO.

These precipitates have the same composition; and when heated lose their water and leave the pure binoxide of tin. They dissolve in caustic potash, and are precipitated unchanged by acids. The binoxide is a feeble base, and sometimes acts as an acid, forming salts with bases.

The *Protochloride of Tin*, SnCl, is obtained by dissolving metallic tin in hot hydrochloric acid. It is soluble in water and used as a deoxidizing agent for the salts of mercury and silver.

The *Bichloride of Tin*, $SnCl_2$, may be prepared by exposing metallic tin to the action of chlorine, or distilling 1 part of powdered tin with 5 parts of the protochloride of mercury. A solution of the bichloride of tin is much used as a mordant.

Protosulphide of Tin, SnS, is prepared by fusing metallic tin with an excess of sulphur, and strongly heating the residue. With hydrochloric acid gently heated it evolves hydrosulphuric acid.

Sesquisulphide of Tin, Sn_2S_3, may be prepared by gently heating the protosulphide with a third part of its weight of sulphur.

The *Bisulphide of Tin*, SnS_2, *Mosaic Gold*, is prepared by raising to a low red heat in a glass flask 12 parts of tin, 6 of the chloride of ammonium, and 7 of flowers of sulphur. The chloride of ammonium and part of the sulphur sublime, the bisulphide of tin collecting at the bottom in brilliant gold-coloured scales, which are used for some purposes as a substitute for gold.

If a solution of the terchloride of gold be added to a dilute solution of the protochloride of tin, a precipitate is obtained which is known as the purple of Cassius, which is used in enamel painting. Tinned vessels are extensively used for domestic purposes. Thin plates of iron are superficially covered with tin. Good pewter is tin hardened with antimony.

Salts of the protoxide of tin pass by exposure to the atmosphere into salts of the binoxide.

Tests.

Protoxide of Tin.—The caustic alkalies precipitate the white hydrate, soluble in excess of the precipitant.

The carbonates of ammonia, potash, and soda, produce a similar precipitate, insoluble in excess. Ammonia gives the same result.

Sulphuretted hydrogen and the sulphide of ammonia give a black precipitate of the protosulphide, soluble in sulphide of ammonium.

Binoxide of Tin.—The caustic alkalies precipitate the white hydrate soluble in excess.

Ammonia and the alkaline carbonates produce a white precipitate of the hydrate, slightly soluble in excess. Sulphuretted hydrogen and sulphide of ammonium a yellow precipitate of the sulphide, soluble in excess.

ANTIMONY.

Symbol, Sb; *Equivalent*, 129.

This metal is found in combination with sulphur. The ore is heated with metallic iron or the carbonate of potash, which removes the sulphur. It is of a bluish colour, highly crystalline, having a metallic lustre. It is so brittle that it can be reduced to powder in a pestle and mortar. It melts below a red heat, and volatilizes at a white heat. It is not acted upon by air at the common temperature. If strongly heated, it burns with a white flame producing crystals of the teroxide of antimony.

There are three oxides of antimony.

Teroxide of Antimony, SbO_3.—This oxide is feebly basic. It may be prepared by burning metallic antimony at the bottom of a large red-hot crucible. There is a sulphate, a nitrate, and an oxalate of this oxide. When the teroxide of antimony, SbO_3, is boiled with the bitartrate of potash, it is dissolved, and the solution, by careful evaporation, yields crystals of tartar emetic, $KO,SbO_3,C_8H_4O_{10} + 2HO$, the tartrate of antimony and potash.

There is an intermediate oxide of antimony, SbO_4, which is the ultimate product of oxidizing the metal by heat, with exposure to the atmosphere. It is an indifferent body.

Antimonic Acid, SbO_5.—When strong nitric acid is made to act on metallic antimony, antimonic acid is produced. When heated carefully it becomes anhydrous, and of a pale straw colour. At a red heat it is decomposed. It is insoluble in water and acids. The acid so prepared is monobasic:—MO,SbO_5; or, (M,SbO_6). If the pentachloride of antimony, $SbCl_5$, be decomposed by excess of water, hydrochloric acid and antimonic acid are produced. The acid so produced is bibasic; —$2MO,SbO_5$, as, KO,HO,SbO_5, or, KH,SbO_7. This latter salt is the only precipitant for soda. It must be added with care, or no result is obtained.

The *Terchloride of Antimony*, $SbCl_3$; *Butter of Antimony.*—This compound is obtained when hydrosulphuric acid is generated by the action of hydrochloric acid on the tersulphide of antimony. The residual solution is then distilled until each drop of the condensed vapour, on falling into the water of the receiver, produces a white precipitate. The receiver is then changed, and the process carried on. The air should be carefully excluded. Alkaline solutions decompose the terchloride, producing the teroxide. Powdered antimony takes fire in chlorine gas, forming $SbCl_3$.

The *Pentachloride of Antimony*, $SbCl_5$, corresponds to antimonic acid. It is prepared by passing a stream of dry chlorine over metallic antimony, gently heated in a porcelain tube.

The *Tersulphide of Antimony*, SbS_3, is the native sulphide from which the metal is obtained.

The *Pentasulphide*, SbS_5, is unimportant.

There is also a compound of antimony and hydrogen, which is prepared by adding the teroxide of antimony, and zinc, to a solution of dilute sulphuric acid. Antimoniuretted hydrogen comes off, which burns with a green flame. When a cold plate is pressed into the flame, metallic antimony is deposited. Antimony is used in medicine; and, with lead, it forms the type-metal of the printer. It is too brittle to be used alone, but its alloys have numerous applications.

The salts of the teroxide of antimony are decomposed by the addition of a large quantity of water, a white basic salt being separated. Tartaric acid dissolves this compound, and, when present in the solution, prevents its formation.

Tests:

Hydrosulphuric acid precipitates from acid solutions of the teroxide the tersulphide of antimony, soluble in sulphide of ammonium and in potash, but insoluble in the bicarbonate of ammonia.

Ammonia precipitates the teroxide of antimony, insoluble in excess.

Carbonate of ammonia produces the same result, with the evolution of carbonic acid.

The fixed caustic alkalies precipitate the teroxide, soluble in excess. The presence of tartaric acid interferes with the reaction.

Metallic zinc precipitates from salts of the teroxide black metallic antimony. In the presence of nitric acid the teroxide is separated.

Nitric acid precipitates antimonic acid from any of its compounds.

Hydrosulphuric acid throws down from acid solutions of antimonic acid an orange-yellow pentasulphide of antimony, SbS_5.

Compounds of antimony, exposed with carbonate of soda to the reducing flame of the blowpipe, yield globules of the metal, a portion of which volatilizes at a high temperature, forming a white incrustation of the teroxide on the charcoal.

COPPER.

Symbol, Cu; *Equivalent,* 31·50.

This metal is sometimes found pure, but generally associated with sulphur and iron. Large quantities of this ore are obtained from the mines of Cornwall and South Australia. The ore is taken to South Wales to be reduced. The process of reduction is rather complicated. The ore is roasted in a reverberatory furnace. The sulphide of iron is converted at a high temperature into the oxide; but the sulphur still exists in combination with the copper. The roasted ore is then heated with silicious sand, which combines with the iron, forming a fusible slag which separates from the sulphide of copper. The iron, by repeating this process, is at last removed. The sulphide of copper then decomposes, giving up its sulphur, and combining with oxygen. The oxide is then reduced by charcoal.

Copper is a good conductor of electricity. It is of a reddish colour, and very malleable. In a dry atmosphere, it undergoes no change; in a moist atmosphere, it becomes covered with a green crust of the carbonate of the metal. Dilute hydrochloric and sulphuric acid act slowly upon copper. Boiling sulphuric acid and cold nitric acid dissolve it readily; the first yielding sulphurous acid, and the second, nitric oxide, or binoxide of nitrogen. There are two oxides of copper which form salts:—the *Protoxide*, CuO; and the *Suboxide*, Cu_2O. The *Protoxide* may be prepared by decomposing the nitrate in a crucible at a red heat. If carbonate of copper in solution be mixed with a solution of caustic potash, a bulky blue precipitate of the hydrated oxide falls, which, on being boiled, becomes of a dark-brown colour: this is the anhydrous oxide. The protoxide unites with all the acids, forming a series of important salts. The *suboxide* of copper (Cu_2O) often occurs, beautifully crystallized, with copper ores. It may be obtained artificially by heating in a covered crucible 5 parts of the black oxide and 4 parts of copper filings, or by adding sugar to a solution of the sulphate of copper, and then adding caustic potash in excess. The sugar deprives the black protoxide of half its oxygen, the suboxide precipitating as a brown-red powder. The salts of suboxide are colourless and very unstable, owing to their tendency to absorb oxygen, and pass into higher oxides. The protoxide gives a green colour to glass: the suboxide, a red colour.

Sulphate of Copper, *Blue Vitriol*, $CuO, SO_3 + 5HO$; or, $(Cu\,SO_4 + 5HO)$.—The old copper sheathing of ships is calcined and dissolved in dilute sulphuric acid: or the salt may be procured by oxidizing the sulphide. This salt crystallizes in large blue crystals, soluble in water; and, when anhydrous, is nearly of a white colour. At a high temperature the salt is decomposed. It combines with potash and ammonia, forming double salts of a blue colour.

Nitrate of Copper, $CuO, NO_5 + 3HO$; or $(Cu, NO_6 + 3HO)$.—This salt is prepared by dissolving the metal in nitric acid. It forms deep blue crystals very soluble in water.

Carbonate of Copper, CuO, CO_2; or, (Cu, CO_3).—Carbonic acid, in various proportions, combines with the oxide of copper. When a solution of the carbonate of soda is added in excess to a solution of the sulphate of copper, a pale blue precipitate is obtained. On heating, it becomes sandy, and of a green colour. It has the following composition: $CuO,CO_2 + CuO,HO + HO$. This salt is used in painting. The green mineral malachite has a similar composition. A natural crystalline compound of the carbonate of copper and water, having a deep blue colour, which has not yet been imitated artificially, has the following composition: $2(CuO,CO_2) + CuO,HO$. Verditer is said to be similar in its composition.

Chloride of Copper, $CuCl + 2HO$, is easily prepared by adding hydrochloric acid to the protoxide of copper. It forms green crystals very soluble in water and alcohol. When slightly heated, it loses its water of crystallization, and becomes of a yellowish-brown colour. At a high temperature, it loses half its chlorine, becoming the subchloride, Cu_2Cl. This compound may also be formed by introducing copper filings or Dutch gold leaf into a jar of chlorine gas.

Arsenite of Copper (*Scheele's Green*) is prepared by mixing a solution of the sulphate of copper with a solution of the arsenite of potash. It is a bright green powder, and is used in painting.

Brass is an alloy of 70 parts of copper and 30 parts of zinc. Gun-metal is an alloy of 90 parts of copper and 10 of tin. Bell-metal contains a larger proportion of tin, Good bronze consists of 91 parts of copper, 2 parts of tin, 6 of zinc, and 1 of lead. The brass of the ancients contained tin instead of zinc.

Tests.

Hydrosulphuric acid, and the sulphide of ammonium, precipitate the black sulphide of copper.

The fixed caustic alkalies throw down the blue hydrate of copper, which, on boiling, loses its water and becomes black.

Ammonia produces a greenish-blue precipitate, which dissolves in excess, producing a beautiful blue azure colour.

The carbonates of the fixed alkalies throw down a greenish-blue carbonate of copper, which, by boiling, becomes black in an excess of the precipitant.

Cyanide of potassium gives a yellow-green precipitate, soluble in excess.

Ferrocyanide of potassium produces a dark, purple-red precipitate.

Copper may be thrown down in very dilute neutral solutions in a pure metallic state, by introducing a bright piece of metallic iron, which becomes at once covered with copper. When heated upon charcoal with carbonate of soda in the inner blowpipe flame, metallic copper is obtained.

LEAD.

Symbol, Pb; *Equivalent*, 103·56,

The most abundant ore of lead is the native sulphide (galena), which not unfrequently contains small quantities of silver. The ore is placed on the floor of a reverberatory furnace. A considerable portion of the sulphide becomes by oxidation converted into the sulphate. The ore is then thoroughly mixed together, the temperature raised, the sulphide and sulphate react upon each other, producing sulphurous acid and metallic lead:—

$$PbO, SO_3 + PbS = 2SO_2 + 2Pb;$$
$$\text{or, } Pb, SO_4 + PbS = 2SO_2 + 2Pb.$$

Lead is very malleable, it has no elasticity, and very little strength. It melts at 600°, and volatilizes at a white heat. In moist air it becomes coated with a grey film of the suboxide. When melted it absorbs oxygen, passing rapidly into the protoxide. All the acids, except nitric, act slowly upon the lead. There are four oxides of lead.

Protoxide, *Litharge*, *Massicot*, PbO.—This oxide is obtained when melted lead is exposed to a current of hot air, or by heating the carbonate to a dull redness. When carefully prepared, it is of a pale yellow colour; when raised to a red heat, it fuses and forms crystals, on cooling, of a brick-red colour. It is slightly soluble in water, giving an alkaline reaction. In a melted state, it dissolves in silicious matter with great energy. It is largely used in the manufacture of glass, upon which it confers great brilliancy. It unites with acids, forming a large class of colourless salts, if the acid be without colour.

Red Oxide of Lead, *Red Lead*, $Pb_3O_4 = 2PbO + PbO_2$.—This oxide is prepared by exposing for a long time to a dull red heat the protoxide of lead which has not been fused. It is a brilliant red powder. By the action of acids it is converted into the protoxide and binoxide. It is used in painting, and sold as a substitute for vermilion.

The *Binoxide* or *Brown Oxide*, PbO_2, is obtained by digesting red lead in dilute nitric acid. At a red heat it is converted into the protoxide. Hydrochloric acid converts it into the chloride. When heated with sulphuric acid, it forms sulphate of lead, one equivalent of oxygen being liberated. The binoxide of lead is used to separate sulphurous acid from gaseous mixtures, by converting the sulphurous acid into sulphuric,—$PbO_2 + SO_2 = PbO,SO_3$.

The *Suboxide of Lead*, Pb_2O, is obtained when the oxalate of lead is heated to a dull red heat in a retort. It absorbs oxygen with avidity, passing into the protoxide.

Nitrate of Lead, PbO,NO_5; or (Pb,NO_6).—If the carbonate of lead be dissolved in dilute nitric acid, nitrate of lead is formed, and the carbonic acid liberated; or the protoxide may be dissolved in dilute nitric acid, and concentrated till the crystals separate.

Carbonate of Lead, White Lead, PbO, CO_2; or, (Pb, CO_3).—This salt is sometimes found crystallized in long white needles, with other metallic ores. It is prepared by adding in the cold the carbonate of soda to the acetate of lead. It is, however, largely manufactured for painting in the following manner:—Into a great number of earthen jars, arranged side by side, a little vinegar is poured. These jars stand on manure or tan; to each jar, a roll of sheet lead is adapted, or bars cast into the shape of a portcullis are placed, so as not to touch the vinegar or project above the jar. The jars are closely covered with boards, and a layer of tan or manure is placed upon the top. Another row of pots is then placed upon the top of these, covered in like manner. By this arrangement a pile of many rows is constructed. After the lapse of some days, the sheets or bars of lead are taken out of the pots, when it is found that they are in great part converted into the carbonate, which is made ready for use by washing and grinding. The nature of the change is not well understood; it is thought that an acetate of lead is first formed, which is decomposed by the carbonic acid of the decomposing manure.

If clean metallic lead be placed in pure water with exposure to the air, a white powder makes its appearance, covering the whole of the lead. This deposit is chiefly due to the action of the carbonic acid dissolved in the water, and consists of a combination of the carbonate and hydrate. The carbonate, although very insoluble in pure water, is slightly soluble in water containing free carbonic acid, which is nearly always the case. If the water were quite pure it would be contaminated with lead, and in course of time the whole of the lead would be dissolved. As all lead compounds are very poisonous, pure water, or water containing much free carbonic acid, cannot therefore be kept in leaden cisterns, or passed through leaden pipes. In common river and spring water, however, a little sulphate is almost always present; this forms a thin film of the insoluble sulphate of lead, which prevents further action. It is for this reason that leaden cisterns may be used without danger.

The *Chloride of Lead*, PbCl.—This salt may be prepared by adding litharge to boiling dilute hydrochloric acid, filtering, and setting aside the filtrate to cool, or by mixing a saturated solution of the acetate of lead with the chloride of sodium.

Iodide of Lead, PbI.—This salt is obtained as a beautiful yellow precipitate when any soluble salt of lead is mixed with a solution of the iodide of potassium.

Tests.

Lead salts are thrown down as a black sulphide on the addition of hydrosulphuric acid or sulphide of ammonium. The insoluble compounds of lead, by digestion with sulphide of ammonium, yield the sulphide if hydrochloric acid is present. The precipitate approaches a brownish-red colour.

The soluble chlorides, or hydrochloric acid, precipitate the white chloride, which is soluble in an excess of water or potash.

Sulphuric acid precipitates the white sulphate.

Ammonia precipitates all the lead salts except the acetate, as the white hydrate, insoluble in excess.

The chromates precipitate the yellow chromate of lead from solutions of the oxide, insoluble in dilute nitric acid.

The fixed caustic alkalies precipitate the white hydrate, soluble in excess.

The alkaline carbonates precipitate the white carbonates, insoluble in excess, but soluble in caustic potash.

The compounds of lead, when heated in the blowpipe on charcoal with the carbonate of soda, yield a globule of metallic lead.

BISMUTH.

Symbol, Bi; *Equivalent*, 213.

This metal is of a reddish colour. It is separated from the earthy matter with which it is associated by heating the mixture in a crucible with charcoal. It volatilizes at high temperatures, is slightly oxidized by contact with air, and dissolves freely in dilute nitric acid.

The *Teroxide of Bismuth*, BiO_3, is the base of all the bismuth salts. It fuses at a high temperature, and forms a powerful flux with silicious substances.

Nitrate of Bismuth, $BiO_3,3NO_5 + 10HO$; or, $(Bi,3NO_6 + 10HO)$.—This salt is prepared by dissolving bismuth in nitric acid. If the solution be left for some time, large crystals of the neutral nitrate of bismuth are deposited. There is a basic nitrate of bismuth, $BiO_3NO_5 + 2HO$. This salt was extensively used as a cosmetic, but it injures the skin. It is occasionally used in medicine.

Eight parts of bismuth, 5 of lead, and 3 of tin, form a fusible metal used for taking impressions from seals. Bismuth also enters into the composition of the best kinds of type.

Tests.

Hydrosulphuric acid and sulphide of ammonium precipitate a dark tersulphide of bismuth, BiS_3.

The alkalies precipitate a white hydrate, insoluble in excess.

The alkaline carbonates precipitate a white carbonate, slightly soluble in an excess of the precipitant.

Chromate potash precipitates a yellow chromate of bismuth, insoluble in caustic potash, but soluble in nitric acid.

Bismuth salts give, when heated upon charcoal with carbonate of soda in the inner flame of the blowpipe, a yellow incrustation.

MERCURY.

Symbol, Hg; *Equivalent*, 100.

This important metal has been known from a very early period. It is found native, but chiefly in combination with sulphur. This ore is sometimes called cinnabar. It is found chiefly in Spain; and large quantities, in a state of great purity, have recently been imported from California. The mercury is separated from the sulphur by heating the ore in an iron retort, with scraps of iron or lime, at a high temperature. The mercury rises in vapour, which passes over into a large chamber, where the mercury is condensed, and sulphurous acid escapes. Mercury is imported in iron bottles, each holding seventy-five pounds. The iron bottles are sometimes used for preparing oxygen from the binoxide of manganese. Mercury is sometimes adulterated with tin and lead. Mercury, at ordinary temperatures, is a liquid. At 662° it boils, and passes into a colourless vapour of great density; at a temperature of —40° F. it solidifies. At ordinary temperatures mercury undergoes no change. When near its boiling point it absorbs oxygen, passing into a dark-red powder, which is the red protoxide, and highest oxide of the metal. At a temperature approaching red heat it is decomposed into metallic mercury and oxygen. It was by this method that Priestley discovered the presence of oxygen gas. Dilute hydrochloric and sulphuric acid have no effect on the metal at ordinary temperatures. Concentrated boiling sulphuric acid converts mercury into the sulphate of the protoxide of the metal, with evolution of sulphurous acid. Nitric acid, at the common temperature, dissolves mercury without difficulty, with evolution of the binoxide of nitrogen.

There are two oxygen compounds of mercury, the grey or suboxide, Hg_2O, and the red or protoxide, HgO. The salts of the latter are the most permanent in their character. The isomorphous relationship of mercury to the other metals is not well established.

The *Suboxide of Mercury*, Hg_2O, is prepared by digesting calomel, Hg_2Cl, with a solution of caustic potash, or by adding caustic potash to the nitrate of the suboxide. It is a dark grey powder, insoluble in water. Mercury and chalk, and grey ointment are preparations of this oxide.

The *Protoxide of Mercury*, HgO, may be prepared by heating the nitrate of the protoxide until the acid is expelled, or by adding an excess of caustic potash to a solution of corrosive sublimate, $HgCl$. A bright yellow precipitate of the protoxide is thrown down. It is slightly soluble in water, and very poisonous.

Sulphate of the Suboxide of Mercury, Hg_2O,SO_3, or Hg, SO_4) is prepared by adding sulphuric acid to a solution of the nitrate of the suboxide.

Sulphate of the Protoxide of Mercury, HgO, SO_3, or (Hg, SO_4), may be obtained by boiling equal weights of sulphuric acid and mercury, until the metal is converted into a white crystalline powder, which is the sulphate of the protoxide of the metal. The solution is evaporated to dryness. Water decomposes this salt, forming an insoluble yellow compound called turbith mineral, $3HgO,SO_3$. By continual washing with hot water, the acid is entirely removed, leaving the original protoxide of the metal.

The *Nitrate of the Suboxide of Mercury*, Hg_2O,NO_5 or (Hg_2NO_6) + $2HO$, is obtained by dissolving mercury in an excess of dilute nitric acid. If the mercury be in excess a salt is formed, containing $3Hg_2O,2NO_5 + 3HO$. Both these salts are decomposed by water. These salts, having a similar composition, may be easily distinguished by rubbing them in a mortar with a little chloride of sodium. The nitrate of the suboxide gives nitrate of soda and calomel. The basic salt gives a black compound of the nitrate of soda, calomel, and protoxide of mercury.

The *Nitrate of the Protoxide of Mercury*, HgO,NO_5, or (Hg,NO_6), is prepared by dissolving the red oxide in excess of nitric acid, and gently evaporating. When the solution, which has this composition, $HgO,NO_5 + 2HO$, has become thick, it is placed under a bell jar over strong sulphuric acid, when large crystals are deposited, consisting of $2(HgO,NO_5) + HO$.

The *Subchloride of Mercury*, *Calomel*, Hg_2Cl, is a very important compound. It may be prepared by pouring a solution of the nitrate of the suboxide into a large excess of dilute solution of common salt. Calomel is usually prepared by the following method. The dry sulphate of the red oxide of mercury is rubbed in a mortar with the same weight of mercury the salt contains, and a quantity of the chloride of sodium. When a uniform mixture has been produced, it is placed in a retort. It sublimes at a temperature below a red heat. The vapour is conveyed into a chamber, and condensed in a finely divided state :—

$$HgO,SO_3 + Hg + NaCl = Hg_2Cl + NaO,SO_3;$$

$$\text{or, } Hg,SO_4 + Hg + NaCl = Hg_2Cl + NaSO_4.$$

Calomel is a white heavy insoluble powder. It is not affected with cold nitric acid, but hot nitric acid dissolves it. It is decomposed by lime water, with production of the suboxide. If calomel should contain any chloride of mercury it is very dangerous. This chloride may be detected by boiling the calomel with water, filtering the solution, and adding caustic potash. Any corrosivesublimate is detected by a yellow precipitate.

The *Protochloride of Mercury*, $HgCl$, may be prepared by dissolving the red oxide in hot hydrochloric acid, when crystals of corrosive sublimate separate as the solution cools. It is generally prepared by subliming a mixture of the dry sulphate of the protoxide of mercury with an equal weight of dry common salt :—

$$HgO,SO_3 + NaCl = HgCl + NaO,SO_3,$$
$$\text{or, } Hg_2SO_4 + NaCl = HgCl + Na_2SO_4.$$

Corrosive Sublimate is a white crystalline compound. It absorbs ammoniacal gas freely, is soluble in water, and forms a series of double salts:—

$$HgCl + NH_4Cl + HO.$$

Organic substances containing nitrogen unite with corrosive sublimate, forming insoluble compounds. It is on this account that albumen is a good antidote to poisoning by corrosive sublimate. Kyan's patent for preserving timber, cordage, &c., consisted in saturating the substances with a solution of corrosive sublimate.

Subiodide of Mercury, Hg_2I, is obtained when a solution of the iodide of potassium is added to a solution of the nitrate of the suboxide of mercury. The suboxide separates as an insoluble dirty yellow precipitate.

The *Protoiodide of Mercury*, HgI, is prepared by adding, in equivalent proportions, a solution of the iodide of potassium to the protochloride of mercury. The precipitate is first yellow, but changes into a brilliant permanent scarlet colour.

Subsulphide of Mercury, Hg_2S, is thrown down as a black precipitate when a solution of hydrosulphuric acid is mixed with a solution of the nitrate of the suboxide of the metal. It is decomposed by heat into sulphur and metallic mercury.

The *Protosulphide of Mercury*, *Cinnabar*, *Vermilion*, HgS.—When hydrosulphuric acid is passed gradually into a solution of corrosive sublimate, a white precipitate is obtained, a compound of the salt and sulphide. An excess of gas converts it into the black sulphide. When this black sulphide is sublimed, it undergoes no change of composition, but it becomes a dark red crystalline compound known as cinnabar and vermilion. When the protosulphide is heated in air, it yields sulphurous acid and metallic mercury. The protosulphide is not attacked by alkalies or strong mineral acids. It dissolves in aqua regia.

Mercury dissolves many of the metals, as gold, silver, lead, tin. A mixture of tin and mercury is used in silvering glass. The compounds are called amalgams.

All the basic salts of the suboxide are white; the neutral salts are yellow.

Salts of the suboxide are reduced by zinc, iron, copper, and the sulphate of the protoxide of iron.

Tests.

With salts of the *suboxide* :—

Hydrosulphuric acid precipitates the black subsulphide, Hg_2S.

The sulphide of ammonium produces the same result, insoluble in excess. It is decomposed into metallic mercury and sulphide of mercury on the addition of potash.

The soluble chlorides and hydrochloric acid precipitate the white subchloride, which is converted into the black suboxide on the addition of any caustic alkali.

The alkaline carbonates precipitate the carbonate of the suboxide.

Salts of the *protoxide* are precipitated as the black protosulphide, on the addition of hydrosulphuric acid and the sulphide of ammonium. This precipitate is soluble in caustic potash if the sulphide of ammonium be in excess. On the addition of a small quantity of these reagents the precipitate is white; but on a greater addition the black sulphide is produced, insoluble in concentrated boiling nitric acid, but easily dissolved in nitro-hydrochloric acid.

Ammonia, and the carbonate of ammonia, give a white precipitate.

All the compounds of mercury, when heated in a test-tube with carbonate of soda, yield globules of metallic mercury easily seen through a lens. Mercury, in neutral and slightly acid solutions, may be detected on the introduction of a piece of bright copper foil. After a few minutes, a grey metallic covering of mercury is produced.

SILVER.

Symbol, Ag; *Equivalent*, 108.

This metal is found in its pure state, but chiefly as the sulphide, bromide, and chloride, and in combination with lead ores. The chief sources of this metal are the mines of the Hartz mountains in Germany, Kongsberg in Norway, and the Andes in America. The ores are too poor to admit of the metal being extracted by smelting, although fuel may be abundant in the neighbourhood. It is therefore obtained by a complicated process of amalgamation with mercury. Pure silver may be obtained by dissolving the metal in pure nitric acid. If copper be present the solution will be of a faint green tint. If gold be present, a black undissolved powder will remain. If the solution be now mixed with hydrochloric acid, a white insoluble precipitate of the chloride of silver will be obtained. This is washed, dried, and mixed with twice its weight of anhydrous carbonate of soda, and heated in an earthen crucible, gradually raised in temperature until it approaches a white heat. Oxygen and carbonic acid are evolved, and metallic silver is found in a button at the bottom of the crucible. Silver is of a white colour, of great lustre, malleable, ductile, and an excellent conductor of heat and electricity. It is unalterable in air and water. When strongly heated in air or oxygen, it absorbs many times its volume of that gas; but when the metal solidifies it parts with its oxygen, giving a rough arborescent appearance to the surface of the metal. The addition of a small portion of copper prevents the absorption of oxygen. When silver is heated with powdered glass, it forms a silicate of an orange colour. Hydrochloric acid has little effect on silver. Hot sulphuric acid converts it into a sulphate. Nitric acid dissolves it easily. The metal has a great affinity for sulphur. Sulphuretted hydrogen tarnishes the surface with a black sulphide. This is the reason silversmiths usually burn their gas outside the shop.

There are three oxides of silver:—

The *Suboxide* Ag_2O, which is unimportant.

The *Protoxide of Silver*, AgO, is a powerful base, neutralizing the most powerful acids, which form colourless salts. It may be prepared by precipitating the nitrate of silver with caustic potash. It is soluble in ammonia and slightly soluble in pure water. At a red heat, it is decomposed into metallic silver and oxygen.

The *Peroxide* is unimportant.

The *Nitrate of Silver*, *Lunar Caustic*, AgO,NO_5, or (Ag,NO_6), is prepared by dissolving metallic silver in nitric acid. The solution is concentrated and evaporated. When melted and poured into cylindrical moulds it forms the *lunar caustic* of the surgeon. When exposed to light, the salt blackens, especially if any organic matter be present. It stains the skin black, and is employed as a dye for the hair, and for the manufacture of marking-ink for linen.

Sulphate of Silver, AgO,SO_3, or (Ag,SO_4).—This salt is prepared by precipitating a concentrated solution of the nitrate of silver with sulphate of potash.

The *Hyposulphate of Silver*, $AgO,S_2O_5 + HO$; or $(Ag, S_2O_6 + HO.)$—This salt is of a white colour, very soluble and permanent.

The *Hyposulphite of Silver*, $AgO,S_2O_2 + HO$; or $(Ag, S_2O_3 + HO$, is a very unstable compound, insoluble, and of a white colour. It combines with any alkaline hyposulphite, forming soluble compounds of a sweet taste. The oxide and chloride of silver give rise to similar compounds with the alkaline hyposulphites.

The *Carbonate of Silver*, AgO,CO_2; or (Ag,CO^3), is prepared by mixing a solution of the nitrate of silver with a solution of the carbonate of soda.

Chloride of Silver, AgCl.—This salt is produced when any silver compound is precipitated with a soluble chloride. It is of a white curdy appearance, quite insoluble in water and nitric acid. It is sometimes found native, and is called horn silver. The chloride of silver is decomposed by exposure to light, especially if any organic substances are present. It is reduced to its metallic state when put into water with metallic zinc or iron, and a little sulphuric acid. It is soluble in ammonia and the cyanide of potassium.

The *Iodide of Silver*, AgI, is obtained when a solution of the iodide of potassium is added to the nitrate of silver. It is nearly insoluble in ammonia. In this respect it is an exception to the silver salts generally. The bromide resembles the chloride.

The *Sulphide of Silver*, AgS, is obtained by precipitating a solution of silver with hydrosulphuric acid.

The salts of silver are mostly without colour, and leave in most cases metallic silver after ignition. Silver compounds are decomposed by zinc, iron, copper, sulphate of the protoxide of iron, and chloride of tin, metallic silver being precipitated.

Tests.

Hydrosulphuric acid and sulphide of ammonium precipitate the black sulphide of silver.

Hydrochloric acid and the soluble chlorides throw down a white curdy precipitate of the chloride of silver, soluble in ammonia, but insoluble in nitric acid.

The fixed caustic alkalies give a brown precipitate. Ammonia gives a brown precipitate of oxide, soluble in excess.

The alkaline carbonates throw down the white carbonate of silver, soluble in an excess of the carbonate of ammonia.

PLATINUM.

Symbol, Pt: *Equivalent*, 98·7.

This metal in a crude state, associated with palladium, rhodium, iridium, osmium, and iron, is found among the diluvial accumulations on the slope of the Ural mountains in Russia, Ceylon, and South America. It has never been seen in the rocks formed of any of the materials accompanying it. It is extracted from the alloy by a very difficult and complicated process.

Platinum is a white metal, very malleable, ductile, and infusible, scarcely yielding at times to the oxyhydrogen blowpipe. It is the heaviest known substance. Neither air, moisture, nor ordinary acids have the slightest effect on platinum. It is dissolved in nitro-hydrochloric acid, and oxidized by the fused hydrate of potash. Spongy platinum and platinum black have the property of condensing gases into their pores, especially oxygen. It is this property which is thought to determine the union of hydrogen and oxygen when a piece of spongy platinum is held before a jet of hydrogen.

There are two oxides of platinum.

The *Protoxide*, PtO, may be obtained by digesting the protochloride with caustic potash: the protoxide of platinum is produced as a black powder, soluble in an excess of potash.

The *Binoxide of Platinum*, PtO_2, is prepared by boiling the bichloride with an excess of caustic soda, then adding acetic acid. A bulky brown powder is produced, which becomes black and anhydrous when gently heated.

The *Bichloride of Platinum*, $PtCl_2$, is the most important compound. This compound is formed when platinum is dissolved in nitro-hydrochloric acid. On evaporating to dryness, a reddish-brown residue is left, very soluble in water and alcohol. The solution has an orange-yellow colour. It forms a variety of double salts with metallic chlorides. When a solution of the bichloride of platinum is added to any solution of potash, a bright yellow crystalline precipitate is produced of the bichloride of platinum and the chloride of potassium, $PtCl_2 + KCl$. A few drops of hydrochloric acid promote the formation of the precipitate.

The *Chloride of Platinum*, $PtCl$, is produced when sulphurous acid is passed into a solution of the bichloride until it ceases to give a precipitate with chloride of ammonium. It is a greenish powder insoluble in water, but soluble in hydrochloric acid.

Tests.

A solution of the binoxide or bichloride of platinum can be recognized by a yellow precipitate with chloride of ammonium or chloride of potassium. Hydrosulphuric acid on the application of a moderate heat precipitates the bisulphide, of a brownish-black colour, insoluble in nitric acid and hydrochloric acid, but soluble in the sulphide of ammonium and nitro-hydrochloric acid. Chloride of tin, in the presence of free hydrochloric acid, gives a reddish-brown colour to solutions of platinum.

GOLD.

Symbol, Au; *Equivalent*, 197.

This metal is widely diffused throughout nature. Small quantities are found in iron pyrites and the granitic rocks. It is always found in a pure metallic state, beautifully crystallized in a cubic form, associated with sand, quartz, and oxide of iron. The sands of various rivers contain this metal in a finely-divided state. This is the gold-dust of commerce.

Gold does not oxidize in air or water. It is not affected by any single acid. A mixture of hydrochloric and nitric acid is used to dissolve it. Gold melts at a temperature of

2016° Fahr. It is the most malleable and ductile of all the metals. The thinnest gold leaf is said not to exceed the $\frac{1}{200000}$th of an inch in thickness.

There are two oxides and two chlorides of gold.

The *Protoxide of Gold*, AuO, is obtained by adding a solution of caustic potash to the protochloride. It is a green powder, slightly soluble. The solution is very unstable, and decomposes into metallic gold and the teroxide of the metal; the latter dissolves, and the former falls in fine powder to the bottom.

The *Teroxide of Gold*, AuO_3, is obtained by adding magnesia to the terchloride. It is a reddish-yellow powder, which dries into a dark-brown colour.

The *Protochloride of Gold*, $AuCl$, is produced by evaporating the terchloride to dryness and exposing it to a heat of about 440°, until chlorine ceases to be given off. It is a yellowish-white mass insoluble in water.

The *Terchloride of Gold*, $AuCl_3$, is the most important compound. It may be produced on a small scale by passing chlorine through water in which pieces of gold leaf are suspended. The gold dissolves, and if the solution be evaporated it yields crystals of the terchloride. It is always produced when gold is dissolved in nitro-hydrochloric acid. It is very deliquescent and soluble in water. The terchloride combines with other metallic chlorides forming double salts. The general formula of these compounds may be thus represented, $MCl + AuCl_3$.

The terchloride of gold, with an excess of the bicarbonate of soda or potash, is used for gilding small ornamental copper articles. They are first cleaned with dilute nitric acid, then boiled for some time in the solution.

Gold coins are alloyed with a proportion of silver or copper to increase their hardness and durability. English standard gold contains $\frac{1}{12}$ of copper.

Gilding on copper is now performed by electrotype. The solution employed for this purpose is obtained by dissolving the oxide or cyanide of gold in a solution of the cyanide of potassium.

The most mportant reducing agents for gold are the sulphate of the protoxide of iron, oxalic acid, and chloride of tin. The sulphate of the protoxide of iron precipitates metallic gold as an impalpable brown powder, which may be collected on a filter. Dried and rubbed with a smooth body, it assumes the yellow colour and metallic lustre of the metal.

The chloride of tin, containing a little bichloride, on the addition of a drop of nitric acid, produces in very dilute solutions of gold a purple precipitate, known as the purple of Cassius.

ARSENIC.

Symbol, As; *Equivalent*, 75.

This metal is of a steel-grey colour. It is sometimes found native, but more frequently in combination with oxygen, sulphur, iron, nickel, and cobalt. These arsenides of the metals are placed in a reverberatory furnace, and the volatile products condensed in a chimney of peculiar construction. The crude arsenious acid is then purified by sublimation. It is then reduced with charcoal in an earthenware retort, and the metal easily sublimes.

Arsenic, when heated, oxidizes into arsenious acid. It combines with the metals, like sulphur and phosphorus, and unites with oxygen in two proportions:—arsenious acid, AsO_3, and arsenic acid, AsO_5. There is no basic compound of this metal. The vapour of arsenic has a strong odour of garlic.

Arsenious Acid, *White Oxide of Arsenic*, AsO_3, is prepared by heating arsenic in contact with air. It is very soluble in hot water, and the acid crystallizes in transparent octohedrons. The alkalies dissolve this compound readily, forming arsenites. Arsenious acid is a powerful poison, In experimenting with arsenic compounds, the student should avoid inhaling the vapours.

Arsenic Acid, AsO_5.—This compound is produced by dissolving arsenious acid in hot hydrochloric acid, and oxidizing with nitric acid so long as red vapours are produced. The solution is then carefully evaporated to dryness. The acid so obtained is white and anhydrous. It slowly and entirely dissolves in water, the solution being slightly acid, which, on being evaporated, yields after some time, crystals of hydrated arsenic acid. When heated, it is decomposed into arsenious acid and oxygen.

Arsenic acid resembles in the closest manner phosphoric acid, forming salts corresponding to the phosphates.

The *Arseniate of Soda*, $2NaO,HO,AsO_5 + 24HO$; cannot be distinguished in its appearance from the phosphate of soda. It may be prepared by adding a solution of carbonate of soda to a solution of arsenic acid until the acid is completely neutralized, and then evaporating.

There are also two other arseniates of soda, having the following composition:— $3NaO,AsO_5 + 24HO$, and $NaO, 2HO,AsO_5 + 24HO$.

There are three sulphides of the metal:—

Realgar, AsS_2, which is found native, and can be formed artificially by heating arsenic with the equivalent proportion of sulphur. It is employed in painting and for making white fire.

Orpiment, AsS_3, is also a natural compound. This may be prepared by precipitating an acid solution of arsenic with hydrosulphuric acid: after the lapse of some time a higher sulphur compound is precipitated.

The *Pentasulphide of Arsenic*, AsS_5, is not important.

The three sulphur compounds of arsenic are all sulphur acids.

Tests.

Arsenious acid can be detected by characters which cannot be mistaken.

The nitrate of silver is added to a solution of arsenious acid and water; on the addition of a drop or two of ammonia, a yellow precipitate of the arsenite of silver is thrown down very soluble in ammonia and nitric acid.

Sulphate of copper, on the addition of a little alkali, added drop by drop, produces a yellow-green precipitate (Scheele's green), soluble in ammonia.

Sulphuretted hydrogen is passed through a solution of arsenious acid, previously acidulated with a few drops of hydrochloric acid; a copious bright yellow precipitate of orpiment is thrown down, soluble in ammonia, and can be reprecipitated by sulphuric acid.

The detection of arsenious acid when mixed with the contents of the stomach, is much more difficult. The presence of organic matter so interferes with the action of the tests just given, as to render their results worthless in a matter of so much importance.

Marsh's plan is, perhaps, the most simple and delicate. The apparatus used consists of a U-shaped tube of large diameter, on one leg of which is blown two bulbs, the uppermost closed by a cock with a small brass stop-cock. The other leg of the tube is open. Slips of pure zinc are placed in the lower bulb by unscrewing the stop-cock. The suspected liquid, which has been acidulated with pure sulphuric acid in the proportion of one ounce of acid to a pint of the liquid, is then introduced among the scraps of zinc. The stop-cock is closed, and hydrogen is generated in the lower bulb, which forces the liquid into the upper bulb. The apparatus now acts by hydrostatic pressure. As soon as the stop-cock is opened the gas is forced through the jet and inflamed; a piece of white porcelain is held in the flame, when any arsenic is at once recognized by the presence of a black, brilliant metallic spot of reduced arsenic on the porcelain.

The evolved hydrogen in the lower bulb combines with the arsenic in the arsenious acid, and the gas which is inflamed is arseniuretted hydrogen, which in contact with the cold porcelain is reduced to a metallic state.

RÉSUMÉ.

POTASSIUM.

Symbol, K; *Equivalent*, 39.

Discovered by Sir H. Davy in 1807.

Physical Properties.—A white, soft metal, capable of being cut with a knife.

Specific Gravity.—0·86.

Chemical Properties.—A simple body absorbing oxygen at all temperatures, and decomposing water in the cold.

Preparation.—By reducing the carbonate of potash and charcoal to powder, and heating in an iron retort; the potassium in a melted state falls into a receiver containing coal-naphtha.

PROTOXIDE OF POTASSA, POTASH.

Symbol, KO; *Equivalent*, 47.

Physical Properties.—A white, solid body. When touched with the fingers it feels soapy. It is fusible at a dull red heat, and volatilizes at a high temperature. Potash unites with water, forming the hydrate of potash, which is the true caustic potash, KO,HO (KHO_2). It is a very powerful alkali, completely neutralizing the strongest acids. On exposure to the air potash absorbs moisture, becomes slightly liquefied; and when converted into the carbonate it melts.

Preparation.—The potash of commerce is an impure carbonate, and is chiefly obtained from wood ashes.

NITRATE OF POTASH, NITRE.

Symbol, KO,NO_5; *Equivalent*, 101.

Physical Properties.—A white salt, crystallizes in long six-sided prisms. It is decomposed at a high temperature.

Chemical Properties.—It is soluble in water, and becomes liquid (deliquesces) in a moist atmosphere. It is a poison when taken in large quantities. With sulphur and carbon it forms gunpowder.

Nitre is found in many parts of the East as a white efflorescence on the surface of the earth. The French manufacture it artificially from urine, chalk, and decomposing animal substances.

Gunpowder differs in its composition according to its use. The following may be taken as the general composition :—

$$KO,NO_5 + S + 3C.$$

SODIUM.

Symbol, Na; *Equivalent*, 23.

Discovered by Sir H. Davy in 1807.

Specific Gravity.—0·97.

Physical Properties.—A white, silver-looking metal, solid at ordinary temperatures, and volatile at a red heat.

Chemical Properties.—Decomposes water, and absorbs oxygen at all temperatures. It is kept in coal-naphtha.

Preparation.—By heating the carbonate of soda and charcoal in iron retorts, the same as described under potassium.

SODA.

Symbol, NaO; *Equivalent*, 31.

Soda is very analogous in all its properties to potash.

The soda of commerce is an impure carbonate, chiefly obtained from common salt.

SULPHATE OF SODA.

Symbol, NaO,SO_3; or (Na,SO_4). *Equivalent*, 71.

Physical Properties.—A white salt, crystallizes in four-sided prisms; soluble in twice its weight of cold water. It is used in the manufacture of carbonate of soda.

CHLORIDE OF SODIUM, COMMON SALT.

Symbol, NaCl; *Equivalent*, 68·5.

Physical Properties.—A white salt, crystallizes in the form of cubes. It is soluble in water, and is obtained by evaporation of brine. Our chief supplies are from the salt mines of Cheshire.

The brine springs of Cheshire yield a large supply. It is sometimes obtained by evaporating sea-water.

AMMONIA.

Symbol, NH_3; Oxide of Ammonium, NH_4O; *Equivalent*, 26.

The compound metal ammonium, of which ammonia is the oxide, has never been isolated.

Chemical Properties.—Ammonia has a powerful odour, and very volatile. It is strongly alkaline, and can be made to replace potash or soda. It is this property of ammonia which gives rise to a series of salts analogous to those of soda and potash.

The salts of ammonia contain an equivalent of water. Hypothetical formula of ammonium, NH_4. The hydrate is the oxide NH_3HO, or NH_4O. All the salts of ammonia are constituted on this theory.

Carbonates of Ammonia.

The neutral anhydrous carbonate of ammonia, $NH_3\ CO_2$, is prepared by the direct union of carbonic acid with ammoniacal gas. It is very soluble in water.

The strong smelling salts of ammonia are prepared by subliming a mixture of sal ammoniac with chalk. It is probably a sesquicarbonate, $2NH_4O,3CO_2$.

Chloride of Ammonium, Sal Ammoniac.

Symbol, NH_4Cl; *Equivalent*, 53·5.

A white, fibrous, crystalline body; volatile at a red heat, and soluble in water and alcohol. It is prepared by acting on the carbonate of ammonia with hydrochloric acid:—

$$NH_4O,CO_2 + HCl = NH_4Cl + HO,CO_2, \text{ or}$$
$$NH_4CO + HCl = NH_4Cl + HO + CO_2.$$

BARIUM.

Symbol, Ba; *Equivalent*, 68·5.

A white, brilliant metal, but speedily oxidises on exposure to air. It is very heavy, poisonous, and decomposes cold water.

BINOXIDE OF BARIUM.

Symbol, BaO_2; *Equivalent*, 84·5.

Is of a greenish-white colour, and amorphous. It can be obtained by heating to redness baryta, BaO, in a current of oxygen or air. The binoxide is deoxidized into baryta at a red heat. This alternate oxidation and deoxidation of baryta has been proposed as an economic source of oxygen.

CALCIUM.

Symbol, Ca; *Equivalent*, 20.

A white metal, discovered by Sir H. Davy in 1809, similar to Barium. Calcium has been seen by very few persons. On exposure to air it passes into the protoxide of calcium, CaO (lime).

LIME.

Symbol, CaO; *Equivalent*, 28.

Lime is a white amorphous body, usually called quicklime. It absorbs moisture with great avidity, passing into the hydrate, CaO,HO (Ca,HO_2), and on exposure to the air it absorbs carbonic acid, becoming the carbonate of lime.

It is more soluble in cold water than hot, forming a solution known as lime-water.

Lime is made from chalk or marble, by strongly heating in a limekiln, which drives off the carbonic acid.

With alumina, silica, and iron, it forms hydraulic cements, which have the property of hardening under water.

Building mortar is a mixture of sand, lime, and water. It hardens on exposure to air. Good mortar ought to contain three parts of road or river sand to one of lime.

CARBONATE OF LIME.

Symbol, CaO,CO_2; or (Ca,CO_3); *Equivalent*, 50.

There are several varieties of carbonate of lime:

The crystallized forms, such as stalactite, spathose, and arragonite; amorphous, such as chalk and limestone; and the granular or crystalline, as marble.

The pure carbonate of lime is without colour; insoluble in pure water but soluble in water containing carbonic acid. It is decomposed by heat into carbonic acid and lime. Carbonate of lime is one source of incrustations in steam-boilers and tea-kettles. By heating water containing this salt the carbonic acid is driven off, and the insoluble carbonate of lime falls to the bottom. Stalactites in caverns are formed by the gradual percolation of water containing the carbonate of lime.

SULPHATE OF LIME, PLASTER OF PARIS, GYPSUM.

Symbol, CaO,SO_3; or (Ca,SO_4); *Equivalent*, 68.

This salt is found in nature both in a crystallized and amorphous state. The anhydrous sulphate is the plaster of Paris; it absorbs water and becomes a hard compact mass. It requires a red heat to deprive it of this water. The sulphate of lime is soluble in 500 parts of cold water. In its crystallized state alabaster is a familiar example.

ALUMINIUM.

Symbol, Al; *Equivalent*, 13·7.

A silver-white metal, discovered by Wöhler in 1828; does not readily oxidize. Aluminium is obtained from a mineral called cryolite, by reduction with potassium or sodium, or by decomposing the chloride of aluminium by the vapour of sodium. Specific gravity, 2·6.

ALUMINA.

Symbol, Al_2O_3; *Combining Equivalent*, 51·4.

Alumina is a very weak base, and its salts have frequently an acid reaction. The ruby and sapphire are crystallized varieties of alumina coloured with some metallic oxide. Emery powder, used for polishing glass and metals, is a coarse kind of crystallized alumina called corundum.

ALUMS.

All the alums are double sulphates. Common alum may be taken as the type, which is a double sulphate of alumina and potash, $Al_2O_3,3SO_3 + KO,SO_3 + 24HO$; or, $(2Al,3SO_4 + K,SO_4 + 24HO)$. Alum is used in the arts for dyeing, and preparing skins.

IRON.

Symbol, Fe; *Equivalent*, 28.

Pure iron is of a greyish-white colour, fusible at a high temperature. It has great tenacity, very ductile. Specific gravity 7·84. On exposure to the atmosphere, iron passes into the oxide; at a red heat it decomposes water.

Oxides of Iron.—The protoxide FeO, which is a black powder, and very oxidizable. The sesquioxide Fe_2O_3, which is of a reddish-brown colour, and is obtained by precipitating with ammonia, any of the salts of the protoxide of iron. The magnetic oxide Fe_3O_4, which is found in nature, and is the product of burning iron wire in oxygen. Ferric acid, FeO_3.

ZINC.

Symbol, Zn; *Equivalent*, 32.

This metal is first mentioned by Paracelsus, who died in 1530. It is of a bluish-white colour, soft, slightly flexible, and at a temperature of 270° very malleable. At a red heat it boils and volatilizes, and if air be present, burns with a green flame, generating the oxide of zinc. Zinc has a specific gravity of 7.

Zinc is much used for roofing, instead of lead. It is durable, but inferior in this respect to lead. Cisterns, pails, and many domestic utensils are now made of zinc.

The protoxide of zinc, ZnO, is a white body, and a powerful base. It may be obtained by the direct oxidation of zinc. It is used as a bad substitute for white lead in painting.

Sulphate of zinc, ZnO,SO_3+7HO, or (Zn,SO_4+7HO) — white vitriol. This salt is amorphous with the sulphate of magnesia. It is of a white colour, soluble in water, but insoluble in alcohol. It is employed in calico printing.

TIN.

Symbol, Sn; *Equivalent*, 58.

Tin is a white silver-looking metal, has great metallic lustre, and is a good reflector of light. It has a specific gravity of 7·29. There are two important oxides of tin, the protoxide SnO, and the binoxide SnO_2. The salts of the protoxide are precipitated with chloride of gold. The precipitate is of a purple colour, and is known as the purple of Cassius.

The salts of the binoxide of tin give a white precipitate with sulphuretted hydrogen.

The salts of tin are used in dyeing. Thin sheets of iron dipped in molten tin, form the tin plate of commerce. A mixture of tin and lead forms the solder of the plumber. Tin and sulphur form mosaic gold. Pewter is tin hardened with antimony. Britannia metal is a compound of variable quantities of copper, tin, bismuth, and antimony.

COPPER.

Symbol, Cu; *Equivalent*, 31·7.

Copper is of a red colour, very ductile and tenacious. When breathed on, it creates a disagreeable odour. It is fusible at a bright red heat, and has a specific gravity of 8·89.

On exposure to a moist atmosphere, copper becomes covered with a film of the carbonate of copper, which is of a green colour, and very poisonous.

There are four oxides of copper—two only are important. The suboxide Cu_2O, and the protoxide CuO. The suboxide is of a red colour, soluble in ammonia. The protoxide is the ordinary base of all the copper salts.

The salts of copper give, with ammonia, first a green, then a blue colour. The sulphate of copper $CuO,SO_3 + 5HO$, or $(Cu,SO_4 + 5HO)$, is of a blue colour, and poisonous. It is used in medicine and painting.

Brass consists of about 30 parts of zinc and 70 of copper. Copper is sometimes externally converted into brass by exposure, when at a red heat, to the vapour of zinc. The gold wire of Lyons is manufactured from copper bars prepared in this manner. Gun metal consists of 90 parts of copper and 10 of tin. Bell metal contains a larger proportion of tin. Bronze consists of 91 parts of copper, 2 of tin, 6 of zinc, and 1 of lead. The brass of the ancients was copper and tin.

LEAD.

Symbol, Pb ; *Equivalent*, 103.

Lead is a bluish-white metal, having great lustre; on exposure to the air it becomes dull, owing to its absorption of oxygen; this thin film protects the metal from further oxidation. It is very malleable; but has little ductility and tenacity. The specific gravity of lead is 11·35. Lead melts at 600°; at a white heat it boils and volatilizes. The salts of lead are without colour, except the chromates; they have a sweet taste, and are poisonous. When clean metallic lead is placed in pure water it becomes covered with a thin film of the hydrated protoxide of lead, formed by the dissolved oxygen of the water. This compound is slightly soluble in water. This deposit, however, is more frequently the carbonate in union with the hydrate of lead, and is a very insoluble compound, except in water containing free

CO_2, which is mostly the case. River and spring waters almost always contain a little combined sulphuric acid, which forms a thin coating of the sulphate of lead, and protects the metal from farther action. If water were quite pure, the lead would in time be dissolved; or if the water contain a large quantity of free carbonic acid, the water cannot be used with safety, the compound formed being highly poisonous.

Soluble lead salts give, with sulphuric and hydrochloric acid, a white precipitate; with hydrosulphuric acid, a black precipitate.

White Lead, PbO,CO_2, or (Pb,CO_3).—White lead is manufactured by placing sheets of lead in earthenware jars containing acetic acid. These jars stand in rows upon manure or tan. The acetic acid converts the lead into an acetate, which is subsequently converted into the carbonate by the carbonic acid of the decomposing manure. The process occupies some days. White lead is extensively used in painting.

MERCURY.

Symbol, Hg; *Equivalent*, 100.

This metal, at ordinary temperatures, is a liquid; at a temperature of 40° below zero, it congeals; at a temperature of 662°, it volatilizes. Mercury has a specific gravity of 13·59.

There are two oxides of mercury, and corresponding sulphides and chlorides,—the suboxide, Hg_2O, and the protoxide, or red precipitate, HgO.

Calomel, Hg_2Cl.—This is a white crystalline body, insoluble in water. It is very useful in medicine. Calomel may be prepared by subliming a mixture of sulphate of the protoxide of mercury, free mercury, and common salt:—

$$HgO,SO_3 + Hg + NaCl = NaO,SO_3 + Hg_2Cl; \text{ or,}$$
$$(Hg,SO_4 + Hg + NaCl = Na,SO_4 + Hg_2Cl).$$

Corrosive Sublimate, $HgCl$, is of a white colour. This salt is soluble in water, to which it communicates a sweet metallic taste. With albumen, it forms a white insoluble precipitate. It is a powerful poison. White of egg is the best antidote.

Corrosive sublimate is prepared by subliming a mixture of the sulphate of the protoxide of mercury with the chloride of sodium :—

$$HgO,SO_3 + NaCl = NaO,SO_3 + HgCl;\ \text{or,}$$
$$(Hg,SO_4 + NaCl = Na,SO_4 + HgCl).$$

Mercury combines with great facility with gold, silver, zinc, tin, lead and bismuth, forming alloys called amalgams. Mercury is employed to extract gold and silver from their ores; and for gilding, plating, and looking-glasses. It is also used for thermometers and barometers.

SILVER.

Symbol, Ag; *Equivalent*, 108.

This is a very white, brilliant-looking metal, having great lustre. It is ductile and malleable; but not so much as copper. It has a specific gravity of 10·5, and melts at a bright red heat. It does not readily oxidize in air, but nitric acid readily oxidizes and dissolves it.

Soluble salts of silver are generally without colour; but on exposure to the light they blacken. Hydrochloric acid and the soluble chlorides give a white precipitate with silver solutions, soluble in ammonia.

Nitrate of silver is the lunar caustic of the surgeon.

The silver used for coin and plate in this country consists of an alloy of silver and copper, fixed by law; 111 parts of silver and 9 of copper is standard silver, and bears the Hall mark of the Goldsmiths' Company. In France three different standards are employed; one for currency, one for plate, and one for ornaments.

PLATINUM.

Symbol, Pt; *Equivalent*, 98·7.

This metal was discovered in 1752, and called white gold. Specific gravity 21. It is of a greyish colour, very difficult to fuse, and does not oxidize on exposure to air.

Platinum can be forged and welded like wrought iron. It is made into crucibles and stills for the chemist and gold refiner.

Deville has lately succeeded in fusing any quantity of platinum by employing the oxyhydrogen flame as the source of heat, in a furnace constructed of infusible non-conducting substances.

GOLD.

Symbol, Au; *Equivalent*, 197.

Gold is one of the most unalterable of the metals. It is very ductile and malleable, the thinnest gold leaf not exceeding the $\frac{1}{200000}$ of an inch in thickness. Specific gravity 19·5. It dissolves in nitro-hydrochloric acid, producing the terchloride of gold, $Au\,Cl_3$, which is thrown down as a red-purple precipitate, on adding protochloride of tin. Oxalic acid gives a brown precipitate.

Gold coinage in this country consists of an alloy in which 22 parts out of 24 are pure gold, and 2 parts of alloy. Good jewellery contains 18 parts out of the 24, and bears the mark of the Goldsmiths' Company. Jewellery with less than this proportion does not receive the mark. It not unfrequently happens that chains, rings, &c., sold as gold, do not contain more than 40 per cent. of the pure metal.

ORGANIC CHEMISTRY.

The chief constituents of organic compounds are hydrogen, oxygen, carbon, and nitrogen, occasionally associated with sulphur and phosphorus. These elements in infinitely varied proportions combine with one another and form innumerable compounds; these compounds unite and form new combinations according to the laws of definite multiple and equivalent proportions.

Complex organic bodies are of a very unstable character; by a slight increase of temperature the carbon combines with the oxygen to form carbonic acid, and the hydrogen with the nitrogen to form ammonia, and a high temperature is attended with the destruction of organic compounds. This unstable character decreases with organic bodies of simpler composition till we arrive at compounds of considerable permanence.

TYPES AND THE LAW OF SUBSTITUTION.

Most elementary and compound bodies may be referred to one or other of several different groups or classes, each class being composed of members possessing a similar constitution, and formed according to a certain pattern or *type*. Thus we may consider $NaCl$, HCl, HI, FeI, &c. to belong to the same class or group, the simplest member of which is perhaps HCl, which may therefore be taken as the type of this particular class of bodies.

There are *four* types recognized by chemists, viz.:—the *hydrogen* type, the *hydrochloric acid* type, the *water* type, and the *ammonia* type.

To the first class belong the metals and metalloids, and compound bodies acting as elements, as cyanogen (C_2N), Ethyl (C_4H_5), Phenyl ($C_{12}H_5$), &c.

To the second, or hydrochloric acid type, belong the hydrogen and metallic compounds of the chlorine group, and other bodies having a similar constitution.

The third, or water type, comprises the metallic oxides, the oxygen acids, anhydrides, oxysalts, alcohols, ethers, &c.

The fourth class includes the metallic derivatives of ammonia, amines, urea, &c.

Common salt differs from HCl merely in containing Na instead of H, in combination with Cl. If we act upon Na with HCl, the H is displaced, and the Na substituted for it, the result being common salt, NaCl. All cases of decomposition are merely examples of this *displacement* of one body, and the *substitution* for it, of another body.

But, besides this kind of substitution, there is another which is more remarkable. It is found that a great number of organic compounds containing H, may have that H displaced by Cl, Br, I, without suffering any great change in their general chemical characters. By this means one, two, three equivalents of H may be displaced by one, two, three equivalents of Cl, I, or Br; and even *all* the H may thus be removed. For example: aniline, $C_{12}H_7N$, may have one equivalent of Cl substituted for one of H, the resulting body chloraniline, $C_{12}H_6ClN$, very much resembling aniline itself. Dutch liquid, $C_4H_4Cl_2$, may have its H successively displaced by Cl, yielding the following compounds: $C_4H_3Cl_3$, $C_4H_2Cl_4$, C_4HCl_5, C_4Cl_6. Many bodies allow of their H being displaced by NO_4, as nitronaphthaline, $C_{20}H_7NO_4$; nitrobenzole, $C_{12}H_5NO_4$, binitronaphthaline, $C_{20}H_6(NO_4)_2$, &c. By this means many new bodies have of late years been discovered by chemists, some of them possessing very interesting properties.

ISOMERIC BODIES.

In organic compounds we have frequent illustration of bodies having the same chemical composition, but differing in properties; and, during the past few years, the number of these isomeric bodies has been greatly increased.

The difference of properties is ascribed to the differing arrangement of the elements. Gum ($C_{12}H_{11}O_{11}$), and cane sugar ($C_{12}H_{11}O_{11}$); starch ($C_{12}H_{10}O_{10}$), and cellulose ($C_{12}H_{10}O_{10}$), are examples of isomerism; acetate of methyle and formic ether are isomeric. The former compound is thought to be made up of acetic acid and wood-spirit ($C_4H_3O_3 + C_2H_3O = C_6H_6O_4$). While the latter compound is similarly thought to be made up of formic acid and ether, $C_2HO_3 + C_4H_5O = C_6H_6O_4$.

NON-AZOTIZED BODIES.

Starch, $C_{12}H_{10}O_{10}$.—This substance is present in almost all vegetables; it is most abundant in the cereals, peas, beans, lentils, turnips, carrots, and potatoes. Tapioca, sago, and arrowroot are varieties of starch extensively used as articles of diet. Starch can be obtained by kneading under a stream of water on a sieve the pulp of potatoes. The starch passes through with the water and settles at the bottom of the vessel in the form of a white insoluble powder. When this process is applied to the meal of any of the cereals a tough, elastic substance remains, which is gluten, and the value of wheat and other grain as articles of food depends greatly on the quantity of gluten. Pure flour contains about ten per cent., and bran contains about fifteen per cent. When starch is mixed with water and heated to the boiling point, the granules burst and form a gelatinous mass, which when dried in thin layers forms a yellow horny substance, which again dissolves in boiling water. This is called *amidin*.

Iodine gives with very weak solutions of starch a deep characteristic blue colour. Under the microscope starch granules vary in form according to the vegetable substance from which they were obtained.

If a solution of starch be boiled with dilute sulphuric acid it is converted into *dextrin*, a substance identical in composition with starch; and if the boiling be continued for some time the dextrin is converted into glucose or grape sugar.

Cellulose.—True cellulose is the material which encrusts the interior of the cells of vegetable substances; it has the same composition as starch.

Lignin is the basis of all woody fibre, and one of its purest forms is well-bleached flax.

Of the several varieties of sugar the most important are cane sugar, glucose or grape sugar, diabetic sugar, sugar of milk, and amorphous sugar.

Cane Sugar, $C_{12}H_{11}O_{11}$.—This is obtained from the juice of the sugar-cane, and on the Continent it is largely manufactured from beetroot; and in the western parts of America it is obtained from the sap of the sugar maple. It is easily crystallized, and is very soluble in hot water, but insoluble in alcohol and ether. When heated it melts into a black syrup, which solidifies on cooling into an amorphous mass known as barley sugar. When cane sugar is boiled with dilute sulphuric acid it is converted into grape sugar.

Sugar, as an article of diet, is not nutritious; its chief value depends on the support afforded by the carbon in keeping up the animal heat by respiration.

Caramel, a dark brown substance, obtained from cane sugar, is used for colouring spirits. It is $C_{12}H_9O_9$.

Grape Sugar, *Glucose*, *Diabetic Sugar*, $C_{12}H_{14}O_{14}$.—This is the sweet principle of grapes, raisins, and most fruits, and a product, as we have seen, of the metamorphosis of starch and cane sugar. It may also be obtained from milk sugar and woody fibre when either of these bodies is boiled for some time with dilute sulphuric acid; the solution is then neutralized with chalk, filtered, and evaporated. The theory of the change is, that starch, $C_{12}H_{10}O_{10}$, takes up four equivalents of water, and produces $C_{12}H_{14}O_{14}$, or glucose. By varying the quantity of water the same theory may be applied to the other bodies. Glucose is less soluble in water than cane sugar, and not so sweet. It is found in considerable quantities in the urine of persons afflicted with diabetes.

Sugar of Milk, $C_{12}H_{12}O_{12}$.—This is obtained by evaporating the whey of milk till it crystallizes. When pure it forms hard white crystals, soluble in water, but insoluble in ether or alcohol. The Tartars make an intoxicating liquor by the fermentation of mares' milk. Sugar of milk is used in preparing the little globules of homœopathic medicines.

Amorphous Sugar, $C_{12}H_{12}O_{12}$.—This may be obtained from the juices of vegetables and honey.

In the bodies previously described we have several cases of isomerism, and the oxygen and hydrogen usually exist in the proportions necessary to form water. The chief value of these substances as articles of food depends on their supplying those elements of respiration by which the animal heat is maintained.

AZOTIZED BODIES.

Animal Albumen.—The serum of the blood and the white of egg are strong solutions of animal albumen. In its natural state it is soluble in cold water, and is easily coagulated by heat; it is also coagulated by acids, the ferrocyanide of potassium and several metallic salts. Albumen in its purest form always contains small quantities of phosphorus, sulphur, and alkaline phosphates. When dried it becomes a yellow, transparent, horny, and brittle substance, insoluble in cold water, but slightly soluble in boiling water. One hundred parts of albumen have the following composition:—

Carbon, 53·5; hydrogen, 7·0; nitrogen, 15·5; oxygen, 22·0; phosphorus, 0·4; sulphur, 1·6 = 100·0.

Vegetable Albumen.—The albuminous substances of the vegetable kingdom are regarded as the nutriment provided by nature for the support of animals. They have nearly the same composition as animal albumen, and are regarded as the constituents out of which animal fibre and tissue are built up, and are called the plastic elements of nutrition, to distinguish them from the respiratory constituents of food.

Animal Fibrin.—This substance may be obtained by stirring fresh blood with a bundle of twigs, and well washing in cold water, to remove the blood globules. It forms long, white, thread-like filaments, which attach themselves to the twigs. It is insoluble in hot and cold water. The muscles of animal bodies are composed of similar fibres, but not quite identical in composition with the fibrin obtained from blood. The soluble condition in which fibrin exists in the blood cannot be preserved; it is only known in its insoluble form.

Vegetable Fibrin.—This substance is found in the juices of plants; it is also found in the seeds of the cereals. Gluten obtained from wheat flour contains a considerable proportion of vegetable fibrin, associated with a nitrogenous body called *gliadin.* This substance is not present in the meal of oats and barley; these yield no gluten, but filaments of pure vegetable fibrin.

Animal Caseine.—This substance is the coagulable portion of milk, and when prepared from well-skimmed milk is nearly pure. When caseine is taken in the form of food it yields albumen and fibrin, and from this substance the blood and muscle of young animals is formed. The decomposition of caseine is attended with the same products as the decomposition of albumen and fibrin.

Vegetable Caseine.—To distinguish this substance from animal caseine it is sometimes called *legumine.* This substance exists in almost all vegetables used as food, but in the largest proportion in lentils, peas, beans, and other leguminous seeds. A vegetable cheese is made by the Chinese from the caseine obtained from peas.

Gelatine.—This substance is found in the skin, bones, tendons, horns, and hoofs of animals: it dissolves in hot water, forming glue, but it may be so purified as to resemble isinglass, and be used for jellies and soups. Dilute solutions of gelatine are precipitated by tannic acid, forming a kind of tannate of gelatine, or leather. It is largely employed in the arts; but as a nutritive article of diet its value has been greatly over-estimated. Isinglass is nearly pure gelatine; it is obtained from the swimming-bladder of the sturgeon; it dissolves in warm water, and is sometimes used by brewers and wine-merchants to clarify their liquors.

The azotized bodies previously described are analogous in their chemical composition. Vegetable albumen, fibrin, and caseine are almost identical with the same substances obtained from blood or milk. It is from these bodies that the waste of muscular tissue and fibre is repaired.

FERMENTATION.

When a solution of grape sugar, to which is added a little yeast, or the saccharine juice of any plant, is exposed to a temperature of about 70°, carbonic acid rises to the surface, and in time the sweet taste disappears, and the solution is found to contain *alcohol.* The change may be expressed as follows:—Grape sugar contains $C_{12}H_{14}O_{14}$, from which we obtain two equivalents of alcohol, four of carbonic acid, and two of water,—

$$\begin{array}{lcl} 2(C_4\,H_6\,O_2) & = & C_8\,H_{12}\,O_4 \\ 4\,C\,O_2 & = & C_4\quad\;\; O_8 \\ 2\,H\,O & = & \quad\;\; H_2\,O_2 \\ \hline & & C_{12}\,H_{14}\,O_{14}. \end{array}$$

The yeast or ferment which is added takes no part in the change, but acts merely by inducing it, and this change is called the vinous fermentation. An infusion of malt or the saccharine juice of fruits contains, besides sugar, some body containing nitrogen, which acts as a ferment; but a pure solution of sugar will not ferment without the addition of a little yeast.

We have referred in a former page to compound organic radicals, that is, compound bodies which enter into combination with, and behave like, elementary bodies, and perform, in fact, the part of elements. Some organic compounds act as metals, and form salts by combining with chlorine, iodine, bromine, sulphur, and cyanogen; they also combine with oxygen, and form compounds, having basic properties analogous to the metallic oxides. The following are examples:—

Basic Radicals.		Oxygen Compounds.	Chlorine Compounds.
Ethyle .	. $C_4\ H_5 = Ae$	AeO	AeCl
Methyle.	. $C_2\ H_3 = Mt$	MtO	MtCl
Amyl .	. $C_{10}\ H_{11} = Ayl$	AylO	AylCl

Ether, or *oxide of ethyle* ($C_4H_5O = AeO$), is obtained artificially by gently heating alcohol with sulphuric acid, which removes an equivalent of water from the alcohol, which has the following composition:—$C_4H_6O_2$; subtracting one equivalent of water we have C_4H_5O. The chloride, bromide, and iodide of ethyle may be prepared by saturating alcohol with the vapour of either hydrochloric, hydrobromic, or hydriodic acid gas, distilling the mixture in a vapour bath, and collecting in a receiver kept cool with iced water.

VEGETABLE ACIDS.

Acetic Acid, $C_4H_3O_3 + HO = AcO_3,HO$ ($Ac = C_4H_3$ = Acetyle).—This acid may be formed either by the oxidation of alcohol or the distillation of wood (pyroligneous acid). If fermented liquors are exposed to the atmosphere the alcohol they contain undergoes a change by the absorption of oxygen; and if the exposure be continued a sufficient length of time, the alcohol is changed into acetic acid, or vinegar. The change may be represented as follows:—

$$C_4H_6O_2 + O_4 = C_4H_3O_3,3HO = AcO_3, HO + 2HO.$$

Tartaric Acid, $C_8H_4O_{10},2HO = \bar{\bar{T}},2HO$. This is the acid principle of grapes and several other fruits. The tartaric acid of commerce is obtained from argol, an impure tartrate of potash, which deposits itself as a crust in the interior of wine casks. This acid is employed in calico-printing; it is used to evolve the chloine from bleaching-powder.

Citric Acid, $C_{12}H_5O_{11}3,HO = \bar{Ci},3HO$.—This is the acid principle of the lemon, lime, orange, and red currant. It is usually prepared by neutralizing lemon juice with chalk, by which process an insoluble citrate of lime is formed; this is decomposed with dilute sulphuric acid. The solution is then heated, filtered, and, on evaporation, crystals of citric acid are deposited.

Malic Acid, $C_8O_4H_8,2HO = \bar{M},2HO$. This is the acid principle of unripe apples and rhubarb. It is usually prepared from the berries of the mountain ash.

Oxalic Acid, C_2O_3, HO, or $HC_2O_4 = HO,\bar{O}$.—This acid is usually found in the vegetable juices of several plants. It exists in the form of an oxalate of potash. It was formerly obtained from a plant known as the oxalis acetosella, but it is now manufactured as an article of commerce by the action of nitric acid on sugar, starch, and some other organic compounds. This acid is very poisonous, and has been the cause of many fatal accidents from its resemblance to Epsom salts. It is, however, easily distinguished from these salts by its sour taste. The best antidote is chalk administered in water; this forms an inert oxalate of lime.

Tannic Acid.—This acid exists in almost all vegetables: it is, however, most abundant in oak bark, and in the fruit and stems of the sumach, and especially in nut galls, which are excrescences produced on the branches and leaves of the oak by the puncture of insects. Tea contains from 8 to 10 per cent. of tannic acid, to which it imparts a strong astringent quality. Tannic acid unites with the gelatine or skins of animals, and forms an insoluble compound; and this is the principle practically applied in the manufacture of leather, and which is effected by steeping the skins for a long time in an aqueous solution of oak bark.

VEGETABLE BASES.

Morphia, Morphine.—Opium is the dried juice of a certain species of poppy, and morphine is its chief active principle. It exists in opium in combination with meconic acid; it is powerfully narcotic and poisonous. It crystallizes in small colourless prisms, which have the following composition:—$C_{34} H_{19} NO_6 + 2HO$.

Quinine and Cinchonine.—The medicinal properties of Peruvian bark are due to these substances. Quinine is a white crystallized body of an intensely bitter taste. It is a valuable medicinal agent, and is administered in the form of a sulphate; a peculiar acid, known as kinic acid, and sulphuric acid, exist in combination with these alkalies in Peruvian bark. Cinchonine and quinine are analogous in composition: $C_{40} H_{24} N_2 O_2$.

Strychnia and Brucia.—These substances are obtained from a variety of plants belonging to the genus strychnos, and especially from the berries of the nux vomica. They are among the most powerful of vegetable poisons; their influence seems to be exerted principally on the nerves and spinal cord, producing violent spasms, which increase in intensity until death. According to Messrs. Nicholson and Abel, strychnine has the following composition, $C_{44} H_{22} N_2 O_4$. Among other important vegetable alkalies may be mentioned *nicotine*, the poisonous principle of tobacco; *aconitine*, extracted from monkshood; *coneine*, prepared from hemlock; *veratrine*, from the hellebore; *hyoscyamine*, from henbane; all these are violent poisons, scarcely inferior to strychnine. Among the less injurious vegetable alkalies may be mentioned *emetine*, the medicinal principle of ipecacuanha; *piperine*, extracted from black pepper; *caffeine* and *theine*, the active principle of coffee and tea.

OILS AND FATS.

Oils and fats, whether of animal or vegetable origin, belong to the same general class of organic compounds, and with the exception of the volatile oils, are analogous in their chemical properties. The general distinction between oils and fats is, that the former are liquids at ordinary temperatures, and the latter are solids. The natural fixed oils and fats are chiefly neutral compounds formed by the union of either oleic, margaric, or stearic acid with glycerine. Thus oleine, margarine, and stearine are compounds of glycerine with a fatty acid. These fats and oils are never found by themselves. Suet consists of stearine, oleine, and margarine; lard and human fat, of margarine and oleine. They are all saponified, by boiling with potash or soda, which unite with the fatty acids to form soaps, while the glycerine unites with the water, and is carried off as a waste product. The cleansing properties of soap are due to the alkaline constituents.

An alkali used alone would act more powerfully as a cleansing agent than soap, but it would tend to the destruction of the fibre and the removal of the colour of dyed fabrics.

Washing powders are merely caustic alkalies; they facilitate washing by providing an excess of alkali, and precipitating the lime present in hard water, but in excess they tend to injure the fabrics, and occasion them to shrink. These powders ought never to be used for blankets or woollen articles.

ORGANIC COLOURING MATTERS.

With the exception of cochineal, all the organic colouring principles are of vegetable origin.

Blue.—The most important of the blue dyes is indigo, which is obtained from several species of American and Asiatic plants. The juice of these plants is colourless, but when their leaves are digested in water and allowed to ferment, a yellow substance is dissolved out, which on exposure to the air becomes blue, and is deposited in the form of a thick sediment. This washed and dried forms the indigo of commerce, which is usually very impure, containing more than half its weight of foreign substances.

It is insoluble in every liquid, except sulphuric acid, with which it forms a compound soluble in water.

Among other important blue dyes are litmus, archil, and cudbear.

Red Madder is the root of the plant *rubia tinctorum.* Its most beautiful colouring constituents are madder red, madder purple, and orange. Turkey red is not a natural product, but is obtained by the action of sulphuric acid on the root. Other substances of this class are Brazil wood, safflower, sandal-wood, and cochineal. The last is a dried species of insect which lives upon the cacti of South America and other warm countries.

Yellow.—The most valuable dyes of this class are fustic, and turmeric, the root of an East Indian plant.

Green.—Chlorophyle is the name given to the green colouring principle of plants. It exists in very minute quantities, and is obtained in a state of purity with great difficulty.

CYANOGEN.

Perhaps the best illustration of a compound radical is cyanogen, which is a compound of carbon and nitrogen, $C_2N = Cy$; and as this compound acts exactly like a metal, it is represented by the symbol Cy.

When animal refuse is heated with carbonate of potash in closed iron vessels, cyanogen gas is evolved, which enters into combination at a high temperature with the potassium of the carbonate of potash, forming a salt known as the cyanide of potassium (KCy). This salt is very unstable, and is decomposed by water; it is therefore necessary to convert it into a more stable compound, which is effected by the addition of sulphide of iron. The iron is dissolved by the aqueous solution of cyanide of potassium, forming cyanide of iron ($FeCy$), and sulphide of potassium (KS). The elements of the cyanide of iron unite with the cyanide of potassium, and form a very permanent salt known as the yellow prussiate of potash, or the ferrocyanide of potassium, $K_2FeCy_3 + 3HO$ or $K_2FeC_6N_3 + 3HO$. When a solution of this salt is added to a solution of any salt of the sesquioxide of iron, Prussian blue is formed.

From the ferrocyanide of potassium all the other compounds of cyanogen and cyanogen itself may be prepared; but the best method of preparing cyanogen gas is by heating, in a small retort or test tube, the dried cyanide of mercury, and collecting the gas over mercury. It is colourless, and has the peculiar odour of peach kernels, and burns with a beautiful purple flame fringed with yellow; at a pressure of four atmospheres it is converted into a liquid.

With hydrogen, cyanogen forms hydrocyanic acid, HCy; with oxygen and water it forms the following compounds: CyO,HO. $Cy_2O_2, 2HO$. $Cy_3O_3, 3HO$.

Cyanogen also combines with chlorine, and with sulphur it forms a new radical, which has the composition of

CyS_2 *sulphocyanogen;* with the metals it forms salts, analogous to chlorides (KCy. $FeCy$. $AgCy$., &c. KCl. $FeCl$. $AgCl$, &c.), and acts precisely the same part in forming compounds as a simple radical.

Hydrocyanic Acid, HCy, prussic acid.—This acid is remarkable for its poisonous properties. It may be obtained by distilling the cyanide of potassium with dilute sulphuric acid, $KCy + HSO_4 = KSO_4 + HCy$. In its pure state it is a colourless transparent liquid, having a very feeble acid reaction; it is extremely volatile, and if a drop be allowed to fall on a piece of glass, a part of the acid will be frozen by its evaporation.

Cyanic Acid, CyO, HO, is a highly volatile liquid; it is difficult to keep in its unaltered condition. The salts are called cyanates.

Fulminic Acid, $Cy_2O_2, 2HO = Cy_2O_4, H_2$.—This acid is not known in its separate state. In combination with the metals it forms a class of salts known as fulminates, since they easily explode. The fulminate of mercury, $2Hg_2O, C_4N_2O_2$, is a most dangerous compound; it explodes violently with the slightest friction, and sometimes spontaneously. The detonating shells thrown under the carriage of Napoleon III., in 1858, were filled with this salt. It is used with powder in the manufacture of percussion-caps.

QUESTIONS *given at the Examination for Certificates in Chemistry, held at the South Kensington Museum in November* 1859.

[Three hours allowed for this Paper.]

INORGANIC CHEMISTRY.

SECTION I.

1. Describe the usual processes for the preparation of oxygen gas.
2. Give an outline of the chemical processes involved in the manufacture of sulphuric acid.
3. Describe the action of sulphur upon the alkaline carbonates at a high temperature, and the deportment of the resulting products under the influence of atmospheric air.

SECTION II.

1. Give an outline of the process of combustion.
2. Describe the manufacture of carbonate of soda.
3. A salt has been found to contain in 100 parts—

17·1	Calcium.
26·5	Phosphorus.
1·7	Hydrogen.
54·7	Oxygen.
100	Salt.

What is the formula of the salt? Give the details of calculation.

The following equivalents to be used:—

Calcium	20
Phosphorus	31
Hydrogen	1
Oxygen	8

Solution.

Dividing the quantity of each element by its equivalent, we obtain numbers expressing the proportion of the equivalents in the compound:—

$$\frac{17{\cdot}1}{20} = {\cdot}855 \text{ eqs. of Ca.}$$

$$\frac{26{\cdot}5}{31} = {\cdot}855 \quad \text{,,} \quad \text{P.}$$

$$\frac{1{\cdot}7}{1} = 1{\cdot}7 \quad \text{,,} \quad \text{H.}$$

$$\frac{54{\cdot}7}{8} = 6{\cdot}837 \quad \text{,,} \quad \text{O.}$$

Dividing each of these quotients by the number of equivalents of H (1·7) we obtained the number of equivalents of each element, united with 1 eq. of H.

$$\frac{{\cdot}855}{1{\cdot}7} = {\cdot}5 \text{ eqs. of Ca.}$$

$$\frac{{\cdot}855}{1{\cdot}7} = {\cdot}5 \quad \text{,,} \quad \text{P.}$$

$$\frac{1{\cdot}7}{1{\cdot}7} = 1 \quad \text{,,} \quad \text{H.}$$

$$\frac{6{\cdot}837}{1{\cdot}7} = 4 \quad \text{,,} \quad \text{O.}$$

Doubling these quotients to get rid of fractional parts of equivalents, we find the salt to be composed of—

1 eq. Ca; 1 eq. P; 2 eqs. H; 8 eqs. O;

and is therefore, $CaPH_2O_8$ = $CaO,2HO,PO_5$, or CaH_2PO_8.

The above operations may be combined in one, as follows:—

$$\frac{17{\cdot}1}{20 \times 1{\cdot}7} = \frac{17{\cdot}1}{34} = {\cdot}5 \text{ Ca or } Ca_1.$$

$$\frac{26\cdot5}{31 \times 1\cdot7} = \frac{26\cdot5}{52\cdot7} = \cdot5 \text{ P or } P_1.$$

$$\frac{1\cdot7}{1 \times 1\cdot7} = \frac{1\cdot7}{1\cdot7} = 1 \text{ H or } H_2.$$

$$\frac{54\cdot7}{8 \times 1\cdot7} = \frac{54\cdot7}{13\cdot6} = 4 \text{ O or } O_8.$$

SECTION III.

1. Describe the method of separating baryta, strontia, and lime.
2. Give the formula of ammonia, its percentage composition, by weight and by volume.
3. Give the chemical history of the phosphoric acids, and their compounds.

SECTION IV.

1. How is magnesia separated from the alkalies?
2. What is meant by the term *hardness* when applied to water? Distinguish the different kinds of hardness in water, and describe a practical method of determining it.
3. Give a brief outline of the chemical history of chlorine and its principal compounds.

PRACTICAL EXAMINATION IN ANALYSIS.

Seven hours allowed.

1. Qualitative analysis of a mixture consisting of two acids and two bases soluble in water or acids.
2. Qualitative analysis of a mixture consisting of two acids and two bases partly or entirely insoluble in water and acids.
3. Qualitative analysis of more complicated mixtures.

ORGANIC CHEMISTRY.

Three hours allowed for this Paper.

SECTION I.

1. Describe the process for preparing cyanogen.
2. Describe the method of preparing hydrocyanic acid, and the transformation which it undergoes in the presence of acids and alkalies.
3. Give an outline of the chemical history of the ferrocyanides and ferricyanides. State their use in distinguishing the two classes of iron salts.

SECTION II.

1. State the formula of cane sugar and glucose (grape sugar), and give the equations representing their change during the process of fermentation.
2. Describe the successive changes which starch undergoes in its transformation, first into alcohol, and lastly into acetic acid.
3. Give a sketch of the alcohols and homologues, of the aldehydes, and of the acids corresponding to them.

SECTION III.

1. Describe the process of etherification.
2. Give a short sketch of the fatty bodies, and of the process of saponification.
3. Describe the relation in which alcohols, either simple or compound, and anhydrides, simple and compound, stand to water. Illustrate each of these groups by a formula.

SECTION IV.

1. Describe one of the processes for determining the equivalent of an organic acid.
2. Give an outline of the method of determining the vapour density of an organic body.
3. An organic base has been found to consist of—

Carbon	77·42
Hydrogen . . .	7·53
Nitrogen	15·05
	100·00

The platinum salt of the base contains 32·99 per cent. of platinum.

What is the formula of the base? Give the details of calculation.

The following equivalents to be used:—

Carbon	6·0
Hydrogen	1·0
Nitrogen	14·0
Chlorine	35·5
Platinum	98·7

Solution.

Divide each quantity by each equivalent respectively:—

$$\frac{77\cdot42}{6} = 12\cdot903 \text{ eqs. C.}$$

$$\frac{7\cdot53}{1} = 7\cdot53 \text{ eqs. H.}$$

$$\frac{15\cdot05}{14} = 1\cdot075 \text{ eqs. N.}$$

Dividing these quotients by the number of eqs. of N., (because this is the element which has the lowest number of equivalents in the compound) we obtain—

$$\frac{12\cdot903}{1\cdot075} = 12 \text{ eqs. of C.}$$

$$\frac{7\cdot53}{1\cdot075} = 7 \text{ ,, H.}$$

$$\frac{1\cdot075}{1\cdot075} = 1 \text{ ,, N.}$$

The *empirical* formula is therefore, $C_{12}\ H_7\ N$.

The *rational* formula, or the formula expressing one *equivalent* of the compound depends on the quantity required to form the double chloride with platinum, and will either coincide with the above or be some multiple of it.

Formulæ of platinum salts:—

Double chloride of potassium and platinum = $KCl,PtCl_2$

,, ,, ammonium ,, = $NH_4Cl,PtCl_2$
or $NH_3HCl,PtCl_2$

Now the compound formed by the base in question with HCl will be, $x(C_{12}H_7N),HCl$, and the double salt with platinum, $x(C_{12}H_7N),HCl,PtCl_2$. Now in 100 parts of this compound there are 32·99 of platinum; it is easy to find the equivalent, therefore, of the platinum salt by the following proportion:—

$$\overset{\text{Pt.}}{32\cdot99} : \overset{\text{Pt.}}{98\cdot7} :: 100 : 299\cdot2$$

The 299·2 parts contain 1 eq. $Pt.Cl_2$, 1 eq. HCl, and and 1 eq. of the base.

$$\begin{array}{lr} \text{Eq. } PtCl_2 = 98\cdot7 + 71 = & 169\cdot7 \\ \text{Eq. HCl} = & 36\cdot5 \\ \hline & 206\cdot2 \end{array}$$

Deducting this from 299·2, we obtain the equivalent of the base, viz. 299·2 − 206·2 = 93.

This corresponds with the empirical formula previously found, for—

$$\begin{array}{lcr} C_{12} & = & 72 \\ H_7 & = & 7 \\ N & = & \underline{14} \\ & & 93 \end{array}$$

And therefore the *empirical* formula is also here the *rational* formula, viz.—

$$C_{12}H_7N, \text{ or } \left.\begin{array}{r} H \\ H \\ C_{12}H_5 \end{array}\right\} N = \textit{phenylamine}, \text{ or } \textit{aniline}.$$

QUESTIONS *given at the Examination for Certificates in Chemistry, held at the South Kensington Museum, in November* 1860.

[Three hours allowed for this Paper.]

INORGANIC CHEMISTRY.

SECTION I.

1. Describe the preparation, properties, and the composition by weight and volume, of nitric oxide.
2. Give an outline of the manufacture of coal gas.
3. Give the composition of the atmosphere, and describe the methods employed in determining its constituents.

SECTION II.

1. State the composition of gunpowder, and explain the reaction which takes place during its combustion.
2. Give an outline of the theory of bleaching.
3. A mineral has been found to contain in 100 parts :—

Sodium . . .	8·73
Aluminium . . .	10·39
Silicium . . .	32·32
Oxygen . . .	48·56
	100·00

What is the formula of the mineral? The following equivalents to be used :—

Hydrogen . . .	1
Oxygen . . .	8
Silicium . . .	21·3
Sodium . . .	23
Aluminium . . .	13·7

Give the details of calculation.

Solution.

Divide the quantity of each element by its equivalent:—

$$\frac{8\cdot73}{23} = \cdot38 \text{ eqs. of Ca.}$$

$$\frac{10\cdot39}{13\cdot7} = \cdot758 \text{ ,, Al.}$$

$$\frac{32\cdot32}{21\cdot3} = 1\cdot517 \text{ ,, Si.}$$

$$\frac{48\cdot56}{8} = 6\cdot07 \text{ ,, O.}$$

Dividing each of these results by ·38 (the number of eqs. of Ca) we obtain

$$\frac{\cdot38}{\cdot38} = 1 \text{ eq. Ca.} \quad \frac{\cdot758}{\cdot38} = 2 \text{ eqs. Al.}$$

$$\frac{1\cdot517}{\cdot38} = 4 \text{ eqs. Si.} \quad \frac{60\cdot7}{\cdot38} = 16 \text{ eqs. O.}$$

Therefore the compound contains $Ca\,Al_2\,Si_4\,O_{16}$, and its composition is $CaO, SiO_2 + Al_2O_3, 3SiO_2$.

SECTION III.

1. Give the composition and the preparation of chloric and perchloric acids.
2. Describe the course of analysis followed in cases of poisoning by arsenic.
3. Give an outline of the doctrine of chemical equivalents.

SECTION IV.

1. How is fluorine detected in the presence of silicates?

2. Give an outline of the manufacture of white lead.
3. Name the more frequently occurring mineral compounds which are insoluble both in water and acids, and describe the process of rendering them soluble for analysis.

PRACTICAL EXAMINATION IN ANALYSIS.

1. Qualitative analysis of a mixture consisting of two acids and two bases soluble in water or acids.
2. Qualitative analysis of a mixture consisting of two acids and two bases partly or entirely insoluble in water and acids.
3. Qualitative analysis of more complex mixtures.

ORGANIC CHEMISTRY.

[Three hours allowed for this Paper.]

SECTION I.

1. Describe the several methods of ascertaining the presence of hydrocyanic acid.
2. Give the method of preparing fulminate of mercury, and explain the constitution of this compound.
3. Give an outline of the chemical processes involved in the nutrition of plants.

SECTION II.

1. Describe the method of preparing iodide of ethyl.
2. Give an outline of the preparation, properties, and decompositions of urea.
3. Describe Gay Lussac's method of determining vapour densities, and develop the general formula:

D = Density of vapour.
p = Amount of substance employed.
V = Volume of vapour observed in cubic centimetres.
t° = Temperature observed.
B = Barometer normal.
b = Barometer observed.
h = Suspended mercury-column.
a = Weight of 1 cubic centimetre of air.
e = Coefficient of expansion of air.

Solution.

It is not intended to enter into the details of Gay Lussac's method of ascertaining the density of the vapour of a volatile liquid; but, in order to render the following formula intelligible, a brief outline of the process is necessary.

A graduated glass receiver is filled with mercury and inverted in a vessel containing the same fluid. A few small and very thin glass globes, filled with a known weight of the volatile liquid, are passed through the mercury to the top of the receiver. By appropriate means, heat is applied to the glass receiver, when the glass globules burst, and the volatile liquid becomes converted into vapour, which by its elasticity expels a portion of the mercury from the receiver. The volume of the vapour is then read off, and the height of the mercury column in the glass receiver also observed. It is evident that as the pressure exerted by the vapour, together with the column of mercury in the receiver, balances the atmospheric pressure, that therefore the true pressure exerted by the vapour alone is equal to the atmospheric pressure, *minus* the height of the suspended mercury. By noting the temperature at which the liquid was converted into vapour, and also the barometric pressure at the time of the experiment, the density of the vapour may be ascertained according to the following formula:—

Weight of 1 c. c. vapour, at observed temperature and pressure $= \frac{p}{V}$.

Ratio of weights of equal bulks of vapour at t° and 0° $= 1 : 1 + et$.

$\therefore$ Weight of 1 c. c. vapour at 0° =

$$1 : 1 + et :: \frac{p}{V} = \frac{p(1 + et)}{V}$$

True pressure of vapour $= b - h$.

The weights of equal volumes of vapour are to one another as their pressures.

$\therefore$ Weight of 1 c. c. vapour at 0° and barometer B

$$= (b - h) : B :: \frac{p(1 + et)}{V}$$

$$= \frac{B}{b - h} \times \frac{p(1 + et)}{V}$$

Weight of 1. c. c. air at 0° and bar. B = a.

$$\therefore D = \frac{B.\,p.\,(1 + et)}{V.\,a.\,(b - h)} \text{ or } (1 + et)\frac{B}{b - h} \cdot \frac{p}{V.\,a}$$

Or, the same result may be obtained as follows:—

$$\text{Volume of vapour at } 0^\circ = 1 + et : 1 :: V = \frac{V}{1 + et}$$

Correcting this for barometric pressure—

True pressure of vapour = b − h.

$$\text{Volume of vapour at } 0^\circ \text{ and bar. } B = B : (b - h) :: \frac{V}{1 + et} = \frac{V\,(b - h)}{B\,(1 + et)}$$

$$\text{Weight of 1. c. c.} = p \div \frac{V\,(b - h)}{B\,(1 + et)} = \frac{p.B.\,(1 + et)}{V\,(b - h)}$$

Weight of 1. c. c. air at 0° and bar. B. = a.

$$\therefore D = \frac{B.\,p.\,(1 + et)}{V.\,a.\,(b - h)} \text{ as before.}$$

SECTION III.

1. Give a sketch of the chief processes by which acetic acid is obtained.
2. Describe the products which are formed by the action of sulphuric acid upon alcohol.
3. Give examples of diatomic and triatomic alcohols; describe their preparation and their chemical characters.

SECTION IV.

1. Give an outline of the process of manufacturing stearic candles.
2. Describe the several methods of determining the nitrogen in organic compounds.
3. Give an outline of the classification of the derivatives of ammonia.

QUESTIONS *given at the Examination of Pupils in Science Classes and Schools, in May* 1861.

INORGANIC CHEMISTRY.

GENERAL INSTRUCTIONS.

Three hours are allowed for this paper.

You are only permitted to attempt eight questions.

You may select your questions from Series I., or from Series II., or from both.

The value attached to the correct answer of each question in Series I. is 8; and those in Series II. is 12·5.

N.B.—A full and exact answer will in all cases gain more marks than an inexact or incomplete answer; though in the former case the question may be the more easy of the two and have less marks attached to it.

SERIES I.

1. Describe the preparation of oxygen by means of binoxide of manganese and sulphuric acid. Represent the reaction by a diagram.
2. What is the difference between a chemical compound and a mechanical mixture?
3. Describe the preparation and properties of ammonia.
4. What takes place when a current of chlorine is passed through a solution of potash?
5. How is sulphuretted hydrogen prepared? What is the action of this gas upon a solution of nitrate of lead?
6. Give a definition of the terms *acid*, *base*, and *salt*.
7. State the constituents of the atmosphere and the proportion by volume in which they are present.
8. What is the meaning of the term *hardness*, applied to water?

9. What is meant by the term alum? Give the formulæ of some alums.
10. Explain the action of sulphuretted hydrogen upon a solution of sesquichloride of iron.
11. What is the chemical difference between the several varieties of iron known as cast iron, wrought iron, and steel?
12. Describe the process of making lime, and explain the difference between quicklime and slaked lime.

SERIES II.

13. A compound of carbon and oxygen has been found to contain in 100 parts 27·3 of carbon and 72·7 of oxygen. What is the chemical formula of the compound?

Combining number of carbon = 6
Combining number of oxygen = 8

Give details of calculation.

Solution.

Dividing these numbers respectively by the eqs. of C and O, we obtain $\frac{27\cdot3}{6} = 4\cdot55$ eqs. of C, and $\frac{72\cdot7}{8} = 9\cdot1$ eqs. of O. These reduced become 1 eq. C and 2 eqs. O, or CO_2.

14. The formula of nitric acid is HNO_6. What is the composition of nitric acid in 100 parts?

Combining number of hydrogen = 1
Combining number of oxygen = 8
Combining number of nitrogen = 14

Give details of calculation.

Solution.

Equivalent of $HNO_6 = 1 + 14 + 48 = 63$; *i. e.*, 63 parts by weight of HNO_6, contain 1 part of H, 14 of

N, and 48 of O. The percentage composition of HNO_6 may then be found by the following simple proportion :—

$$\begin{array}{ccccccc} & & & \text{H.} & & & \\ 63 & : & 100 & :: & 1 & : & 1{\cdot}59, \text{ or } 1{\cdot}6 = 1\tfrac{3}{5}, \text{ H.} \\ & & & & \text{N.} & & \\ & & & :: & 14 & : & 22{\cdot}22, \text{ or } 22{\cdot}2 = 22\tfrac{1}{5}, \text{ N.} \\ & & & & \text{O.} & & \\ & & & :: & 48 & : & \underline{76{\cdot}19}, \text{ or } 76{\cdot}2 = 76\tfrac{1}{5}, \text{ O.} \\ & & & & & & 100{\cdot}00 \end{array}$$

15. State the principal groups of the chemical elements and the characteristic terms of these groups.
16. Describe a chemical process by which hydrogen, and one by which oxygen, may be prepared from water.
17. State the composition of the several oxides of iron, and of some of the salts which they are capable of forming.
18. State the composition of nitrous and nitric oxides by weight and volume.
19. Describe some of the processes which are employed for separating the metals from their oxides.
20. Give the theory of the processes of bleaching by exposure to air and by treatment with chloride of lime.
21. One volume of chlorine unites with 1 volume of hydrogen to hydrochloric acid, 1 volume of oxygen unites with 2 volumes of hydrogen to water, 1 volume of nitrogen unites with 3 volumes of hydrogen to ammonia. State the volumes of hydrochloric acid, water-vapour, and ammonia thus produced, and the specific gravities of these three compounds.

Specific gravity of hydrogen = 1
Specific gravity of oxygen = 16
Specific gravity of nitrogen = 14
Specific gravity of chlorine = 35·5

Solution.

1. 1 vol. H unites with 1 vol. Cl without condensation, producing, therefore, 2 vols. HCl.

$$\begin{array}{rl} \text{Weight of 1 vol. H} &= 1 \\ \text{Weight of 1 vol. Cl} &= \underline{35\cdot5} \\ \text{Weight of 2 vols. HCl} &= 36\cdot5 \end{array}$$

$\therefore$ Sp. gr. or weight of 1 vol. HCl $= \frac{1}{2}$ of $36\cdot5 = 18\cdot25$.

2. 1 vol. O unites with 2 vols. H, condensing $\frac{1}{3}$, so as to form, therefore, 2 vols. steam.

$$\begin{array}{rl} \text{Weight of 2 vols. H} &= 2 \\ \text{Weight of 1 vol. O} &= \underline{16} \\ \text{Weight of 2 vols. HO} &= 18 \end{array}$$

$$\therefore \text{Sp. gr. water vap.} = \frac{18}{2} = 9.$$

3. 1 vol. N and 3 vols. H unite and condense $\frac{1}{2}$, so as to form 2 vols. NH_3.

$$\begin{array}{rl} \text{Weight of 1 vol. N} &= 14 \\ \text{Weight of 3 vols. H} &= \underline{3} \\ \text{Weight of 2 vols. } NH_3 &= 17 \end{array}$$

$$\therefore \text{Sp. gr. } NH_3 = \frac{17}{2} = 8\cdot5.$$

22. Describe the composition of the phosphates as illustrations of the salts of monobasic, bibasic, and tribasic acids.

23. Give an outline of the manufacture of sulphuric acid.

24. Give an outline of the manufacture of carbonate of soda.

ORGANIC CHEMISTRY.

GENERAL INSTRUCTIONS.

Three hours are allowed for this paper.

You are only permitted to attempt eight questions.

You may select your questions from Series I., or from Series II., or from both.

The value attached to the correct answer of each question in Series I. is 8; and those in Series II. is 12·5.

N.B.—A full and exact answer will in all cases gain more marks than an inexact or incomplete answer; though in the former case the question may be the more easy of the two and have less marks attached to it.

SERIES I.

1. What is the composition of cyanogen, and how is it prepared?
2. Describe the method of making hydrocyanic acid.
3. Give an outline of determining the carbon and hydrogen in organic compounds.
4. State the composition of marsh gas and the mode of preparing it.
5. State the composition of olefiant gas and the mode of preparing it.
6. What is meant by the term *substitution?* Give some examples of substitution.
7. What is meant by the term *organic radical?* Give some examples of organic radicals.
8. Give the names and formulæ of some of the simplest organic acids.

9. Give the names and formulæ of some of the simplest organic bases.
10. What becomes of oxalate of lime when it is heated?
11. What is soap?
12. State the nature of the green arsenical colour which is used in the manufacture of paper-hangings.

SERIES II.

13. Describe the process of alcoholic fermentation.
14. Describe the process of acetification.
15. Describe the process of etherification.
16. State the preparation, the composition, the properties, and the principal decompositions of oxalic acid.
17. Describe the method of preparing aniline.
18. Give some examples of alcohols homologous with ethylic alcohol, and also of the acids, hydrated and anhydrous, and of the ethers, simple, mixed, and compound, belonging to them.
19. Give some examples of organic bases, and explain the relation in which these substances stand to ammonia.
20. A substance contains in 100 parts,—

Carbon . . .	38·7
Hydrogen . .	16·2
Nitrogen . .	45·1
	100·0

What is the formula of the compound?

Combining number of carbon = 6
Combining number of hydrogen = 1
Combining number of nitrogen = 14

Give details of calculation.

Solution.

Divide the quantity of each element by its eq.

$$\frac{38\cdot7}{6} = 6\cdot45 \text{ eqs. C.}$$

$$\frac{16\cdot2}{1} = 16\cdot2 \text{ eqs. H.}$$

$$\frac{45\cdot1}{14} = 3\cdot22 \text{ eqs. N.}$$

These numbers show the *proportion* of the equivalents of the elements in the compound. They must now be reduced to whole numbers. To do this the easiest plan is to divide each number by the lowest number of equivalents expressed above, viz. 3·22 eqs. N.

Dividing, then, by 3·22 we get—

$$\frac{6\cdot45}{3\cdot22} = 2 \text{ eqs. C}; \quad \frac{16\cdot2}{3\cdot22} = 5 \text{ eqs. H};$$

$$\frac{3\cdot22}{3\cdot22} = 1 \text{ eq. N.}$$

The composition of the body is C_2H_5N, and it may be represented as $\left.\begin{array}{c} C_2H_3 \\ H \\ H \end{array}\right\} N$, or methylamine.

21. Describe the preparation of fulminate of mercury; and give the composition of this compound.
22. What is meant by the term *isomerism?* Give some example of isomerism.
23. Describe the manufacture of prussiate of potash.
24. Give an outline of the process of saponification.

QUESTIONS *given at the Examination for Certificates in Chemistry, held at the South Kensington Museum, in November* 1861.

[Three hours allowed for this Paper.]

INORGANIC CHEMISTRY.

SECTION I.

1. Describe the usual process for the preparation of chlorine.
2. How is nickel separated from cobalt?
3. The formula usually adopted by chemists for felspar is $KO,SiO_3, + Al_2O_3, 3SiO_3$. In this formula $K = 39$, $Al = 13{\cdot}7$, $Si = 21$, $O = 8$. What is the formula of felspar if the following equivalents be taken: $K = 39$, $Al = 13{\cdot}7$, $Si = 28$, $O = 16$?

Solution.

Empirical formula of felspar $= KAl_2\ Si_4\ O_{16}$, which becomes (using the higher equivalent numbers for Si and O,) $KAl_2\ Si_3\ O_8$. Now, as potash, KO, and alumina, $Al_2\ O_3$, become respectively $KO\frac{1}{2}$ and $Al_2\ O\frac{3}{2}$, it is very evident that the whole formula must be doubled, to do away with the half equivalents. It then stands thus, $K_2\ Al_4\ Si_6\ O_{16}$. Again, the formula for silicic acid, SiO_3, becomes $Si\frac{3}{4}\ O\frac{3}{2}$, and on removing the fractions, it will be seen that the Si and O are united (using the higher equivalents) in the proportion of 1 eq. of Si and 2 eqs. of O, or SiO_2.

The formula $K_2\ Al_4\ Si_6\ O_{16}$ may therefore be written

$K_2O,\ Al_4O_3,\ 6SiO_2$ or $K_2O,3SiO_2 + (Al_2)_2O_3,3SiO_2$,

the compound being regarded as a tersilicate of potash combined with the silicate of alumina.

SECTION II.

1. The specific gravity of nitrogen, referred to that of hydrogen as unity, is 14. What is the specific gravity of ammonia referred to that of hydrogen as unity? Give the method of calculation.

Solution.

Hydrogen and nitrogen have the same combining volume, therefore the weights of their volumes are in the ratio of their equivalents.

Again, 3 vols. of H and 1 vol. of N combine, and condense into 2 vols. of ammonia, NH_3.

Specific gravity of H	1
" " of N	14

$$\therefore \text{Specific gravity of } NH_3 = \frac{1 \times 3 + 14}{2} = \frac{17}{2} = 8 \cdot 5.$$

2. Describe the method of estimating nitrogen by volume.

3. A salt has been found to contain in 100 parts:—

Phosphorus	12·65
Magnesium	9·79
Nitrogen	5·71
Hydrogen	6·54
Oxygen	65·31
	100·00

What is the formula of the salt? The following equivalents to be used:—

Hydrogen	1
Oxygen	8
Nitrogen	14
Phosphorus	31
Magnesium	12

Give details of calculation.

Solution.

Divide the quantity of each element by its equivalent:

$$\frac{12{\cdot}65}{31} = {\cdot}408 \text{ eqs. of P}; \quad \frac{9{\cdot}79}{12} = {\cdot}816 \text{ eqs. of Mg}.$$

$$\frac{5{\cdot}71}{14} = {\cdot}408 \text{ eqs. of N}; \quad \frac{6{\cdot}54}{1} = 6{\cdot}54 \text{ eqs. of H};$$

$$\frac{65{\cdot}31}{8} = 8{\cdot}163 \text{ eqs. of O}.$$

Dividing these quotients by ·408,

$$\text{Eq. of P} = \frac{{\cdot}408}{{\cdot}408} = 1; \quad \text{eq. of Mg} = \frac{{\cdot}816}{{\cdot}408} = 2.$$

$$\text{Eq. of N} = \frac{{\cdot}408}{{\cdot}408} = 1; \quad \text{eq. of H} = \frac{6{\cdot}54}{{\cdot}408} = 16;$$

$$\text{Eqs. of O} = \frac{8{\cdot}163}{{\cdot}408} = 20.$$

The salt is therefore $Mg_2\,NPH_{16}\,O_{20}$.

The P is doubtless in the state of PO_5, the N in the form of NH_4O, and the Mg as MgO.

The formula may therefore be represented as

$$2MgO,NH_4O,PO_5 + 12HO; \text{ or, } Mg_2NH_4,PO_8 + 12HO.$$

SECTION III.

1. Describe a process by which pure oxygen gas may be prepared from atmospheric air.
2. Give an outline of the manufacture of nitric acid, and state the quantities of sulphuric acid (HSO_4), nitrate of potassium (KNO_6), or nitrate of sodium ($NaNO_6$), which are required to produce 1 lb. of nitric acid (HNO_6).

Solution.

$H = 1$; $O = 8$; $N = 14$; $S = 14$; $K = 39$; $Na = 23$.

Formula showing the production of nitric acid :—

$$(1.)\ \left\{\begin{matrix}Na\\ \text{or}\\ K\end{matrix}\right\} NO_5 + HSO_4 = HNO_6 + \left\{\begin{matrix}Na\\ \text{or}\\ K\end{matrix}\right\} SO_4.$$

$$(2.)\ \left\{\begin{matrix}Na\\ \text{or}\\ K\end{matrix}\right\} NO_5 + 2HSO_4 = HNO_6 + \left\{\begin{matrix}Na\\ \text{or}\\ K\end{matrix}\right\} SO_4 + HSO_4.$$

In the first, sulphate of potash or soda is left in the retort; in the second, bisulphate of soda or potash is left.

According to the first formula 1 eq. of HSO_4 is required for the decomposition of 1 eq. of $NaNO_6$ or KNO_6.

Eq. of HSO_4 = 49
,, $NaNO_6$ = 77
,, KNO_6 = 93
,, HNO_6 = 63

So that 49 parts by weight of HSO_4 will produce from 77 parts of $NaNO_6$, or 93 parts of KNO_6, 63 parts of HNO_6. The quantities of each of these required for the production of 1 lb. of HNO_6 may be easily determined by proportion :—

HNO_6.		HNO_6.		HSO_4.		HSO_4.
63	;	1	::	49	:	12·4 ozs. (avoirdupois.)
				$NaNO_6$.		$NaNO_6$.
			::	77	:	1 lb. 3·5 ozs. ,,
				KNO_6.		KNO_6.
			::	93	:	1 lb. 7·619 ozs.

According to the second formula, these quantities remain the same with the exception of HSO_4, of which double the above quantity is employed, viz., 12·4 ozs. × 2 = 1 lb. 8·8 ozs.

3. How many volumes of oxygen gas are required to burn one volume of marsh gas, and one volume of olefiant gas: and how many volumes of carbonic acid gas are produced by their combustion?

Solution.

By volume $C_2 H_4$ is composed of 4 vols. H and 2 vols. C vapour (hypothetical) condensed into 2 vols.

So that 1 vol. of $C_2 H_4$ contains 1 vol. C vapour, + 2 vols. H.

Now CO_2 contains equal volumes of C vapour (hypothetical) and of O, Therefore 1 vol. C vapour requires 1 vol. O to unite with it so as to form 1 vol of CO_4.

Again, HO contains H and O in proportion of 2 vols. H to 1 vol. O.

Therefore 2 vols. H require for complete combustion 1 vol. O.

So that 1 vol. $C_2 H_4$, containing 1 vol. C and 2 vols. H, requires altogether 2 vols. O for complete combustion, from which 1 vol. of CO_2 is formed.

Again, $C_4 H_4$ is composed, by volume, of 4 vols. C vapour and 4 vols. H condensed into 2 vols.

So that 1 vol. $C_4 H_4$ contains 2 vols. C vapour + 2 vols. H.

While the H then remains the same as in $C_2 H_4$, the volume of the C is doubled, requiring twice as much O, or 2 vols.

Quantity of O required by the 2 vols. H = 1 vol.

So that the total quantity of O required is 3 vols., of which 2 vols. combine with the C, forming 2 vols. CO_2.

SECTION IV.

1. Describe the chemical reactions involved in the etching of glass.
2. State the succession of processes by which aluminium may be extracted from alum shale.

3. What is meant by the term disintegration? Give an outline of the several phases of the disintegration of granite.

PRACTICAL EXAMINATION IN ANALYSIS.

1. Qualitative analysis of a mixture consisting of two acids and two bases soluble in water or acids.
2. Qualitative analysis of a mixture consisting of two acids and two bases partly or entirely insoluble in water and acids.
3. Qualitative analysis of more complex mixtures.

State the method of examination in detail, and tabulate the result of the analysis, pointing out particularly the substances of which traces only have been found.

ORGANIC CHEMISTRY.

[Three hours allowed for this Paper.]

SECTION I.

1. What is the action of heat upon formiate of ammonium?
2. Describe the processes for transforming benzol into aniline.
3. An organic acid has been found to contain in 100 parts,—

Carbon . . .	60·87
Hydrogen . . .	4·35
Oxygen . . .	34·78
	100·00

The analysis of the silver salt of this acid has given the following result: 1·052 parts of the silver salt left upon ignition 0·4637 parts of metallic silver. What is the formula of the acid? Give details of calculation.

Solution.

Dividing by each equivalent we obtain the following quotients, expressing the proportionate number of equivalents in the compound:

$$\text{Eqs. of C} = \frac{60\cdot 87}{6} = 10\cdot 145.$$

$$\text{,,} \quad \text{H} = \frac{4\cdot 35}{1} = 4\cdot 35.$$

$$\text{,,} \quad \text{O} = \frac{34\cdot 78}{8} = 4\cdot 35.$$

Dividing these quotients by 4·35, we have

Eq. of O = 1
,, H = 1

Eqs. of C combined with 1 eq. of O or H = 2·33

Multiplying these numbers by 3, to clear of fractional parts,

Eqs. of C = 7
,, H = 3
,, O = 3

The empirical formula is therefore $C_7 H_3 O_3$.

Again, 1·052 parts of the silver salt yield ·4637 parts of silver; the quantity of the salt necessary to yield 108 parts or 1 eq. of Ag may be easily found by the following proportion:—

Ag.		Ag.		Silver Salt.		Silver Salt.
·4637	:	108	::	1·052	:	245

This 245 is then the equivalent of the silver salt, and deducting the equivalent of AgO, 108 + 8 = 116, we obtain 245 − 116 = 129 for the equivalent of the *dry acid* or *anhydride*.

The *real* acid will contain in addition HO, and its equivalent will therefore be 129 + 9 = 138.

Now the equivalent of the empirical formula, $C_7 H_3 O_3$, given above, is 42 + 3 + 24 = 69, which is just half the *real* equivalent 138. So that it is evident this formula $C_7 H_3 O_3$ must be doubled so as to express one equivalent of the acid. It then becomes $C_{14} H_6 O_6$ or $C_{14} H_5 O_5$, HO.

SECTION II.

1. Describe the processes of transforming ferrocyanide of potassium respectively into urea and cyanuric acid.
2. State the preparation, the composition, the properties, and the principal transformations of oxalic acid.

3. A certain variety of malt has been found to contain 10 per cent. of starch ($C_{12} H_{10} O_{10}$) and 20 per cent. of sugar ($C_{12} H_{12} O_{12}$). Describe the changes which starch and sugar undergo during their transformation into acetic acid ($C_4 H_4 O_4$), and state the quantity of malt required for the production of 100 pounds of acetic acid, supposing the whole amount of starch and sugar to be converted.

Equation (1) expressing the change of sugar into alcohol;

$$(1.)\ C_{12}H_{12}O_{12} = 2(C_4H_6O_2) + 4CO_2.$$

Equations (2, 3) expressing the change of starch into sugar and alcohol;

$$(2.)\ C_{12}H_{10}O_{10} + 2HO = C_{12}H_{12}O_{12}.$$
$$(3.)\ C_{12}H_{12}O_{12} = 2(C_4H_6O_2) + 4CO_2.$$

Equation (4) showing the conversion of alcohol into acetic acid;

$$(4.)\ C_4H_6O_2 + O_4 = C_4H_4O_4 + 2HO.$$

It will be seen from these equations, that 1 eq. sugar produces 2 eqs. alcohol; and as 1 eq. alcohol yields 1 eq. acetic acid, therefore from 1 eq. sugar 2 eqs. of acetic acid are formed. Similarly, 1 eq. starch produces 2 eqs. acetic acid.

Now—

Eq. sugar $C_{12}H_{12}O_{12} = 72 + 12 + 96 = 180$

Eq. starch $C_{12}H_{10}O_{10} = 72 + 10 + 80 = 162$

Eq. acetic acid $C_4 H_4O_4 = 24 + 4 + 32 = 60$

2 eqs. „ „ $= 60 + 2 = 120$

Therefore 180 lbs. sugar yield 120 lbs. acetic acid; and 162 lbs. starch yield the same quantity.

The quantity of starch in the malt is 10 per cent., and of sugar 20 per cent.

Then by proportion,

Starch.		Starch.		$\bar{A}$		$\bar{A}$
162	:	10	::	120	:	$7\frac{11}{27}$
Sugar.		Sugar.				
180	:	20	::	120	:	$13\frac{1}{3}$

Total quantity of acetic acid from 100 lbs. malt $= 7\frac{11}{27} + 13\frac{1}{3} = 20\frac{20}{27}$ lbs.

The quantity of malt necessary for the production of 100 lbs. acetic acid can then easily be found by proportion;

$\bar{A}$.		$\bar{A}$.		Malt.		
$20\frac{20}{27}$	:	100	::	100	:	$462\frac{1}{7}$ lbs. of malt required.

SECTION III.

1. Express in formulæ the transformation of oxalic ether into oxamat of ethyl and oxamide.
2. Describe the method of preparing the chlorides of the acid-radicals, and their deportment under the influence of water, alcohol, and ammonia.
3. In determining the vapour-density of an organic compound, according to Dumas's process, the following observations were recorded :—

Temperature of air . .	12° C.
Temperature of vapour . .	152° C.
Weight of glass globe filled with the vapour *minus* weight of globe filled with air . .	0·386 grammes.
Capacity of globe . .	292 cubic centimetres.
Barometer . .	0·762 metres.

100 cubic centimetres of air at 0° C. and bar. 0·760 metres weigh 0·12932. What is the vapour-density of the compound referred to hydrogen as unity? Give details of calculation.

Solution.

1st. To find the weight of the air contained in the globes, viz., 292 c. c. at 12° C. and ·760 m.

Co-efficient of expansion of air for 1° = ·00367.

1. c. c. of air at 0° becomes $1 + (\cdot 00367 \times 12) = 1\cdot 044$ c. c. at 12°.

$\therefore$ 1·044 c. c. at 12°, reduced to 0° become 1 c. c. and 292 c. c. at 12°, reduced to 0° become

$$\frac{292}{1\cdot 044} = 279\cdot 694 \text{ c. c.}$$

Correcting for barometric pressure; as the volumes of gases are inversely as the pressure, 279·694 c. c. at a pressure of ·762 m. become at ·760 m.—

$$\frac{279\cdot 694 \times 762}{760} = 280\cdot 43 \text{ c. c.}$$

Now 100. c. c. air at 0° and ·760 m. weigh ·12932 grm.

$\therefore$ 280·43 c. c. air at 0° and ·760 m. weigh

$$\frac{\cdot 12932 \times 280\cdot 43}{100} = \cdot 3626 \text{ grm.}$$

i. e., the weight of air in globe = ·3626 grm.

2nd. To find the weight of substance in the state of vapour.

As the globe and vapour − globe and air = ·386 grm.

$\therefore$ Weight of vapour = ·386 + ·3626 = ·7486 grm.

Hence 292 c. c. of vapour at 152° C. and ·762 m. weigh ·7486 grm.

3rd. To reduce this volume to 0° and ·760 m.

1. c. c. at 0° becomes $1 + (\cdot 00367 \times 152) = 1\cdot 55784$ at 152°, *i. e.*, 1·55784 at 152°, when cooled to 0°, become 1 c. c.

∴ 292 c. c. vapour at 152°, cooled to 0° become—

$$\frac{292}{1 \cdot 55784} = 187 \cdot 44.$$

Correcting this for pressure, 187·44 at a pressure of ·760 m. become $\frac{187 \cdot 44 \times 762}{760} = 187 \cdot 9$ c. c. which weigh ·7486 grm.

Weight of 100. c. c. vapour at 0° and ·760 m = $\frac{\cdot 7486 \times 100}{187 \cdot 9} = \cdot 3984.$

Weight of 100. c. c. air at 0° and ·760 m. = ·12932.

∴ Density of vapour = $\frac{\cdot 3984}{\cdot 12932} = 3 \cdot 0807.$

Now the density of hydrogen compared with air is as ·0693 to 1.

Therefore, density of vapour referred to hydrogen as unity = $\frac{3 \cdot 0807}{\cdot 0693} = 44 \cdot 4.$

SECTION IV.

1. Give some examples of monoatomic, diatomic, and triatomic alcohols. State their formulæ.
2. Describe a process by which alcohol may be converted into propionic acid, and acetic acid into glycolic acid.
3. Calculate the quantities of binoxide of manganese and hydrochloric acid which are necessary to convert, respectively, into chloroform ($C_2 H Cl_3$) and chloride of carbon ($C_2 Cl_4$) the quantity of marsh gas evolved by the distillation of 1 lb. of acetic acid ($C_4 H_4 O_4$) with an excess of lime. Give details of calculation.

Solution.

Equation (1) representing the resolution of acetic acid into marsh gas and carbonic acid:—

$$(1.)\ C_4H_4O_4 = C_2H_4 + 2CO_2.$$

Equation (2) showing the ultimate production of chloroform from marsh gas:—

$$(2.)\ C_2H_4 + 6Cl = \underbrace{C_2HCl_3}_{\text{Chloroform.}} + 3HCl.$$

Equation (3) representing the ultimate conversion of marsh gas into perchloride of carbon:—

$$(3.)\ C_2H_4 + 8Cl = C_2Cl_4 + 4HCl.$$

Equation (4) showing the liberation of Cl from the action of HCl on MnO_2:—

$$(4.)\ MnO_2 + 2HCl = MnCl + 2HO + Cl.$$

It will be seen from these equations that 1 eq. acetic acid liberates 1 eq. marsh gas, which requires 6 eqs. chlorine for its conversion into chloroform; 6 eqs. chlorine are formed from 6 eqs. MnO_2 and 12 eqs. HCl.

Eq. acetic acid = 60; 6 eqs. MnO_2 = $43{\cdot}6 \times 6 = 261{\cdot}6$;

$$12 \text{ eqs. HCl} = 36{\cdot}5 \times 12 = 438.$$

Therefore the quantities of MnO_2 and of HCl required for the conversion of the marsh gas evolved from 60 lbs. acetic acid are 261·6 lbs. MnO_2 and 438 lbs. HCl.

Dividing each of these by 60, we obtain,

Quantity of MnO_2 required to convert into C_2HCl_3 the marsh gas from 1 lb. acetic $= \frac{261{\cdot}6}{60} = 4{\cdot}36$ lbs. MnO_2.

Quantity of HCl required to convert into C_2HCl_3 the marsh gas from 1 lb. acetic $= \frac{438}{60} = 7{\cdot}3$ lbs. HCl.

Again, as $C_2 H_4$ requires for its conversion into perchloride of carbon 8 eqs. Cl, (see equation 2), we obtain the quantities of MnO_2 and HCl necessary for the production of $C_2 Cl_4$ from the $C_2 H_4$ evolved from 1lb of $C_4 H_4 O_4$, by a very simple proportion:—

Eqs. Cl. Eqs. Cl. MnO_2. MnO_2.
6 : 8 :: 4·36 : 5·813 lbs. MnO_2 required.
HCl.
6 : 8 :: 7·3 : 9·73 lbs. HCl. required.

QUESTIONS *given at the Examination of Pupils in Science Classes and Schools, in May* 1862.

INORGANIC CHEMISTRY.

GENERAL INSTRUCTIONS.

Three hours are allowed for this paper.

You are only permitted to attempt eight questions.

You may select your questions from Series I., or from Series II., or from both.

The value attached to the correct answer of each question in Series I. is 8; and of each question in Series II. is 12·5.

N.B.—A full and exact answer will in all cases gain more marks than an inexact or incomplete answer; though in the former case the question may be the more easy of the two and have less marks attached to it.

SERIES I.

1. What methods are generally employed for the preparation of hydrogen? Represent the reactions by diagrams or formulæ.
2. How is chlorine prepared?
3. What is the action of nitric acid upon copper?
4. What is the action of chorine upon a solution of ammonia?
5. What is the composition of water by weight and by volume?
6. How is bleaching powder prepared?
7. Express the preparation of nitric acid in formulæ.
8. What is the composition of the principal mineral from which lead is obtained?

9. The formula of oxide of silver is AgO, the equivalent of silver is 108, the equivalent of oxygen is 8. How many pounds of oxide of silver are necessary to make 100 lbs. of oxygen?

Solution.

Eq. oxide of silver $\mathrm{AgO} = 108 + 8 = 116$.
From 116 lbs. AgO, 8 lbs. of O will be obtained;

$$\begin{array}{lcccccccc} & \mathrm{O.} & & \mathrm{O.} & & \mathrm{AgO.} & & \\ \text{therefore by proportion} & 8 & : & 100 & :: & 116 & : & 1450\,\text{lbs.} \end{array}$$

AgO required to make 100 lbs. O.

10. What is the source of hydrochloric acid, and how is it prepared?
11. On what principle does the hardening of mortars depend?
12. How is caustic potash prepared?

SERIES II.

13. How are barium, strontium, and calcium detected when mixed?
14. State the composition of the principal iron ores, and describe the process of extracting the iron from them.
15. How are the manganates and permanganates prepared?
16. Describe the manufacture of sesquioxide of chromium from chrome iron ore.
17. How is chloric acid prepared?
18. What is the composition of gunpowder, and what reaction takes place during its explosion?
19. How many compounds of phosphorus and hydrogen are there, and how are they obtained?
20. How are the oxides of nitrogen prepared?
21. How is hydrate of baryta procured from native carbonate and sulphate of baryta?

22. How are the chlorides of mercury obtained?

23. A substance has been found to contain—

Arsenic . . .	65·22
Oxygen . . .	34·78
	100·00

The combining number of arsenic is 75
The combining number of oxygen is 8
What is its formula?
Give details of calculation.

Solution.

Divide the quantity of each element present by its equivalent:—

$$\frac{65\cdot22}{75} = \cdot87 \text{ eqs. As}; \quad \frac{34\cdot78}{8} = 4\cdot347 \text{ eqs. O.}$$

These numbers represent the proportion in equivalents in which the As and O are combined. It will be at once seen that the number of equivalents of O is 5 times the number of eqs. of As; therefore the formula of the compound is AsO_5.

24. The formula of sulphate of copper is $CuSO_4+5HO$.
What is its percentage composition?

Combining number of copper = 31·5
Combining number of sulphur = 16
Combining number of oxygen = 8
Combining number of hydrogen = 1
Give details of calculation.

Solution.

Collecting the elements together, 1 eq. sulphate of copper contains—

1 eq. Cu = 31·5
1 eq. S = 16
9 eqs. O = 72
5 eqs. H = 5

∴ eq. of sulphate of copper = 124·5.

In 124·5 parts, then, of sulphate of copper there are 31·5 of Cu, 16 of S, 72 of O, and 5 of H. The quantity of each element in 100 parts may now be easily found by the following proportion :—

124·5	:	100	::	Cu, 31·5	:	25·3	parts of Cu.
				S, 16	:	12·85	,, S.
				O, 72	:	57·83	,, O.
				H. 5	:	4·02	,, H.
						100·00	

ORGANIC CHEMISTRY.

General Instructions.

Three hours are allowed for this paper.

You are only permitted to attempt eight questions.

You may select your questions from Series I., or from Series II., or from both.

The value attached to the correct answer of each question in Series I. is 8; and of each question in Series II. is 12·5.

N.B.—A full and exact answer will in all cases gain more marks than an inexact or incomplete answer; though in the former case the question may be the more easy of the two and have less marks attached to it.

Series I.

1. How is sulphocyanide of potassium prepared?
2. What is the action of chlorine upon olefiant gas?
3. In what relation does aldehyde stand to alcohol and acetic acid?
4. Give examples of monobasic, bibasic, and tribasic organic acids.
5. What is the action of nitric acid upon benzol?
6. What is meant by the term, series of fatty acids?
7. What is the action of excess of baryta on benzoic acid at a high temperature?
8. Describe the process for determining nitrogen by its conversion into ammonia.
9. Describe the mode of manufacturing oxalic acid.
10. Explain the manufacture of Prussian blue from yellow prussiate of potash.

11. What are homologous bodies? Give some examples.
12. State the nature of the ingredients of methylated spirit.

SERIES II.

13. What is the action of heat and dehydrating agents on formiate of ammonium and acetate of ammonium?
14. What is the action of solution of ammonia upon oxalic ether? What is the constitution of the principal product?
15. How is chloride of acetyl prepared?
16. By what succession of processes may wood be converted into spirit of wine?
17. Describe the chemical processes involved in the brewing of beer..
18. How may we pass from the ethyl series to a member of the next higher series?
19. Give an example of a diatomic alcohol, and what are the acids that may be derived from it?
20. Give an outline of the manufacture of white lead.
21. How is benzoic alcohol prepared?
22. What is the action of chloride of benzoyl on dry carbonate of ammonium, and what is the constitution of the product?
23. A compound has been found to contain—

Carbon . .	40·00
Hydrogen . .	6·67
Oxygen . .	53·33
	100·00

What is its formula?

Combining number of carbon . .	1
Combining number of hydrogen .	1
Combining number of oxygen . .	8

Give details of calculation.

Solution.

Divide the quantity of each element present by its equivalent:—

$$\frac{40\cdot00}{6} = 6\cdot66, \text{ or } 6\tfrac{2}{3} \text{ eqs. of C,}$$

$$\frac{6\cdot67}{1} = 6\cdot67, \text{ or } 6\tfrac{2}{3} \text{ eqs. of H,}$$

$$\frac{53\cdot33}{8} = 6\cdot66, \text{ or } 6\tfrac{2}{3} \text{ eqs. of O,}$$

It will be seen that the compound contains an equal number of equivalents of C, H, and O: its empirical formula then is CHO, and therefore its real formula cannot with certainty be pointed out; it may be $C_2H_2O_2$, or $C_3H_3O_3$, or $C_4H_4O_4$ (acetic acid).

24. **What are simple, mixed, and compound ethers, and in what relation do they stand to water?**

QUESTIONS *given at the Examination for Certificates in Chemistry, held at the South Kensington Museum, in November* 1862.

[Three hours allowed for this Paper.]

INORGANIC CHEMISTRY.

SECTION I.

1. Describe a *chemical* process by which hydrogen, and one by which oxygen may be prepared from water.
2. What is the action of chlorine upon hydrate of lime at the common temperature and at the temperature of boiling water?
3. V volumes of sulphuretted hydrogen are decomposed by an excess of bromine: how many volumes of hydrobromic acid are generated by this reaction? State the mode of reasoning whereby the result has been arrived at.

Solution.

1 vol. HS contains 1 vol. H and $\frac{1}{6}$ vol. S vapour, as shown by the following calculation:—

Sp. gr. (or weight of 1 vol.) HS = 1·177.

Composition by weight = 16 parts S + 1 part H.

1 vol. ∴ of HS contains $\frac{16}{17}$ of its weight of S = $\frac{16}{17}$ of 1·177 = 1·108,

and $\frac{1}{17}$ of its weight of H = $\frac{1}{17}$ of 1·177 = ·069.

Now ·069 is the sp. gr. of H, or weight of 1 vol.

and 1·108 is $\frac{1}{6}$ of the sp. gr. of S, or $\frac{1}{6}$ the weight of 1 vol. S.

∴ 1 vol. HS contains 1 vol. H and $\frac{1}{6}$ vol. S vapour; and when V vols. HS are decomposed, V vols. H are liberated.

Again, H and Br unite to HBr, in equal volumes, without condensation, which may be shown in the same manner as above, or as follows:—

Sp. gr. Br vap. = 5·393
„ H = ·0695

Weight of 1 vol. H + 1 vol. Br vap. 5·4625

The sp. gr. HBr vap. being 2·73, it follows that the 1 vol. H and 1 vol. Br. have formed 2 vols. HBr; and that for every vol. of H there are 2 vols. of HBr. Consequently V vols. H when liberated from HS by action of Br, form 2 V vols. HBr.

SECTION II.

1. The formula of alum is $KO,SO_3 + Al_2O_3, 3SO_3 + 24\ HO$: what is the composition of the salt in 100 parts? The follow equivalents to be used:—

H = 1
O = 8
S = 16
K = 39
Al = 13·7

Give such numerical details as will show the method of calculation.

Solution.

Collecting the elements eq. alum contains—

1 eq. K = 39
2 eqs. Al = 27·4
4 eqs. S = 64
40 eqs. O = 320
24 eqs. H = 24

Eq. alum = 474·4

The percentage composition may be readily determined from these figures by simple proportion:—

Alum.		Alum.		K.		
474·4	:	100	::	39	:	8·22 K per cent.
				Al.		
			::	27·4	:	5·77 Al „
				S.		
			::	64	:	13·5 S „
				O.		
			::	320	:	67·45 O „
				H.		
			::	24	:	5·06 H „

2. Describe the method of estimating chlorine, bromine, and iodine, when the three elements occur together.

3. V volumes of a hydrocarbon Cn Hn is submitted to combustion. How many volumes of oxygen are required for its complete combustion and how many volumes of carbonic acid gas are generated?

Solution.

Gaseous hydrocarbons of the Cn Hn series contain 2n volumes of C, and 2n volumes of H, condensed into 4 volumes.

Therefore, 4 vols. Cn Hn are resolved into 2n vols. C (vap.), and 2n vols. H.

By proportion V vols. Cn Hn contain $\frac{2Vn}{4}$ vols. of C, and the same number of vols. of H.

Now CO_2 contains equal vols. of C vap. and O condensed one-half. Therefore to unite with $\frac{2Vn}{4} = \frac{Vn}{2}$ vols. C vap. require $\frac{Vn}{2}$ vols, O; and these condensing to one-half there will be $\frac{Vn}{2}$ vols. of CO_2 formed.

Again, as water (vapour) is composed of 2 vols. H to 1 vol. O, to unite with $\frac{Vn}{2}$ vols. of H will require

$$\frac{Vn}{2 \times 2} = \frac{Vn}{4} \text{ vols. O};$$

and therefore total No. vols. O required for complete combustion of V. vols. of $Cn\,Hn = \frac{Vn}{2} + \frac{Vn}{4} = \frac{3\,Vn}{4}$

SECTION III.

1. Give a definition of the terms acid, base, and salt.
2. Give an outline of the processes involved in the manufacture of sulphuric acid.
3. The formula of the mineral stilbite usually adopted by chemists is $CaO, SiO_3 + Al_2O_3, 3\,SiO_3 + 6\,HO$. In this formula the symbols have the following values:—

H = 1
O = 8
Al = 13·7
Si = 21
Ca = 20

What is the formula of stilbite, if the following equivalents be used:—

H = 1
O = 16
Al = 27·4
Si = 28
Ca = 40

Give such numerical details as will show the mode of calculation.

Solution.

Collecting the elements:—

1 eq. of the compound contains 1 eq. Ca
4 eq. Si
2 eq. Al
22 eq. O
6 eq. H

With the higher equivalents these numbers become:—

½ eq. Ca	and doubling these to get rid of fractions they are	1 eq. Ca
3 „ Si		6 „ Si
1 „ Al.		2 „ Al
11 „ O		22 „ O
6 „ H.		12 „ H.

Now in lime the calcium and oxygen are to each other as 20 : 8 = 40 : 16. Therefore lime is CaO as before.

In alumina the Al and O are to each other as 13·7 × 2 : 24 = 27·4 × 2 = 48. Therefore alumina is Al_2O_3 as before.

In silica the Si and O are as 21 : 24 = 28 : 32. Therefore silica is now SiO_2.

Lastly, in water the H and O are as 1 : 8 = 2 : 16 and the formula of water is therefore H_2O.

The formula then of stilbite becomes—

$$CaO, 3SiO_2 + Al_2O_3, 3SiO_2 + 6H_2O,$$

being regarded as a tersilicate of lime combined with the neutral silicate of alumina.

SECTION IV.

1. Describe one of the methods of estimating the amount of oxygen in atmospheric air.
2. The specific gravity of chlorine, referred to that of hydrogen as unity, is 35·5. What is the specific gravity of hydrochloric acid referred to that of hydrogen as unity?

Solution.

Weight of 1 vol. H = 1
" " Cl = 35·5

H and Cl unite in equal vols., without condensation, to form HCl,

∴ Weight of 2 vols. HCl = 36·5
and " 1 " " = 18·75

which is ∴ the sp. gr. of HCl.

3. Describe the action of sulphur on the alkaline carbonates at a high temperature, and the deportment of the resulting product under the influence of atmospheric air and acids.

ORGANIC CHEMISTRY.

[Three hours allowed for this Paper.]

SECTION I.

1. What is soap?
2. How is acetic acid converted into acetic anhydride?
3. The analysis of a basic compound consisting of carbon, hydrogen, and nitrogen, furnished the following results:—

0·250	grammes of substance gave on combustion	
0·632	,,	carbonic acid (CO_2)
0·337	,,	water.

The base forms, with hydrochloric acid and bichloride of platinum, a crystalline platinum salt:—

0·500	grammes of double salt left on ignition	
0·168	,,	platinum.

What is the percentage composition of the base, and what is its formula? Give such numerical details as will show the method of calculation.

Equivalents to be used—

C = 12
H = 1
N = 14
O = 16
Pt = 98·7
Cl = 35·5

Solution.

(1). To find the C; eq. $CO_2 = 12 + 32 = 44$

$$44 : {\cdot}632 :: 12 : {\cdot}17237$$

(2). To find the H. Water = H_2O.

$$18 : 2 :: {\cdot}337 : {\cdot}03744$$

3. To find the N.

$.250 - (\cdot 17237 + \cdot 03744) = \cdot 250 - \cdot 20981 = \cdot 04019$

Dividing each of these numbers by the eqs. of C, H, and N, respectively—

$$\frac{\cdot 17237}{12} = \cdot 01436, \text{ or } 1436 \text{ eqs. C.}$$

$$\frac{\cdot 03744}{1} = \cdot 03744, \text{ or } 3744 \text{ eqs. H.}$$

$$\frac{\cdot 04019}{14} = \cdot 00287, \text{ or } 287 \text{ eqs. N.}$$

Divide each of these by the number of eqs. of N = 287.

$$\frac{1436}{287} = 5 \text{ eqs. C}; \quad \frac{3744}{287} = 13 \text{ eqs. H}; \quad \frac{287}{287} = 1 \text{ eq. N.}$$

Empirical formula = $C_5H_{13}N$.

Again, in platinum salts 1 eq. Pt. is combined. By proportion it is easy to find the weight of the salt which contains 1 eq. of Pt.

Pt.		Pt.		Pt. salt.		Pt. salt.	
168	:	98·7	::	500	:	293·7	= eq. of Pt. salt.

Now the platinum salts of bases of the ammonia type (of which this is one) possess the general composition X HCl, $PtCl_2$, where X stands for the base. To obtain the equivalent therefore of X, we must deduct from the eq. of the Pt. salt (293·7) 1 eq. Pt., 3 eqs. Cl, and 1 eq. H.

Eq. Pt salt = 293·7

1 eq. Pt = 98·7
3 eq. Cl = 106·5
1 eq. H = 1

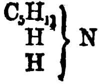
206·2

Eq. of base = 87·5

Or, omitting fractional parts (owing to experimental inaccuracies), eq. of base = 87. This number coincides with the formula given above, $C_5H_{13}N_1$, for—

$$\begin{array}{lcrcr} C_5 & = & 5 \times 12 & = & 60 \\ H_{13} & = & 13 \times 1 & = & 13 \\ N & = & 14 & = & \underline{14} \\ & & & & 87 \end{array}$$

Therefore $C_5H_{13}N$ is the true formula of the base, which may be represented, $\left.\begin{matrix} C_5H_{11} \\ H \\ H \end{matrix}\right\} N$ = Amylamine.



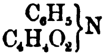
$$\left.\begin{matrix} C_5H_{11} \\ H \\ H \end{matrix}\right\} N = \text{Amylamine}$$

SECTION II.

1. How is benzoic ether prepared?
2. Describe the method of determining carbon and hydrogen in an organic compound.
3. Describe the action of metallic zinc upon iodide of ethyl, and the products which are formed in the reaction.

SECTION III.

1. Describe the composition and the preparation of acetone.
2. State the preparation, the composition, and the three principal transformations of oil of bitter almonds.
3. Given all the inorganic reagents of the laboratory and of organic substances, moreover, benzoic acid $C_7H_6O_2$, and succinic acid $C_4H_6O_4$, how is the substance



$$\left.\begin{matrix} C_6H_5 \\ C_4H_4O_2 \end{matrix}\right\} N$$

prepared?

Express the several reactions in equations.

SECTION IV.

1. What does the term substitution mean in chemistry? Give some examples of substitutions.
2. How is alcohol converted into glycocol?
3. Describe the succession of processes by which olefiant gas is transformed into lactic acid.

PRACTICAL EXAMINATION IN ANALYSIS.

[Seven hours allowed.]

1. Qualitative analysis of a mixture consisting of two acids and two bases soluble in water or acids.
2. Qualitative analysis of a mixture consisting of two acids and two bases partly or entirely insoluble in water and acids.
3. Qualitative analysis of more complex mixtures.

State the method of examination in detail, and tabulate the result of the analysis at the bottom of the page.

EXPLANATION OF SOME OF THE TERMS USED IN CHEMISTRY.

Acetates.—Salts formed by the union of acetic acid with any base.

Æriform Fluids.—Bodies which exist in a gaseous form, such as the air.

Affinity.—A term used to express the force by which bodies combine chemically.

Albumen.—A substance containing N, which exists in the serum of the blood, the white of eggs, butchers' meat, and other bodies.

Alcohol.—Highly rectified spirits of wine.

Allotropic.—A body having different physical forms; carbon, for instance, occurs as plumbago, charcoal, and the diamond.

Alembic.—An apparatus used for sublimations and distillations.

Alkalies.—Substances such as soda, lime, potash, and ammonia; they have a strong tendency to combine with acids and form alkaline salts.

Alloy.—A combination of two or more metals, such as copper and zinc, to form brass. Gold used for coinage is alloyed with copper or silver.

Amalgam.—A mixture of mercury with any metal.

Amorphous.—Without definite form; the opposite of the crystallized form of bodies. Charcoal is amorphous; the diamond is crystallized.

Ammoniacal Salts.—Salts formed by the union of any acid with ammonia, or by the union of the radical of a hydrogen acid with ammonium.

Analysis.—The separation of a compound body either into its proximate or elementary constituents.

Anhydrous.—Destitute of water.

Anti-septic.—Possessing the power of preventing putrefaction.

Argillaceous.—A term applied to substances which contain alumina, such as clay.

Arsenites.—Salts formed by the union of arsenious acid with any base.

Arseniates.—Salts formed by the union of arsenic acid with any base.

Atom.—The smallest ultimate particle of matter.

Atmospheres.—This word is used to express the additional pressure given to fluids. If, for instance, a vessel be filled with carbonic acid gas, and a pressure applied of 15 lbs. upon every square inch of the surface, it is said to be a pressure of one atmosphere: if 30 lbs., two atmospheres; if 45 lbs., three atmospheres, and so on.

Azote.—Another name for nitrogen.

Barometer.—An instrument which shows the variation of the pressure of the atmosphere by the rise and fall of a column of mercury contained in a glass tube.

Base.—Any substance which unites with an acid to form a salt.

Bath, Sand.—Vessels partly filled with dry sand, in which retorts are placed. Baths are sometimes of mercury, water, or steam.

Borates.—Salts formed by the union of boracic acid with a base.

Button.—The name given to the piece of metal which is found at the bottom of a crucible after a metallic ore or oxide has been reduced.

Calcareous.—A term usually applied to combinations of lime with carbonic acid.

Caoutchouc.—India-rubber.

Calcination.—The application of a red heat to bodies, so as to deprive them of moisture.

Capillary Attraction.—A term applied to that species of attraction which causes the rise of sap in vegetables, or the rise of fluids in small tubes.

Carbonates.—Salts formed by the union of carbonic acid with any base.

Carbides.—Substances in which carbon forms one of the constituents, as carbide of iron.

Caustic.—Possessing the power of burning.

Caseine.—An organic body of which the curd of milk is composed.

Catalysis.—An action of contact or surface.

Centigrade.—The thermometric scale, in which the freezing point is 0° and the boiling point 100°.

Chromates.—Salts formed by the combination of any base with chromic acid.

Concentration.—The act of increasing the specific gravity of bodies. It is usually applied to fluids by evaporating a portion of the water.

Combustion.—A term generally applied to express the phenomena of the evolution of light and heat when a substance is burned; but the gradual decay of vegetable and animal substances, and the oxidation of the metals, are instances of slow combustion.

Condensation.—The bringing of the parts of a gas or vapour nearer together by pressure or cold.

Crystallization.—An operation by which bodies pass from a fluid to a solid state, and assume definite geometrical forms.

Crystallization, Water of.—That portion of water which combines with salts in crystallizing, and which cannot be removed without breaking up the compound.

Decant.—To pour off a liquid from the residue.

Decomposition.—The breaking up of the constituent parts of a compound body.

Decrepitation.—The sudden decomposition of bodies, attended with a crackling noise, when thrown into a red-hot crucible.

Deflagration.—The vivid combustion that is produced when the nitrates and some other salts are exposed to a red heat.

Deliquescence.—Signifies the absorption of moisture from the atmosphere, by which salts become liquid.

Desiccation.—The operation of thoroughly drying a substance.

Deoxidise.—To deprive of oxygen.

Detonation.—An explosion.

Digestion.—The continual soaking of a solid substance in a liquid, with or without the application of heat.

Dimorphism.—The property which some bodies possess of crystallizing in two distinct forms.

Ductility.—A property by which metals may be drawn out without fracture.

Effervescence.—A motion occasioned by the sudden liberation of gaseous bodies, as when hydrochloric acid is poured on chalk, or a solution of tartaric acid is added to carbonate of soda.

Efflorescence.—A term applied to salts when they lose part of their water of crystallization on exposure to the air, and become covered with powder, or pulverulent. This appearance is sometimes observable on brick walls.

Elasticity.—The force by which bodies endeavour to recover their natural position, from whence they had been displaced by external pressure.

Elastic Fluids.—A name given to vapours and gases.

Element.—A body that cannot be resolved into any simpler form.

Empyreumatic.—A disagreeable odour arising from the burning of animal oils and vegetable substances in close vessels.

Eremacausis.—Slow combustion, as the rotting of wood.

Evaporation.—The conversion of fluids into vapour by heat.

Eudiometer.—An instrument used for investigating the composition of different gases and the purity of atmospheric air.

Fahrenheit.—Thermometric scale. The boiling point is 212°, and the freezing point 32°.

Fibrin.—A white fibrous substance which may be obtained by agitating warm blood with a birch rod, and which chiefly composes muscular tissue.

Filtration.—A process for separating liquids in a state of purity from solid bodies.

Filtrate.—The clear liquid which passes through the filter.

Flux.—A substance which when heated with another forms a glass, and assists it to fuse.

Fusion.—The rendering of a body fluid by heat.

Gaseous.—Having the nature and properties of a gas.

Gasometer.—An apparatus for collecting and preserving gases.

Gelatine.—A chemical term for glue and animal jellies.

Gluten.—A vegetable substance analogous to gelatine.

Haloid.—A term for iodine, chlorine, bromine, and fluorine, which produce salts by their union with the metals, as chloride of sodium.

Hermetically Sealed.—A term applied to the closing of the end of a glass tube, so as to render it air-tight.

Homogeneous.—Alike in nature and properties.

Hydrometer.—An instrument for determining the specific gravity of fluids.

Hygrometer.—An instrument for determining the quantity of moisture in the atmosphere.

Hydrate.—A compound formed by the union of water with a compound.

Hydracids.—Acids formed by the union of hydrogen with some other element, as hydrochloric acid.

Ignition.—The act of exposing to a red heat. Strictly speaking, in ignition the body is not changed, whereas in combustion it is.

Isomorphous.—Bodies having the same crystalline form, and capable of replacing each other; thus anhydrous arsenic acid (AsO_5), and phosphoric acid (PO_5), are isomorphous.

Isomeric.—Bodies having the same formula but of different characters. Sugar and gum are isomeric.

Lixiviation.—The act of dissolving salts.

Litmus Paper.—Paper coloured blue with an infusion of litmus, and used as a test for acids.

Menstruum.—The fluid in which a solid body is dissolved.

Metamorphis.—The transformation of chemical compounds.

Mordants.—Substances which have an affinity for particular colours and fabrics, and employed by dyers to unite the colour with the cloth : alum is of this class.

Nascent.—In the moment of formation. A gas, for instance, is said to be nascent at the moment of its liberation from some compound.

Neutralize.—The act of rendering a solution neither acid nor alkaline.

Nitrates.—Salts formed by the union of nitric acid with a base, or with HNO_6, and a basic radical.

Oxacids.—All acids containing oxygen.

Pellicle.—A thin skin which forms on solutions when heated.

Phosphates.—Salts formed by the union of phosphoric acid with any base.

Precipitate.—The substance dissolved in a fluid, which falls to the bottom on the addition of a reagent, and is left on the filter in the act of filtration.

Pulverize.—To reduce to powder.

Radical.—A term applied to the root or foundation of compounds. A simple radical is synonomous with an elementary body. Chlorine is the radical of hydrochloric acid.

Sublimation.—A process by which volatile bodies are condensed by cold into a solid form. The soot of chimneys is a familiar instance of this kind.

Sulphates.—Salts formed by the union of sulphuric acid with any base, or with HSO_4, and a basic radical.

Synthesis.—The forming of a compound by the union of its elements, as contrasted with analysis.

Triturate.—To rub together.

INDEX.

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