UNIT 1

Some Basic Concepts in Chemistry

SYLLABUS

Importance of studying chemistry.
Physical quantities and their SI units, dimensional analysis, Precision and significant figures.
Classification of matter.
Laws of chemical combination.
Dalton's Atomic Theory.
Mole concept, atomic, molecular and molar masses.
Percentage composition and molecular formula.
Stoichiometry of chemical reactions.

SECTION—I

GENERAL INFORMATION AND IMPORTANCE OF STUDYING CHEMISTRY

1.1. Introduction

Chemistry is defined as that branch of science which deals with the study of composition, structure and properties of matter and the changes which the matter undergoes under different conditions and the laws which govern these changes.

What does the term 'matter' mean in the above definition ? Though it will be discussed in detail later in this unit, at the moment it is sufficient to mention that all substances (solid, liquid or gaseous) which are present in our body or present around us which may be visible or some of which may not be visible (*e.g.* air) are made of **matter.** Further, two basic units of which all substances are made up are **atoms** and **molecules.** They combine with each other according to certain laws, called **Laws of chemical combination.** The changes which they undergo under different conditions may be physical or chemical changes (called **chemical reactions**). All measurements involving mass, volume, length, density etc. require **precision** or **accuracy.** Further, as chemistry deals with all substances present in our body and all those present around us, it has a great **importance** and broad **applications.**

The aim of this unit is to study all these basic concepts, some of them briefly whereas some others in detail.

1.2. Importance of Studying Chemistry

Chemistry plays a very important role in our daily life. It has helped us to meet all our requirements for a better life such as food, good health, comforts etc. Without the knowledge of chemistry, our life would have been very dull and dreary.

Some of the important applications of chemistry are given below :

(1) **Supply of food.** With increase in population, the need for the overall amount of food has increased manifold. Moreover with increase in the standards of life, there has been increase in the quality and variety of food. Chemistry has helped to achieve these goals as follows :

(*i*) It has provided *chemical fertilizers* such as urea, calcium superphosphate, sodium nitrate, ammonium sulphate etc. which have increased the yield of crops and fruits.

(*ii*) It has helped to protect the crops from insects and harmful bacteria by the use of certain effective *insecticides, fungicides* and *pesticides*.

(*iii*) The use of *preservatives* has helped to preserve food products like jam, butter, squashes etc. for longer periods.

(*iv*) It has given methods *to test the presence of adulterants* thereby ensuring the supply of pure food-stuff.

(2) **Contribution to better health and sanitation.** Chemistry has contributed towards better health and sanitation in a number of ways as follows :

(*i*) It has provided mankind with a large number of *life saving drugs*. Today dysentery and pneumonia are curable due to discovery of sulpha drugs and penicillin. Likewise epidemics such as cholera, plague and small-pox are now things of the past.

Similarly, life saving drugs like *cis platin* and *taxol* have been found to be very effective for cancer therapy and AZT (Aziodothymidine) is used for AIDS victims. These compounds are obtained from animals and plants or by synthetic methods.

(*ii*) *Analgesics* have reduced pain of different types. *Antibiotics* (like chloromycetin, streptomycin etc.) have helped to curb infection and cure from diseases like typhoid and tuberculosis. *Tranquilizers* have helped to reduce tension and bring about calm and peace to patients suffering from mental diseases. *Antiseptics* such is dettol are used to stop infection of the wounds. *Disinfectants* such as phenol are used to kill the micro-organisms present in drains, toilets, floors etc. A low concentration of chlorine *i.e.* 0·2 to 0·4 parts per million parts (ppm) is used for sterilization of water to make it fit for drinking purposes.

(*iii*) Discovery of *anaesthetics* has made surgical operations more and more successful.

(*iv*) The use of *insecticides* such as DDT and Gammexane has reduced the hazards of diseases caused by rats, mosquitoes and flies.

(*v*) New and *more effective medicines* are being discovered from time to time which are replacing the older less effective medicines. For example, quinine has been replaced by more effective antimalarials such as chloroquin, camoquin and primaquin.

(*vi*) *Synthetic vitamins* and *tonics* have contributed significantly towards better health.

In fact, the use of more effective medicines, vitamins etc. and better sanitary conditions have helped to *increase the average life span*.

(3) **Saving the environment.** Refrigerants like chlorofluorocarbons (CFCs) which destroy the ozone layer have been replaced by environment-friendly chemicals. However, Green house gases like CH_4, CO_2 etc. are still posing a challenge to the chemists.

(4) **Increase in comforts, pleasures and luxuries.** Chemistry has placed a large number of utility goods at our disposal which have added to our comforts, pleasures and luxuries. A few of these are given below :

(*i*) *Synthetic fibres (cloth)*. In addition to the natural fibres like cotton, wool, silk etc., chemistry has helped in the production of synthetic fibres such as terylene, nylon, rayon etc. which are more comfortable, durable and attractive. They are easy to wash, dry quickly and do not need ironing. Further, chemistry has provided a large number of synthetic dyes which impart bright and fast colour to the clothes. It has also provided chemicals to make these clothes fire-proof and water-proof, if necessary.

(*ii*) *Building materials*. By supplying steel and cement, chemistry has helped in the construction of safer homes and multi-storeyed buildings and dams and bridges which can last for centuries.

(*iii*) *Supply of metals*. Metals like gold, silver, copper, iron, aluminium, zinc and a large number of their alloys are used for making ornaments, utensils, coins and many industrial and agricultural implements.

(*iv*) *Articles of domestic use*. Chemistry has made our homes more comfortable by supplying a large number of articles of domestic use such as detergents, oils and fats, sugar, paper, glass, unbreakable plastic wares, paints, varnishes, cosmetics, perfumes, cooking gas etc. Electroplating has given us nickel, chromium, silver and gold-plated articles which are more attractive, durable and are not prone to rusting. By use of refrigerants like ammonia, liquid sulphur dioxide and freon, we can make our homes cooler in summer.

(*v*) *Entertainment*. Cinema, one of the common sources of entertainment and also video-cameras as well as simple cameras make use of films which are made of celluloid (a chemical compound) and coated with suitable chemicals. Similarly, fire-works which amuse us on festivals and marriages are chemical products. Phonograph records used for listening music are made of polyvinyl chloride, a chemical compound.

(5) **Transport and communication.** Almost all means of transport including automobiles (scooter, cars, trucks, buses etc.) aeroplanes, helicopters, railways use either petrol or diesel (petroleum products) or coal which are all chemical products. Chemistry has also given high quality fuels which have made possible the landing of Appolo on the moon and Viking on the Mars. Knowledge of chemistry has also helped in the development of telephone and telegraph as important means of communication.

(6) **Nuclear or Atomic energy.** In view of the decreasing coal and petroleum resources, the world will soon be facing an energy crisis. Chemistry has come to the rescue by providing an alternate source of energy which is nuclear energy. A chemical process for the production of the compound, uranium hexafluoride, made possible the enrichment of U-235 which was used first for making atomic bomb and later in the nuclear reactor for production of electricity. Nuclear energy can also be employed for digging tunnels, blasting mountains and in mining as well.

(7) **Applications in industry.** Chemistry has played an important and useful role towards the development and growth of a number of industries *e.g.* glass, cement, paper, textiles, leather, dye, paints, pigments, petroleum, sugar, plastics, pharmaceuticals etc. It has also helped in the greater production of sulphuric acid, nitric acid, ammonia, hydrogenated oils etc. by providing suitable catalysts.

Similarly, it has helped in the synthesis of new materials having specific magnetic, electrical and optical properties which are used in the production of superconducting ceramics, conducting polymers, optical fibres etc.

(8) **Applications in war.** Chemistry has also increased the striking power of man in war times. It is responsible for the discovery of highly explosive substances such as TNT, nitroglycerine and dynamite, poisonous gases like mustard gas, lewisite and phosgene and many other deadly weapons such as atom bomb and hydrogen bomb.

The progress in chemistry can also cause many problems. For example, nuclear energy is useful but disposal of nuclear waste poses a serious problem to humanity. Similarly, phonograph records have added to our pleasure for listening music but they are made of polyvinyl chloride, produced from vinyl chloride which can cause liver cancer in industrial workers. Antibiotics have eliminated infectious diseases but their overuse is very harmful. Likewise, insecticides have increased the food supply but they do a lot of harm to birds, fishes and useful insects. Use of polythene bags is posing a serious threat to sewerage system. Further, chemistry has given drugs like LSD, cocain, brown sugar etc. which are proving a curse to the society.

Thus though chemistry can be regarded as greatest benefactor of humanity, yet it can prove to be a blessing or a curse for humanity depends upon the uses to which it is put and how a relative balance is maintained between benefits and problems caused by progress in chemistry.

SECTION—II

MEASUREMENTS IN CHEMISTRY

1.3. Physical Quantities

In everyday life, we come across a number of measurements *e.g.* we buy vegetables in kilograms, milk in litres, cloth in metres etc. However during scientific studies, in addition to the measurements of mass, volumes and lengths, we come across the measurement of a number of other quantities such as temperature, pressure, concentration, force, work, density etc.

All such quantities which we come across during our scientific studies are called physical quantities.

Evidently, the measurement of any physical quantity consists of two parts :

(*i*) the number, and

(*ii*) the unit.

For example, if an object weighs 4·5 kg, it involves two parts : (*i*) 4·5 *i.e.* the number and (*ii*) kg *i.e.* the unit.

Thus the main aims of this section are :

(*i*) To see how accurately or precisely the number has been expressed *i.e. concept of significant figures*.

(*ii*) To study the units of measurement *i.e. the S.I. units*.

(*iii*) To derive the units of any physical quantity and to check the accuracy of any scientific equation by seeing that the dimensions of both sides of the equation are same *i.e. concept of dimensional analysis*.

1.4. S.I. Units

A unit is defined as the standard of reference chosen to measure any physical quantity.

Since early times, different types of units of measurements have been very popular in different parts of the world *e.g.* sers, pounds etc. for mass ; miles, furlongs, yards etc. for distances. However these units are quite cumbersome because of no uniformity in the conversion factors involved *e.g.* 1 mile = 1760 yards, 1 yard = 3 feet, 1 foot = 12 inches.

In view of the difficulties mentioned above, French Academy of Science, in 1791, introduced a new system of measurements called 'metric system' in which the different units of a physical quantity are related to each other as multiples of powers of 10, *e.g.* 1 km = 10^3 m, 1 cm = 10^{-2} m etc. This system was found to be so convenient that scientists all over the world immediately adopted this system for reporting scientific data and gradually most of the countries have also switched over to this system for measurements of everyday use. India started following metric system since 1957.

The metric system, as put forward earlier, was further improved by the General Conference of Weights and Measures (*Conference Generale des Poids et Mesures*, CGPM) which met in October 1960 in France. The improved system of units has been accepted internationally and is called International System of Units or in short SI Units (for Systeme Internationale in French).

With greater accuracy in measurement, the 'unit' definitions and hence the system of units is improved from time to time. To maintain uniformity all over the world, each nation has National Metrology Institute (NMI). In India, this responsibility has been assigned to National Physical Laboratory (NPL), New Delhi.

Seven Basic Units. The seven basic physical quantities on which the International System of Units is based, their symbols, the names of their units (called the base units) and the symbols of these units are given in Table 1.1.

TABLE 1.1. Seven basic physical quantities and their SI units

PHYSICAL QUANTITY	SYMBOL	SI UNIT	SYMBOL
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Thermodynamic temperature	T	kelvin	K
Amount of the substance	n	mole	mol
Luminous intensity	I_v	candela	cd

The SI unit of mass viz. **kilogram** has been defined as the *mass of platinum-iridium* (Pt-Ir) *cylinder that is stored in an air-tight jar at International Bureau of Weights and Measures in France.*

The SI unit of length viz **metre** was originally defined as the *length between two marks on a Pt-Ir bar kept at a temperature of 0°C* (273 K). However, now it has been redefined by CGPM as the *length of the path travelled by light in vacuum during a time interval of 1/299,792,458 of a second.*

Though the S.I. unit of temperature, 'kelvin', yet it is very common to express temperatures in degree celsius (°C). The two are related to each other as

Temperature in degree kelvin (K) = Temperature in °C + 273·15.

However, it may be remembered that size of 1°K = size of 1°C.

Similarly, though the S.I. unit of length is metre, yet it is very common to express length in Angstrom (Å) or nanometres (nm) or picometres (pm). These are related to S.I. unit as follows :

$$1\text{ Å} = 10^{-10}\,m,\quad 1\text{ nm} = 10^{-9}\,m,\quad 1\text{ pm} = 10^{-12}\,m$$

Further, it is important to understand that the **term 'weight' should not be used in place of 'mass'.** They have different meaning. Mass is the quantity of matter contained in the sample and for the given sample, it is constant and does not depend upon the place. Weight is the force with which the body is attracted towards the earth (W = mg). Thus, it depends upon the acceleration due to gravity 'g' which varies from place to place.

Derived units. The units of all other physical quantities are derived out of those of the basic physical quantities. The units thus obtained are called the derived units. Some commonly used physical quantities and their derived units are given in Table 1.2.

TABLE 1.2. Some commonly used physical quantities and their derived units

PHYSICAL QUANTITY	DEFINITION	UNIT	SYMBOL
Area	Length square	Square metre	m^2
Volume	Length cube	Cubic metre	m^3
Density	Mass/unit vol.	Kilogram per cubic metre	$kg\,m^{-3}$
Velocity	Distance/unit time	Metre per second	$m\,s^{-1}$
Acceleration	Speed change/unit time	Metre per second per second	$m\,s^{-2}$
Force	Mass × acceleration	Newton	$N = kg\,m\,s^{-2}$*
Pressure	Force/unit area	Pascal (Newton per sq. metre)	$Pa = N\,m^{-2} = kg\,m^{-1}\,s^{-2}$
Work, energy	Force × Distance	joule	$J = N\,m = kg\,m^2\,s^{-2}$**
Frequency	Cycles/sec.	Hertz	$Hz = s^{-1}$
Electric charge	Current × time	Coulomb	$C = A\,s$
Potential difference	—	Volt	$V = kg\,m^2\,s^{-3}\,A^{-1} = JA^{-1}\,s^{-1} = JC^{-1}$
Electric resistance	Pot-diff./current	ohm	$\Omega = V\,A^{-1}$
Electric conductance	Reciprocal of resistance	ohm^{-1}	$\Omega^{-1} = A\,V^{-1}$

Subsidiary Units. Quite often we require units that may be multiples or fractions of the base units. The SI system recommends the multiples such as 10^3, 10^6, 10^9 etc. and fractions such as 10^{-3}, 10^{-6}, 10^{-9} etc. *i.e.* the powers are the multiples of 3. These are indicated by special prefixes. These alongwith some other fractions or multiples in common use, alongwith their prefixes are given in Table 1.3 below :

*Newton is defined as the force that gives a mass of 1 kg an acceleration of 1 m s^{-2} so that f = ma = (1 kg) (1 m s^{-2}) = 1 kg m s^{-2} = 1 N.

**Joule is the work done when a displacement of 1 metre takes place by a force of 1 newton so that W = f×d = (1 N) (1 m) = (1 kg m s^{-2}) (1 m) = 1 kg m^2 s^{-2} = 1 J. Thus 1 J = 1 Nm = 1 kg m^2 s^{-2}.

TABLE 1.3. Some commonly used prefixes with the base units

PREFIX	SYMBOL	MULTIPLYING FACTOR	EXAMPLE
deci	d	10^{-1}	1 decimetre (dm) = 10^{-1} m = 0·1 m
centi	c	10^{-2}	centimetre (cm) = 10^{-2} m = 0·01 m
milli	m	10^{-3}	1 millimetre (mm) = 10^{-3} m
micro	μ	10^{-6}	1 micrometre (μm) = 10^{-6} m
nano	n	10^{-9}	1 nanometre (nm) = 10^{-9} m
pico	p	10^{-12}	1 picometre (pm) = 10^{-12} m
femto	f	10^{-15}	1 femtometre (fm) = 10^{-15} m
atto	a	10^{-18}	1 atto metre (am) = 10^{-18}m
zepto	z	10^{-21}	1 zepto metre (zm) = 10^{-21}m
yocto	y	10^{-24}	1 yocto metre (ym) = 10^{-24}m
deka	da	10^{1}	1 dekametre (dam) = 10^{1} m = 10 m
kilo	k	10^{3}	1 kilometer (km) = 10^{3} m = 1000 m
mega	M	10^{6}	1 megametre (Mm) = 10^{6} m
giga	G	10^{9}	1 gigametre (Gm) = 10^{9} m
tera	T	10^{12}	1 terametre (Tm) = 10^{12} m
peta	P	10^{15}	1 petametre (Pm) = 10^{15} m
exa	E	10^{18}	1 exametre (Em) = 10^{18} m
zetta	Z	10^{21}	1 zetta metre (Zm) = 10^{21}m
yotta	Y	10^{24}	1 yotta metre (Ym) = 10^{24}m

As volume is very often expressed in litres, it is important to note that the equivalence in SI units is

1 litre (1L) = 1 dm^3

and 1 millilitre (1 ml) = 1 cm^3*

Some Important Points to Remember About S.I. Units

1. The unit named after a scientist is started with a small letter and not with a capital letter *e.g.* unit of force is written as newton and not as Newton.

Likewise unit of heat or work is written as joule and not as Joule.

2. Symbols of the units do not have a plural ending like '*s*'. For example we have 10 cm and not 10 cms.

3. Words and symbols should not be mixed e.g. we should write either joules per mole or J mol^{-1} and not joules mol^{-1}.

4. Prefixes are used with the basic units *e.g.* kilometer means 1000 m (because meter is the basic unit).

Exception. Though kilogram is the basic unit of mass, yet prefixes are used with gram because in kilogram, kilo is already a prefix.

5. A unit written with a prefix and a power is a power for the complete unit *e.g.* cm^3 means (centimeter)^3 and not centi (meter)^3.

*From 1901 to 1964, a litre was defined as the volume of 1 kg of water at 4°C. During this period, a millilitre was very slightly larger than a cubic centimetre. In 1964, the litre was redefined as exactly equal in volume to 1000 cubic centimetres, thereby removing the confusion.

1.5. Precision and Significant Figures

1.5.1. Difference between Precision and Accuracy. The measurement involving counting of whole numbers of identifiable objects *e.g.* eggs, bananas, tables, chairs etc. can be known accurately *i.e.* they consist of *exact numbers*. Similarly, defined quantities are also exact. For example, there are exactly 60 seconds in exactly 1 minute. But many scientific measurements involving some measuring devices cannot be known accurately. The accuracy of any such measurement depends upon

(*i*) the accuracy of the measuring device used, and (*ii*) the skill of its operator

For example, suppose we measure the length of a room using a tape or by pacing it off. If we pace it off, we are likely to get different count each time or to have a fraction of a pace left over. Thus the result of measurement is not exactly correct. Accuracy is clearly related to the way the measurement is done. Using a tape, measure is more accurate than pacing off a length. One method to judge the quality of measurement is to repeat it. Since each measurement is likely to give slightly different result, take the average value. *If the average value of different measurements is close to the correct value, the measurement is said to be accurate (the individual measurements may not be close to each other).*

If the values of different measurements are close to each other and hence close to their average value, the measurement is said to be precise. (The average value of different measurements may not be close to the correct value). The precision depends upon the measuring device as well as the skill of the operator.

For example, suppose the actual length of the room is 10·5 *m*. Four different persons report the result of their five measurements as follows* :

Measurement (*m*)	1	2	3	4	5	Average (*m*)
Person A	10·3	10·4	10·5	10·6	10·7	10·5
Person B	10·0	10·1	10·2	10·3	10·4	10·2
Person C	10·1	10·3	10·5	10·7	10·9	10·5
Person D	10·0	10·7	10·9	11·1	11·3	10·8

Measurement by person A is both accurate and precise.

Measurement by person B has poor accuracy but good precision.

Measurement by person C has poor precision but good accuracy (just by chance/luck)

Measurement by person D has poor accuracy and poor precision.

These results may be represented diagrammatically as shown in Fig. 1.1.

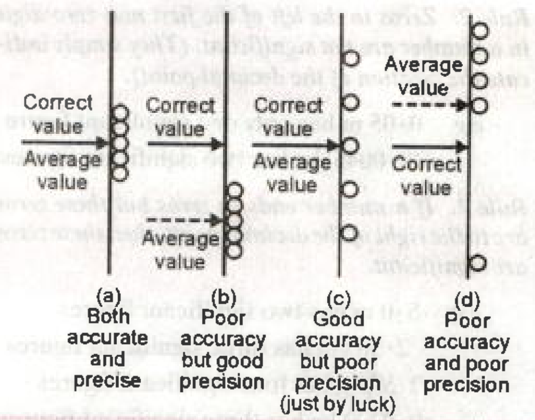

FIGURE 1.1. Understanding accuracy and precision.

A measurement can have a good accuracy but poor precision because different measurements may give a correct average (though generally poor precision corresponds to poor accuracy). The converse is not true. Good precision does not necessarily mean good accuracy. To give one example, suppose we mistake 2 kg weights for 1 kg weights. Our precision may be excellent but our accuracy will be very poor. Such errors in measurement when the same mistake is made repeatedly are called **systematic errors.** They do not affect the precision but they often affect the accuracy of a measurement.

1.5.2. Significant Figures. In the light of what has been discussed above, it is necessary to indicate in the reported result how accurately it has been measured. For example, suppose the mass of an object has been determined to be 14·5678 g. If the accuracy of the analytical balance used is 0·0001 g,

*In the Table given above, five measurements of any person have been shown as different. In actual practice some of these may be same also. Further, the values have been arranged in ascending order. In fact, the results of different measurements may not have any order.

this means that the actual mass of the object is $14 \cdot 5678 \pm 0 \cdot 0001$ g *i.e.* it lies between $14 \cdot 5677$ and $14 \cdot 5679$. Thus in the expressed mass, the first five digits are certain but the last involves an uncertainty of $\pm$ 1. All measured quantities are reported in such a way that only the last digit is uncertain (usually by $\pm$ 1).

The total number of digits in a number including the last digit whose value is uncertain is called the number of significant figures.

For example, in the above value *i.e.* $14 \cdot 5678$g, there are *six* significant figures. Similarly, the number $2 \cdot 5$ has two significant figures while the number $2 \cdot 014$ has four significant figures.

1.5.3. Rules for determining the number of significant figures. The following rules are applied in determining the number of significant figures in any reported quantity :

Rule 1. All non-zero digits as well as the zeros between the non-zero digits are significant.

e.g. 576 cm has three significant figures
$0 \cdot 48$ g has two significant figures
5004 has four significant figures
$2 \cdot 05$ has three significant figures

Rule 2. Zeros to the left of the first non-zero digit in a number are not significant. (They simply indicate the position of the decimal point).

e.g. $0 \cdot 05$ m has only one significant figure
$0 \cdot 0045$ kg has two significant figures.

Rule 3. If a number ends in zeros but these zeros are to the right of the decimal point, then these zeros are significant.

e.g. $5 \cdot 0$ m has two significant figures
$2 \cdot 50$ cm has three significant figures
$2 \cdot 500$ g has four significant figures
$0 \cdot 0200$ kg has three significant figures.

Rule 4. If a number ends in zeros but these zeros are not to the right of a decimal point, these zeros may or may not be significant.

For example, 10500 g may have three, four or five significant figures. This ambiguity is removed by expressing the value in an exponential form. For example, the above mass may be written in three different exponential forms as follows :

$1 \cdot 05 \times 10^4$ g, which has *three* significant figures
or $1 \cdot 050 \times 10^4$ g, which has *four* significant figures
or $1 \cdot 0500 \times 10^4$ g, which has *five* significant figures

In these cases, the significant figures of only the first factor are counted (remembering that all zeros to the right of a decimal point are significant). Thus in such cases, the general notation is

$$N \times 10^n$$

where N = a number with a single non-zero digit to the left of the decimal point
and n = an integer.

The above method of expressing a number is called **Scientific** or **Exponential Notation.**

1.5.4. Rules for determining the number of significant figures in answers involving calculations. To get the final result of any experiment, usually calculations are required which involve addition, subtraction, multiplication and division of different numbers. These numbers may have different accuracies *i.e.* may contain different number of significant figures or decimal places. The final result involving these numbers, therefore, cannot be more accurate or precise than the least precise number involved in a particular calculation. The following rules are applied in determining the number of significant figures in the answer of any particular calculation :

Rule 1. The result of an addition or subtraction should be reported to the same number of decimal places as that of the term with least number of decimal places. The number of significant figures of different numbers have no role to play.

Example 1.

		$4 \cdot 523$
		$2 \cdot 3$
		$6 \cdot 24$
Actual sum	=	$13 \cdot 063$
Reported sum	=	$13 \cdot 1$*

***Rounding off.** The actual sum of $13 \cdot 063$ has been reported as $13 \cdot 1$. This is because the result is to be reported only upto one decimal place. The digit at the second decimal place is 6 which is greater than 5. Hence the digit at the first decimal place has been increased by 1. This method is called *rounding off*. The general procedure for rounding off is as follows :

(*i*) If the digit just next to the last digit to be retained is less than 5, the last digit is taken as such and all other digits on its right are dropped.

(*ii*) If the digit is greater than 5, the last digit to be retained is increased by 1 and all other digits on its right are dropped.

(*iii*) If the digit is equal to 5, the last significant figure is left unchanged if it is even and is increased by 1 if it is odd.

EXAMPLES : $1 \cdot 234 = 1 \cdot 23$, $1 \cdot 236 = 1 \cdot 24$, $1 \cdot 235 = 1 \cdot 24$, $1 \cdot 225 = 1 \cdot 22$.

First number has three decimal places, second has one and third has two. Hence answer should be reported only upto one decimal place. Note that the significant figures in the three numbers are 4, 2 and 3 respectively.

Example 2.

$$\begin{array}{lr} & 7\cdot621 \\ & 6\cdot243 \\ & 1\cdot020 \\ \text{Actual sum} = & \overline{14\cdot884} \\ \text{Reported sum} = & 14\cdot884 \end{array}$$

Each number has three decimal places. So the answer is also reported upto three decimal places. Further note that the significant figures in each of the three numbers is 4 but result has five significant figures.

Example 3.

$$\begin{array}{lr} & 154\cdot2 \\ & 6\cdot1 \\ & 23 \\ \text{Actual sum} = & \overline{183\cdot3} \\ \text{Reported sum} = & 183 \end{array}$$

The last number is an exact number (involving no decimal place). Hence answer is also reported as an exact number.

Example 4.

$$\begin{array}{lr} & 18\cdot4215 \\ & -6\cdot01 \\ \text{Actual Diff.} & \overline{12\cdot4115} \\ \text{Reported Diff.} = & 12\cdot41 \end{array}$$

As the second number has two decimal places only while the first has four, the answer is reported upto the two decimal places only.

Example 5.

$$\begin{array}{lr} & 29\cdot25 \\ & -12\cdot0234 \\ \text{Actual Diff.} = & \overline{17\cdot2266} \\ \text{Reported Diff.} = & 17\cdot23 \end{array}$$

As the first number has two decimal places only while the second has four, the answer is reported only upto two decimal places.

Rule 2. The result of a multiplication or division should be reported to the same number of significant figures as is possessed by the least precise term used in the calculation.

Example 1.

$$\begin{array}{lr} & 4\cdot327 \\ & \times\, 2\cdot8 \\ \text{Actual Product} = & \overline{12\cdot1156} \\ \text{Reported Product} = & 12 \end{array}$$

The first number has four significant figures while the second has two. The actual product has been rounded off to give a reported product of 12 *i.e.* containing two significant figures only. This is because the least precise term in the calculation (viz. 2·8) has only two significant figures.

Example 2. $0\cdot46 \div 15\cdot734$ gives

Actual quotient $= 0\cdot029236$

Reported quotient should be $0\cdot029$ containing only two significant figures because the least precise term in calculations (viz. 0·46) has two significant figures only.

It may be noted that both the above rules, in fact, may be interpreted as follows :

"The reported answer should not be more precise than the least precise term used in the calculation."

This generalization helps to check the reported answer in cases where doubt arises. For example, in example 1, the precision of the least precise term is 0·1 part in 2·8 *i.e.* 1 part in 28 or nearly 35 parts per thousand (*i.e.* 35 p.p.t). The precision of the reported answer is 1 part in 12 or nearly 83 p.p.t. Thus the reported result is not more precise than the least precise term. If we take the reported answer as 12·1, the precision will be 0·1 in 12·1 or 1 part in 121 *i.e.* nearly 8 p.p.t. which is more precise than the least precise term. Hence the reported result should be 12 and not 12·1.

Rule 3. If a calculation involves a number of steps, the result should contain the same number of significant figures as that of the least precise number involved, other than the exact numbers.

Example.

$$\frac{42\cdot967 \times 0\cdot02435}{0\cdot34 \times 4} = 0\cdot7692988$$

(Actual result)

Leaving the exact number 4, the least precise term has two significant figures. Hence after rounding off, the reported result will be $0\cdot77$ *i.e.* containing two significant figures. *Alternatively*, the above rule is applied as follows :

First the number of significant figures that the answer should contain is decided (i.e. it should be equal to that of the least precise term, other than the exact number). Before carrying out the mathematical operations, every number is rounded off to contain one significant figure more than the answer would have. The answer obtained is then rounded off to contain the required number of significant figures.

Thus in the above example, the answer should have two significant figures. Hence every number is first rounded off to contain three significant figures. Thus

$$\frac{43 \cdot 0 \times 0 \cdot 0243}{0 \cdot 340 \times 4} = 0 \cdot 7683088$$

After rounding off to two significant figures, reported answer should be $0 \cdot 77$.

REMEMBER

The rules that have been stated above apply only to non-integral measured quantities because only in these cases the uncertainty in measurement has significance. These do not apply to exact numbers where uncertainty has no significance e.g. there are exact 12 eggs in a dozen of eggs, there are exact 1000 g in 1 kg, there are exact 3 feet in 1 yard and so on.

PROBLEMS ON SIGNIFICANT FIGURES

EXAMPLE 1. *What is the difference between 5·0 g and 5·00 g ?*

Solution. Though they look to be equivalent but scientifically they are different. 5·0 g has two significant figures and hence its precision is 0·1 part in 5 *i.e.* 20 p.p.t. 5·00 has three significant figures and hence its precision is 0·01 parts in 5 *i.e.* 2 p.p.t. Hence 5·00 g is more precise measurement than 5·0 g.

EXAMPLE 2. *How many significant figures are there in each of the following numbers ?*

(i) $6 \cdot 005$

(ii) $6 \cdot 022 \times 10^{23}$

(iii) 8000

(iv) $0 \cdot 0025$

(v) π

(vi) the sum $18 \cdot 5 + 0 \cdot 4235$

(vii) the product $14 \times 6 \cdot 345$.

Solution. *(i) Four* because the zeros between the non- zero digits are significant figures.

(ii) Four because only the first term gives the significant figures and exponential term is not considered.

(iii) Four. However, if expressed in scientific notation as 8×10^3, it will have only one significant figure, as $8 \cdot 0 \times 10^3$, $8 \cdot 00 \times 10^3$ or $8 \cdot 000 \times 10^3$, it will have 2, 3 or 4 significant figures.

(iv) Two because the zeros on the left of the first non-zero digit are not significant.

(v) As $\pi = \frac{22}{7} = 3 \cdot 1428571.....$, hence it has infinite number of significant figures.

(vi) Three because the reported sum will be only upto one decimal place *i.e.* 18.9.

(vii) Two because the number with least number of significant figures involved in the calculation (*i.e.* 14) has two significant figures.

EXAMPLE 3. *Express the following to four significant figures :*

(i) $6 \cdot 45372$ *(ii)* $48 \cdot 38250$ *(iii)* 70000

(iv) $2 \cdot 65986 \times 10^3$ *(v)* $0 \cdot 004687$.

Solution. *(i)* $6 \cdot 454$ *(ii)* $48 \cdot 38$

(iii) $7 \cdot 000 \times 10^4$ *(iv)* $2 \cdot 660 \times 10^3$

(v) $0 \cdot 004687$.

EXAMPLE 4. *A sample of nickel weighs* $6 \cdot 5425$g *and has a density of* $8 \cdot 8$ g/cm^3. *What is the volume? Report the answer to correct decimal place.*

Solution. $\text{Volume} = \frac{\text{Mass}}{\text{Density}} = \frac{6 \cdot 5425 \text{ g}}{8 \cdot 8 \text{ g/cm}^3}$

$$= 0 \cdot 74 \text{ cm}^3$$

The result should have two significant figures because the least precise term ($8 \cdot 8$) has two significant figures.

EXAMPLE 5. *Express the result of the following calculation to the appropriate number of significant figures* $\frac{3 \cdot 24 \times 0 \cdot 08666}{5 \cdot 006}$ **(N.C.E.R.T.)**

Solution. $\frac{3 \cdot 24 \times 0 \cdot 08666}{5 \cdot 006} = 0 \cdot 0560883$

(Actual result)

As $3 \cdot 24$ has least number of significant figures viz 3, the result should contain 3 significant figures only. Hence the result will be reported as $0 \cdot 0561$ (after rounding off).

PROBLEMS FOR PRACTICE

1. How many significant figures are there in each of the following numbers ?

   (*i*) 6·200 (*ii*) 0·052
   (*iii*) $7 \cdot 5 \times 10^4$
   (*iv*) 0·00050
   (*v*) 67·32 – 6·3
   (*vi*) 4·2+7·589
   (*vii*) $(5 \cdot 56)^2\ (8 \cdot 24)/(3 \cdot 6)$
   (*viii*) $18 \cdot 567/(8 \cdot 1 \times 2)$

   [Ans. (*i*) 4 (*ii*) 2 (*iii*) 2 (*iv*) 2 (*v*) 3 (*vi*) 3 (*vii*) 2 (*viii*) 2]

2. What is the number of significant figures in

   (*i*) Avogadro's number ($6 \cdot 0 \times 10^{23}$) and

   (*ii*) Planck's constant ($6 \cdot 62 \times 10^{-34}$ Js)?

   [Ans. (*i*) 2 (*ii*) 3]

3. Express the number 45000 in exponential notation to show

   (*i*) two significant figures

   (*ii*) four significant figures.

   [Ans. (*i*) $4 \cdot 5 \times 10^4$ (*ii*) $4 \cdot 500 \times 10^4$]

4. Criticize the following statements :

   (*i*) A fossil was studied 25 years ago and was found to be 25000 years old. Now it must be 25025 years old.

   (*ii*) The population of a city is 605,000. A family of five persons shifts to some other city. The population of the city must now be 604,995.

HINTS FOR DIFFICULT PROBLEMS

1. (*i*) Zeros to the right of the decimal point are significant.

   (*ii*) Zeros to the left of the first non-zero digit are not significant.

   (*iii*) When expressed as $7 \cdot 5 \times 10^4$, ony significant figures of 7·5 are to be considered.

   (*iv*) Apply rules given in Hints (*i*) and (*ii*) above.

   (*v*) 67·32—6·3 = 61·02. The result is to be reported to same number of decimal places as that of the term with least number of decimal places (viz 6·3 with only one decimal place). Hence after rounding off, reported result = 61·0, (which has there significant figures).

   (*vi*) 4·2 + 7·589 = 11·789. As it is to be reported to one decimal place (as in (*v*) above), after rounding off, reported result = 11·8 (having three significant figures).

   (*vii*) As least precise term (viz 3·6) has two significant figures, the reported result should have two significant figures.

   (*viii*) Leaving exact number 2, the least precise term (8·1) has two significant figures.

2. See Hint to Q. 1. (*iii*).

3. See Hint to Q. 1. (*iii*).

4. (*i*) $25000 = 2 \cdot 5 \times 10^4$. It has only two significant figures. Expressing the new value in terms of same number of significant figures, the age of the fossil after 25 years can be taken as same.

ADD TO YOUR KNOWLEDGE

1. Quite often the uncertainty in measurement is expressed in terms of percentage by putting ± sign before it e.g. 250 ± 1% etc. If the same instrument is used for measuring different quantities, then smaller the quantity to be measured, greater is the percentage uncertainty. For example, if a balance has uncertainty in measurement equal to ± 1 mg, then if we weigh 100 *g* on it, the result can be reported as 100 ± 0·001%. If same balance is used to weigh 10 *g*, the result reported will be 10 ± 0·01%, and if 1 *g* is weighed, the result reported will be 10 ± 0·1%. Hence smaller the quantity to be measured, more precise should be the instrument.

2. While reporting the result to correct significant figures in any calculation, the exact number (*e.g.* 4 in example on page 1/10) is left out as it does not affect the number of significant figures. This is because *an exact number is considered to have an infinite number of significant figures.*

1.6. Dimensional Analysis

Any calculation involving the use of the dimensions of the different physical quantities involved is called dimensional analysis.

It is used for any one of the following purposes:

(1) To convert a physical quantity given in one type of units into some other units. It consists of the following steps :

(*i*) *First determine the 'unit conversion factor/factors'.*

e.g. for conversion of pounds (lb) into kilograms (kg) or vice versa, 1 kg = 2·205 lb.

$$\therefore \quad 1 = \frac{2\cdot205\ \text{lb}}{1\ \text{kg}} \quad \text{or} \quad 1 = \frac{1\ \text{kg}}{2\cdot205\ \text{lb}}$$

Similarly, for conversion of inches into cm or vice versa,

$$1\ \text{inch} = 2\cdot54\ \text{cm}$$

$$\therefore \quad 1 = \frac{2\cdot54\ \text{cm}}{1\ \text{inch}} \quad \text{or} \quad 1 = \frac{1\ \text{inch}}{2\cdot54\ \text{cm}}$$

The quantities such as 2·205 lb per kg or 1 kg per 2·205 lb, 2·54 cm per inch or one inch per 2·54 cm etc. are called *'unit conversion factors'*.

(*ii*) *Multiply the given physical quantity with the unit conversion factors, retaining the units of the physical quantity as well as that of the unit conversion factors in such a way that all units cancel out leaving behind only the required units*. If the unit conversion factor has not been used correctly, the answer will come out to be in wrong units.

(*iii*) *If the conversion involves a number of steps, each conversion factor is used in such a way that the units of the preceding factor cancel out.*

As an illustration, a few examples are given below :

EXAMPLE 1. *A man weighs 175 lb. Express his weight in kg. Given that 1 kg = 2·205 lb.*

Solution. 1 kg = 2·205 lb

$$\therefore \quad 1 = \frac{2\cdot205\ \text{lb}}{1\ \text{kg}} = \frac{1\ \text{kg}}{2\cdot205\ \text{lb}}$$

$$\text{Hence } 175\ \text{lb} = 175\ \cancel{\text{lb}} \times \frac{1\ \text{kg}}{2\cdot205\ \cancel{\text{lb}}}$$

$$= \mathbf{79\cdot4\ kg.}$$

The units 'lb' cancel out from the numerator and denominator and the answer is in the required units *i.e.* kg.

Note. If the conversion factor were not correctly used, the answer would have absurd units *e.g.* if we write

$$175\ \text{lb} = 175\ \text{lb} \times \frac{2\cdot205\ \text{lb}}{1\ \text{kg}}$$

the answer will have the units $\text{lb}^2\ \text{kg}^{-1}$.

EXAMPLE 2. *How many inches are there in 3·00 km ? Given that 1 km = 1000 m, 1 m = 1·094 yd, 1 yd = 36 in.*

Solution. The unit conversion factors will be

$$1 = \frac{1000\ \text{m}}{1\ \text{km}} = \frac{1\ \text{km}}{1000\ \text{m}}$$

$$1 = \frac{1\cdot094\ \text{yd}}{1\text{m}} = \frac{1\text{m}}{1\cdot094\ \text{yd}}$$

$$1 = \frac{36\ \text{in}}{1\ \text{yd}} = \frac{1\ \text{yd}}{36\ \text{in}}$$

Here the conversion involves a number of steps. Hence the unit conversion factors are applied in such a way that the units of the preceding factor cancel out. Thus

$$3\cdot00\ \text{km} = 3\cdot00\ \cancel{\text{km}} \times \frac{1000\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1\cdot094\ \cancel{\text{yd}}}{1\cancel{\text{m}}} \times \frac{36\ \text{in}}{1\ \cancel{\text{yd}}}$$

$$= \mathbf{1\cdot18 \times 10^5\ in.}$$

Note that the answer has been reported to contain three significant figures (because 1000 m and 36 in. are exact numbers).

EXAMPLE 3. *Express the following in S.I. base units using power of 10 notation (example 2·54 mm = 2·54 × 10^{-3} m)*

(a) 1·35 mm (b) 1 day (c) 6·45 mL (d) 48 µg (e) 0·0426 in **(N.C.E.R.T.)**

Solution. (*a*) S.I. unit of length is '*m*'

1 m = 100 cm, 1 cm = 10 mm

∴ Unit conversion factors are

$$\frac{1\ \text{m}}{100\ \text{cm}} = \frac{100\ \text{cm}}{1\ \text{m}} = 1, \quad \frac{1\ \text{cm}}{10\ \text{mm}} = \frac{10\ \text{mm}}{1\ \text{cm}} = 1$$

$$\therefore 1\cdot35\ \text{mm} = 1\cdot35\ \cancel{\text{mm}} \times \frac{1\ \cancel{\text{cm}}}{10\ \cancel{\text{mm}}} \times \frac{1\ \text{m}}{100\ \cancel{\text{cm}}}$$

$$= \mathbf{1\cdot35 \times 10^{-3}\ m}$$

(*b*) S.I. unit of time is 'sec' (*s*)

1 day =

$$1 \text{ day} \times \frac{24 \text{ hours}}{1 \text{ day}} \times \frac{60 \text{ min}}{1 \text{ hour}} \times \frac{60 \text{ s}}{1 \text{ min}}$$

$$= 8 \cdot 6400 \times 10^4 s$$

(*c*) S.I. unit of volume is 'm^3'

$$6 \cdot 45 \text{ mL} = 6 \cdot 45 \text{ mL} \times \frac{1 \text{ cm}^3}{1 \text{ mL}}$$

$$\times \frac{1 \text{ m}}{100 \text{ cm}} \times \frac{1 \text{ m}}{100 \text{ cm}} \times \frac{1 \text{ m}}{100 \text{ cm}}$$

$$= \mathbf{6 \cdot 45 \times 10^{-6} \text{ m}^3}$$

(*d*) S.I. unit of mass is 'kg'

$$48 \,\mu g = 48 \,\mu g \times \frac{10^{-6} g}{1 \,\mu g} \times \frac{1 \text{ kg}}{1000 \text{ g}}$$

$$= \mathbf{4 \cdot 8 \times 10^{-8} \text{ kg}}$$

(*e*) S.I. unit of length is '*m*'

$$0 \cdot 0426 \text{ in} = 0 \cdot 0426 \text{ in} \times \frac{2 \cdot 54 \text{ cm}}{1 \text{ in}} \times \frac{1 \text{ m}}{100 \text{ cm}}$$

$$= 1 \cdot 082 \times 10^{-3} \text{ m}$$

(2) In solving problems. For this purpose, units are written alongwith all the numbers. The units are then cancelled in the same manner as the numbers. If the problem has been correctly solved, the answer will have correct units.

EXAMPLE. 1. *What is the mass (in grams) of an aluminium block whose dimensions are 2·0 in. × 3·0 in. × 4·0 in. and whose density is 2·7 g/cm³ ? Given that 1 in. = 2·54 cm.*

Solution. Here unit conversion factors are

$$1 = \frac{2 \cdot 54 \text{ cm}}{1 \text{ in.}} = \frac{1 \text{ in}}{2 \cdot 54 \text{ cm}}$$

and $$1 = \frac{2 \cdot 7 \text{ g}}{1 \text{ cm}^3} = \frac{1 \text{ cm}^3}{2 \cdot 7 \text{ g}}$$

Hence required mass (in g)

$$= 2 \cdot 0 \text{ in} \times 3 \cdot 0 \text{ in} \times 4 \cdot 0 \text{ in} \times \frac{2 \cdot 54 \text{ cm}}{1 \text{ in}}$$

$$\times \frac{2 \cdot 54 \text{ cm}}{1 \text{ in}} \times \frac{2 \cdot 54 \text{ cm}}{1 \text{ in}} \times \frac{2 \cdot 7 \text{ g}}{1 \text{ cm}^3}$$

$$= \mathbf{1 \cdot 1 \times 10^3 \text{ g.}}$$

EXAMPLE 2. *The mass of precious stones is expressed in terms of 'carat'. Given that 1 carat = 3·168 grains and 1 gram = 15·4 grains, calculate the total mass of a ring in grams and kilograms which contains 0·500 carat diamond and 7·00 gram gold.*

(N.C.E.R.T.)

Solution. The unit conversion factors to be used will be

$$1 = \frac{1 \text{ carat}}{3 \cdot 168 \text{ grains}} = \frac{3 \cdot 168 \text{ grains}}{1 \text{ carat}}$$

$$1 = \frac{1 \text{ gram}}{15 \cdot 4 \text{ grains}} = \frac{15 \cdot 4 \text{ grains}}{1 \text{ gram}}$$

$$0 \cdot 500 \text{ carat} = 0 \cdot 500 \text{ carat} \times \frac{3 \cdot 168 \text{ grains}}{1 \text{ carat}}$$

$$\times \frac{1 \text{ gram}}{15 \cdot 4 \text{ grains}}$$

$$= 0 \cdot 10 \text{ gram}$$

∴ Total mass of the ring

$$= 7 \cdot 00 + 0 \cdot 10 \, g = 7 \cdot 10 \, g$$

$$= 7 \cdot 10 \, g \times \frac{1 \text{ kg}}{1000 \, g} = \mathbf{0 \cdot 0071 \text{ kg}}.$$

(3) To check the accuracy of any equation. This is done by finding the dimensions of both sides of the equation which must be same.

EXAMPLE *Using the method of dimensional analysis, verify the validity of the following equations:*

(*i*) $c = \nu \lambda$ (*ii*) $S = ut + \frac{1}{2} at^2$

Solution. (*i*) Dimensions of L.H.S.

i.e. $c = ms^{-1}$

Dimension of R.H.S. *i.e.*

$$\nu \lambda = s^{-1} \times m = ms^{-1}$$

Hence from dimensions point of view, the equation is correct.

(*ii*) Dimensions of L.H.S. i.e. S = metres (m)

Dimensions of R.H.S. i.e. $ut + \frac{1}{2} at^2$

$$= ms^{-1} \times s + \frac{1}{2} m \, s^{-2} \times s^2$$

$$= m + m = \text{metres } (m)$$

Hence from dimensions point of view, the equation is correct.

PROBLEMS FOR PRACTICE

1. Convert 16·1 km to miles using the following units equivalents :
   1 km = 1000 m, 1 ft = 12 inches
   1 m = 100 cm, 1 mile = 1760 yd
   1 inch = 2·54 cm, 1 yd = 3 ft [Ans. **10·0 miles**]
2. What is the weight in pounds of a gold bar 12·0 inches long, 6·00 inches wide and 3·00 inches thick? The density of gold is 19·3 $g\ cm^{-3}$. Given 1 inch = 2·54 cm, 1 lb = 453·6 g. [Ans. **151 lb**]
3. Express the following in SI units :
   (*i*) 5′6″, the average height of an Indian man.
   (*ii*) 80 miles per hour, the average speed of a roadways bus.
   (*iii*) 100 pounds, the average weight of an Indian girl. (Take 1 lb = 454 *g*)
   (*iv*) – 10°C, the lowest temperature in Simla.
   (*v*) 2 litres of milk, the average consumption of a family of 4 persons.
   (*vi*) 14 pounds per square inch (atmospheric pressure)
   [Ans. **(*i*) 1·68 m (*ii*) 35·8 ms^{-1} (*iii*) 45·4 kg (*iv*) 263·15 K (*v*) 2 dm^3 (*vi*) 9841·3 kg m^{-2}**]
4. Convert the following into kilograms :
   (*i*) 500 Mg (mass of loaded jumbo jet)
   (*ii*) 1 fg (mass of human DNA molecule) (*N.C.E.R.T.*)
   [Ans. **(*i*) 5 × 10^{15} kg (*ii*) 10^{-18} kg**]
5. Convert the following into metre
   (*i*) 40 Em (thickness of Milky way galaxy)
   (*ii*) 1·4 Gm (diameter of Sun)
   (*iii*) 41 Pm (distance of nearest star) (*N.C.E.R.T.*)
   [Ans. **(*i*) 4 × 10^9 *m* (*ii*) 1·4 × 10^9 *m* (*iii*) 41 × 10^{15} m**]
6. Using the unit conversion factors, express
   (*i*) 1·54 mm s^{-1} into pm μs^{-1}
   (*ii*) 2·66 *g* cm^{-3} to $\mu g\ \mu m^{-3}$. (*N.C.E.R.T.*)
   [Ans. **(*i*) 1·54×10^3 pm μs^{-1} (*ii*) 2·66×10^{-6} $\mu g\ \mu m^{-3}$**]
7. Vanadium metal is added to steel to impart strength. The density of vanadium is 5·96 *g*/cm^3. Express this in S.I. units (kg/m^3). (*N.C.E.R.T.*)
   [Ans. **5960 kg/m^3**]
8. "The star of India" sapphire weighs 563 carats. If one carat is equal to 200 mg, what is the weight of the gemstone in grams ? (*N.C.E.R.T.*)
   [Ans. **112·6 g**]

HINTS FOR DIFFICULT PROBLEMS

1. $$16\cdot1\ km = 16\cdot1\ \cancel{km} \times \frac{1000\ m}{1\ \cancel{km}} \times \frac{100\ \cancel{cm}}{1\ \cancel{m}} \times \frac{1\ \cancel{inch}}{2\cdot54\ \cancel{cm}} \times \frac{1\ \cancel{ft}}{12\ \cancel{inches}} \times \frac{1\ \cancel{yd}}{3\ \cancel{ft}} \times \frac{1\ mile}{1760\ \cancel{yd}}$$
   = 10·0 miles.
   Note that the least precise term (2·54) has three significant figures.
2. $$\text{Weight in pounds} = 12\cdot0\ \cancel{in} \times 6\cdot00\ \cancel{in} \times 3\cdot00\ \cancel{in} \times \frac{2\cdot54\ \cancel{cm}}{1\ \cancel{in}} \times \frac{2\cdot54\ \cancel{cm}}{1\ \cancel{in}} \times \frac{2\cdot54\ \cancel{cm}}{1\ \cancel{in}} \times \frac{19\cdot3\ \cancel{g}}{1\ \cancel{cm^3}} \times \frac{1\ lb}{453\cdot6\ \cancel{g}}$$
   = 150·60491 *lb* = 151 *lb* (after rounding off because least precise term has 3 significant figures)
3. (*i*) $$5'6'' = 66'' = 66\ \cancel{in} \times \frac{2\cdot54\ \cancel{cm}}{1\ \cancel{in}} \times \frac{1\ m}{100\ \cancel{cm}} = 1\cdot6764\ m = 1\cdot68\ m$$
   (after rounding off to have 3 significant figures as that of the least precise term)

   (*ii*) $$\frac{80\ miles}{1\ hour} = \frac{80\ \cancel{miles}}{3600\ s} \times \frac{1760\ \cancel{yd}}{1\ \cancel{mile}} \times \frac{3\ \cancel{ft}}{1\ \cancel{yd}} \times \frac{12\ \cancel{in}}{1\ \cancel{ft}} \times \frac{2\cdot54\ \cancel{cm}}{1\ \cancel{in}} \times \frac{1\ m}{100\ \cancel{cm}}$$
   $$= 35\cdot76\ ms^{-1} = 35\cdot8\ ms^{-1}$$
   (*iii*) $$100\ lb = 100\ \cancel{lb} \times \frac{454\ \cancel{g}}{1\ \cancel{lb}} \times \frac{1\ kg}{1000\ \cancel{g}} = 45\cdot4\ kg$$
   (*iv*) –10°C = –10 + 273·15 K = 263·15 K
   (*v*) 2 L
   $$= 2\ \cancel{L} \times \frac{1000\ \cancel{cm^3}}{1\ \cancel{L}} \times \frac{1\ dm}{10\ \cancel{cm}} \times \frac{1\ dm}{10\ \cancel{cm}} \times \frac{1\ dm}{10\ \cancel{cm}}$$
   $$= 2\ dm^3$$
   (*vi*) $$\frac{14\ \cancel{lb}}{\cancel{1 inch^2}} \times \frac{1\ kg}{2\cdot205\ \cancel{lb}} \times \frac{\cancel{1\ inch}}{2\cdot54\ cm} \times \frac{\cancel{1\ inch}}{2\cdot54\ cm} \times \frac{100\ \cancel{cm}}{1\ m} \times \frac{100\ \cancel{cm}}{1\ m}$$
   = 9841·3 kg per square metre.

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4. (*i*) 1 Mg *i.e.* megagram $= 10^6\ g$

$\therefore\quad 500\ \text{Mg} = 500 \times 10^6\ g = 5 \times 10^5$ **kg.**

(*ii*) 1 fg *i.e.* femto gram $= 10^{-15}\ g = 10^{-18}$ **kg**

5. (*i*) 1 Em *i.e.* exametre $= 10^{18}\ m$

$\therefore\quad 40\ \text{Em} = 40 \times 10^{18}\ m$

$= 4 \times 10^{19}\ m.$

(*ii*) 1 Gm *i.e.* gigametre $= 10^9\ m$

$\therefore\quad 1\cdot 4\ \text{Gm} = 1\cdot 4 \times 10^9\ m$

6. (*i*) $$\frac{1\cdot 54\ \cancel{\text{mm}}}{1\ \cancel{s}} \times \frac{1\ \cancel{\text{cm}}}{10\ \cancel{\text{mm}}} \times \frac{1\ \cancel{m}}{100\ \cancel{\text{cm}}} \times \frac{1\ \text{pm}}{10^{-12}\ \cancel{m}} \times \frac{10^{-6}\ \cancel{s}}{1\ \mu s}$$

$$= 1\cdot 54 \times 10^3\ \text{pm}\ \mu s^{-1}$$

(*ii*) $$\frac{2\cdot 66\ \cancel{g}}{1\ \text{cm}^3} \times \frac{1\ \mu g}{10^{-6}\ \cancel{g}} \times \frac{100\ \cancel{\text{cm}}}{1\ \cancel{m}} \times \frac{100\ \cancel{\text{cm}}}{1\ \cancel{m}} \times \frac{100\ \cancel{\text{cm}}}{1\ \cancel{m}} \times \frac{10^{-6}\ \cancel{m}}{1\ \mu m} \times \frac{10^{-6}\ \cancel{m}}{1\ \mu m} \times \frac{10^{-6}\ \cancel{m}}{1\ \mu m}$$

$$= 2\cdot 66 \times 10^{-6}\ \mu g\ \mu m^{-3}.$$

7. $$\frac{5\cdot 96\ \cancel{g}}{1\ \cancel{\text{cm}^3}} \times \frac{1\ \text{kg}}{1000\ \cancel{g}} \times \frac{100\ \cancel{\text{cm}}}{1\ m} \times \frac{100\ \cancel{\text{cm}}}{1\ m} \times \frac{100\ \cancel{\text{cm}}}{1\ m} = 5960\ \text{kg}/m^3$$

8. $$563\ \cancel{\text{carats}} \times \frac{200\ \cancel{\text{mg}}}{1\ \cancel{\text{carat}}} \times \frac{1\ g}{100\ \cancel{\text{mg}}} = 112\cdot 6\ g$$

SOME USEFUL CONVERSION FACTORS

(*a*) From given unit to another unit

1 mile = 1760 yards	1 metric ton = 1000 kg
1 yard = 3 feet	1 kg = 1000 g
1 foot = 12 inches	1g = 1000 mg
1 inch = $2\cdot 54$ cm	1 lb = $453\cdot 6$ g
1 litre = 1000 ml = 1000 cm^3	
1 quart = $0\cdot 9463$ litre	
or 1 litre = $1\cdot 056$ quarts	

(*b*) From given unit to S.I. unit

1 Å = 10^{-10} m	1 atm = 760 mm or torr
	= 101, 325 Pa or Nm^{-2}
	= $1\cdot 013 \times 10^6$ dynes/cm^2
	1 bar = $10^5\ \text{Nm}^{-2} = 10^5$ Pa
1 a.m.u. = $1\cdot 66053 \times 10^{-27}$ kg	1 mm or 1 torr = $133\cdot 322$ Pa or Nm^{-2}
t°C = $t + 273\cdot 15$ K	
$\simeq t + 273$ K	1 calorie = $4\cdot 184$ J
1 litre = $10^{-3}\ m^3 = 1\ \text{dm}^3$	1 erg = 10^{-7} J
1 dyne = 10^{-5} N	1 electron volt (eV) = $1\cdot 6022 \times 10^{-19}$ J.

ADD TO YOUR KNOWLEDGE

1. The SI unit of temperature, viz. Kelvin was given in honour of the great British scientist 'Lord Kelvin'.
2. Temperatures in Kelvin are expressed as 25 K, 35 K etc. and not as 25°K, 35°K etc.
3. Fermi is a unit of length used in Nuclear Physics. 1 Fermi = 10^{-13} cm = 10^{-15} *m*.

SECTION—III

CLASSIFICATION OF MATTER

1.7. What is Matter ?

Everything that is around us and the presence of which can be felt with the help of any of our five senses *i.e.* sight, touch, smell, hearing and taste is called matter. In fact, this whole universe is made up of only two things viz. matter and energy. Experience has shown that all types of matter possess mass and occupy space. Hence

Matter is defined as any thing that occupies space, possesses mass and the presence of which can be felt by any one or more of our five senses.

The examples of matter are innumerable. A few of these include clothes, iron, gold, plastics, wood, water, milk, petrol, kerosene oil, air etc.

1.8. Classification of Matter

There are two ways of classifying matter :

(A) Physical Classification

(B) Chemical Classification.

(A) Physical Classification. Based on *physical state* under ordinary conditions of temperature and pressure, matter is classified into the following three types :

(1) Solids (2) Liquids (3) Gases.

A substance is said to be solid if it possesses a definite volume and a definite shape e.g. sugar, iron, gold, wood etc.

A substance is said to be liquid, if it possesses a definite volume but no definite shape. They take up the shape of the vessel in which they are put. e.g. water, milk, oil, mercury, alcohol etc.

A substance is said to be gaseous if it neither possesses a definite volume nor a definite shape. This is because they fill up the whole vessel in which they are put. e.g. hydrogen, oxygen, carbon dioxide, air etc.

(B) Chemical Classification. Broadly speaking, all kinds of matter may be classified into the following two types :

(1) Homogeneous (2) Heterogeneous

The word 'material' is commonly used for all kinds of matter whether homogeneous or heterogeneous.

A material is said to be homogeneous if it has uniform composition and identical properties throughout.

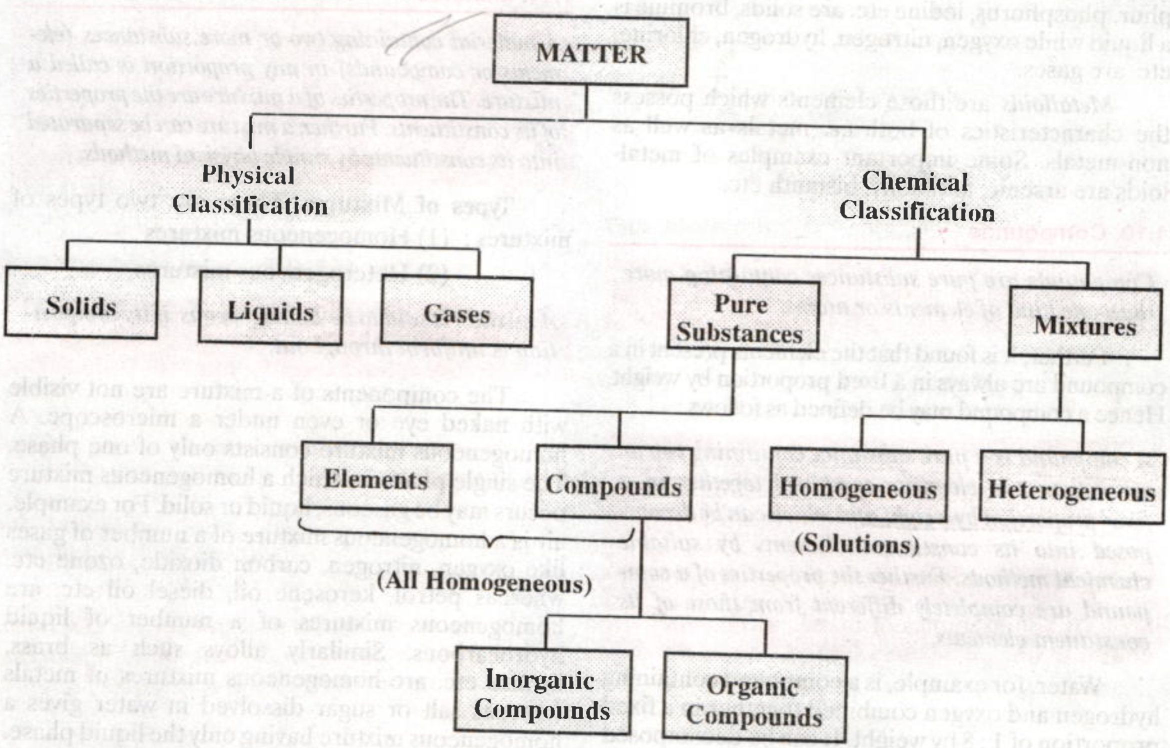
Since any distinct portion of matter that is uniform throughout in composition and properties is called a "phase", hence

A material is said to be homogeneous, if it consists of only one phase. On the other hand, a material is said to be heterogeneous if it consists of a number of phases.

The composition of each phase is uniform throughout within itself but is quite different from those of the other phases. The different phases are separated from each other by distinct boundaries. For example, water, iron, salt, a solution of salt in water, air (which contains a number of gases), motor oil (which contains a number of hydrocarbons) etc. are homogeneous whereas a mixture of ice and water, salt and sand, iron and sulphur etc. are heterogeneous.

However, a more scientific way of classification, called the *'chemical classification of matter'* is briefly described below :—

All types of materials are believed to be made up of "substances". A material containing only one substance is called a "pure substance". On the other hand, materials containing more than one substance are not pure and are called "mixtures". Pure substances are further classified into two types, called "Elements" and "Compounds". Mixtures are also of two types, namely "Homogeneous mixtures" and "Heterogeneous mixtures". Pure substances (*i.e.* elements and compounds) are always homogeneous. Homogeneous mixtures are also called "solutions". The single phase in which a solution occurs may be gaseous, liquid or solid. Air, salt dissolved in water and brass (an alloy of copper and zinc) are examples of gaseous, liquid and solid solutions respectively. Thus the classification of matter may be sketched as shown below.



Now we shall take up a detailed discussion of the three main classes of matter *i.e.* elements, compounds and mixtures.

1.9. Elements

An element is usually defined as the simplest form of a pure substance with definite physical and chemical properties and which can neither be broken into nor built from simpler substances by any chemical or physical method.

This definition is, however, not correct since chemists and physicists together have recently shown that elements can be broken down into and synthesized from simpler substances. With the discovery of atom as a fundamental unit of matter, the definition has now been modified as follows :—

An element is defined as a pure substance that contains only one kind of atoms.

Carbon, sulphur, iron, lead, gold, mercury, oxygen and nitrogen are some examples of familiar elements. The number of elements known to-date is 110. Of these, 92 occur in nature in the earth's crust and the remaining have been prepared artificially in the laboratory through nuclear reactions. The naturally occurring elements are distributed in the crust of the earth in varying proportions.

Types of Elements. Depending upon the physical and chemical properties, the elements are further subdivided into three classes, namely, (1) Metals (2) Non-metals and (3) Metalloids.

Metals are those elements which reflect light and hence possess lustre. They are good conductors of heat and electricity, malleable (*i.e.* hammerred to form sheets) and ductile (*i.e.* can be drawn into wires), exist as solids at room temperature (except mercury) and possess high density. Some examples of common metals are copper, silver, gold, aluminium, iron, lead, tin, nickel, chromium, mercury etc. Majority of the elements (*i.e.* about 65%) are metals.

Non-metals are those elements which do not reflect light and hence do not possess lustre (the only exception being iodine). Further they are brittle, poor conductors of heat and electricity (except graphite) and exist in all the three states, *e.g.* sulphur, phosphorus, iodine etc. are solids, bromine is a liquid while oxygen, nitrogen, hydrogen, chlorine, etc. are gases.

Metalloids are those elements which possess the characteristics of both *i.e.* metals as well as non-metals. Some important examples of metalloids are arsenic, antimony, bismuth etc.

1.10. Compounds

Compounds are pure substances containing more than one kind of elements or atoms.

Further, it is found that the elements present in a compound are always in a fixed proportion by weight. Hence a compound may be defined as follows :

A compound is a pure substance containing two or more than two elements combined together in a fixed proportion by weight and which can be decomposed into its constituent elements by suitable chemical methods. Further the properties of a compound are completely different from those of its constituent elements.

Water, for example, is a compound containing hydrogen and oxygen combined together in a fixed proportion of 1 : 8 by weight. It can be decomposed into its constiuent elements *i.e.* hydrogen and oxygen by passing electricity through water (after acidifying it to make it a good conductor of electricity). Further the properties of water are completely different from its constituents, hydrogen and oxygen. For example, hydrogen is a combustible gas, oxygen gas is a supporter of combustion whereas water is neither combustible nor a supporter of combustion but extinguishes fire.

A few other common examples of compounds are carbon dioxide, sulphur dioxide, sulphuric acid, nitric acid, hydrochloric acid, caustic soda, washing soda, baking soda, common salt, copper sulphate, nitre etc.

Types of Compounds. All the compounds may be divided into the following two categories :

(1) Organic compounds
(2) Inorganic compounds.

Organic compounds are the compounds containing carbon and a few other elements like hydrogen, oxygen, nitrogen, sulphur, halogens etc. These were originally obtained only from animals and plants.

Inorganic compounds are the compounds containing any two or more elements out of more than 110 elements known so far. These are usually obtained from minerals and rocks.

1.11. Mixtures

A material containing two or more substances (elements or compounds) in any proportion is called a mixture. The properties of a mixture are the properties of its constituents. Further, a mixture can be separated into its constituents by simple physical methods.

Types of Mixtures. There are two types of mixtures : (1) Homogeneous mixtures

(2) Heterogeneous mixtures

A mixture is said to be homogeneous if its composition is uniform throughout.

The components of a mixture are not visible with naked eye or even under a microscope. A homogeneous mixture consists only of one phase. The single phase in which a homogeneous mixture occurs may be gaseous, liquid or solid. For example, air is a homogeneous mixture of a number of gases like oxygen, nitrogen, carbon dioxide, ozone etc. whereas petrol, kerosene oil, diesel oil etc. are homogeneous mixtures of a number of liquid hydrocarbons. Similarly alloys such as brass, bronze etc. are homogeneous mixtures of metals whereas salt or sugar dissolved in water gives a homogeneous mixture having only the liquid phase.

All homogeneous mixtures are called *solutions*. Whereas gases can be mixed in any proportions, there is a limit in case of liquid and solid solutions depending upon their mutual solubility.

A mixture is said to be heterogeneous if its composition is not uniform throughout.

A heterogeneous mixture consists of two or more distinct phases. The components of a heterogeneous mixture are visible with naked eye or at least under a microscope. For example, the components of a mixture of iron, sulphur and common salt can be seen with naked eye, lying side by side but the components of milk (which is also a heterogeneous mixture though it looks to be homogeneous), the fats suspended in clear liquid can be seen under a microscope. A few other examples of the heterogeneous mixtures include smoke (a mixture of carbon particles + air), gun powder (a mixture of carbon, sulphur and nitre), oil in water, copper sulphate and sand etc. Strictly speaking, air is also a heterogeneous mixture if dust particles are taken into consideration.

The various characteristics mentioned in the definition of 'mixture' can be explained taking the example of a mixture of iron filings, sulphur and common salt. *Firstly*, they can be mixed in any proportion. *Secondly*, the mixture can be tested to see that it gives the tests of iron, sulphur and common salt. *Thirdly*, the mixture can be separated into its components by simple physical methods *e.g.* iron can be separated by simply moving the magnet into the mixture whereas sulphur and common salt can be separated by making use of their solubility in different solvents *i.e.*, sulphur can dissolve in carbon disulphide whereas common salt is soluble in water. So any one of the two can be dissolved and then filtered.

1.12. Difference between A Compound and A Mixture

The important points of difference between a mixture and compound are summarized below :

MIXTURE	COMPOUND
1. The constituents of a mixture may be present in any ratio.	1. The constituents of a compound are always present in a fixed ratio by mass.
2. Mixture may or may not be homogeneous in nature.	2. Compounds are always homogeneous in nature.
3. The properties of a mixture are midway between those of its constituents.	3. The properties of a compound are entirely different from those of its constituents.
4. The constituents of a mixture can be easily separated by simple mechanical means.	4. The constituents of a compound cannot be easily separated by simple mechanical means. Energy in the form of heat or light is often required.
5. Mixtures are formed as a result of a physical change.	5. Compounds are formed as a result of a chemical change.
6. When a mixture is formed, no heat, light or electrical energy is absorbed or evolved.	6. Formation of a compound is always accompanied by absorption or evolution of heat, light or electrical energy.
7. The melting and boiling points of mixtures are usually not sharp.	7. Chemical compounds possess sharp melting and boiling points.

ADD TO YOUR KNOWLEDGE

1. 20 carat gold is a mixture of 20 parts by weight of gold and 4 parts by weight of copper. Pure gold is 24 carat.
2. Iodized salt used as a table salt is a mixture of NaCl and a small amount of NaI.
3. Cement is a mixture of a number of silicates.
4. The existence of an element in two or more chemically similar but physically different forms is called **allotropy** and the different forms are called **allotropes** *e.g.* diamond, graphite, wood charcoal, lamp black etc. are allotropes of carbon.
5. The existence of a compound in different crystalline forms is called **polymorphism** and the different forms are called **polymorphs** *e.g.* ZnS has two polymorphs called zinc blende and wurtzite.
6. The existence of different compounds with similar chemical composition in the same crystalline form is called **isomorphism** *e.g.* $ZnSO_4.7\,H_2O$ and $FeSO_4.7\,H_2O$ are isomorphs. Similarly, alums, $[M_2SO_4M_2'(SO_4)_3, 24\,H_2O]$, are isomorphs.
7. Substances which absorb moisture from the air are called **hygroscopic substances** *e.g.* anhydrous $CuSO_4$, P_2O_5, quicklime (CaO) etc.
8. Solid substances which absorb a **large amount of moisture** from the air so that they become wet or pass into solution are called **deliquescent** and the phenomenon is called **deliquescence** *e.g.* NaOH, KOH, $MgCl_2$, $CaCl_2$ etc.

ADD TO YOUR KNOWLEDGE CONTD.

9. Some crystalline solids *e.g.* $CuSO_4.5H_2O$, $Na_2CO_3.10H_2O$, $Na_2SO_4.7H_2O$ etc. when exposed to air lose their water partly or wholly. Such substances are called **efflorescent** and this property is called **efflorescence.**

10. The separation of a mixture of gases based on their difference in the rates of diffusion is called **atmolysis** *e.g.* U^{235} and U^{238} are separated by converting them into their hexafluorides which are gaseous and have different rates of diffusion.

SECTION—IV

LAWS OF CHEMICAL COMBINATION

1.13. Introduction

One of the most important aspects of the subject of chemistry is the study of chemical reactions. These chemical reactions take place according to the certain laws, called the 'laws of chemical combination'. These are :

1. Law of Conservation of Mass
2. Law of Constant Composition
3. Law of Multiple Proportions
4. Law of Reciprocal Proportions
5. Law of Combining Volumes (Gay Lussac's Law of Gaseous Volumes).

The first four laws deal with the mass relationships whereas the fifth law deals with the volumes of the reacting gases. Let us now discuss each of these laws one by one.

1.14. Law of Conservation of mass

This law which deals with the masses of the reactants and the products of a chemical reaction (or a physical change) was studied by the great French chemist *Antoine Lavoisier* in 1774. This law may be stated as follows :–

In all physical and chemical changes, the total mass of the reactants is equal to that of the products.

Thus according to this law, there is no increase or decrease in the total mass of matter during a chemical or a physical change. In other words,

Matter can neither be created nor destroyed.

Hence this law is also called the *Law of indestructibility of matter*. The following experiments illustrate the truth of this law.

(a) When matter undergoes a physical change. A piece of ice (solid water) is taken in a small conical flask. It is well corked and weighed. The flask is now heated gently to melt the ice (solid) into water (liquid).

$$\text{Ice} \xrightarrow{\text{Heat}} \text{Water}$$

The flask is again weighed. It is found that there is no change in the weight though a physical change has taken place.

(b) When matter undergoes a chemical change. The following chemical changes illustrate the law :

(*i*) *Precipitation Reaction*. Landolt took solutions of sodium chloride and silver nitrate separately in the two limbs of the Landolt's tube (Fig. 1.2). The limbs were then corked and the tube was weighed. Thereafter, the tube was tilted to allow the two solutions to mix. As a result of the following chemical reaction, a curdy white precipitate of silver chloride is formed :–

$$NaCl\,(aq) + AgNO_3(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$$

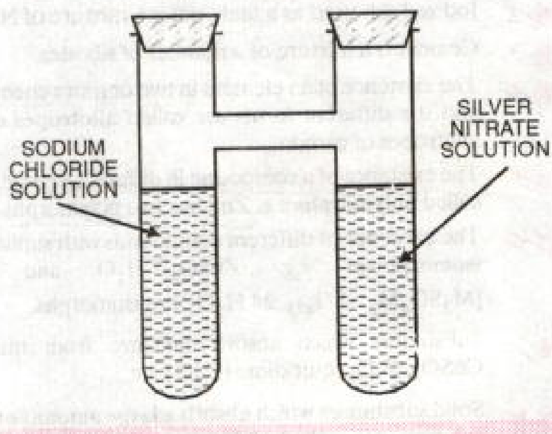


FIGURE 1.2. Landolt's tube.

After the reaction, the tube was again weighed. It was found that weight practically remained unchanged.

(ii) Decomposition of Mercuric oxide. 100 g of mercuric oxide when heated in a closed tube, decomposed to produce 92.6 g of mercury and 7·4 g of oxygen gas :

$$\underset{100\ g}{HgO(s)} \longrightarrow \underset{92\cdot6\ g}{Hg(l)} + \underset{7.4\ g}{\tfrac{1}{2}O_2(g)}$$

Thus, during the above decomposition reaction, matter is neither gained nor lost.

Law of conservation of mass and energy. In a number of reactions, especially the nuclear reactions, the mass of the products is found to be slightly less than the mass of the reactants. The question arises–where is this mass lost ? According to Einstein, this mass is converted into energy according to the equation

$$E = mc^2$$

where m is the mass lost and c is the velocity of light. Hence the law of conservation of mass has been modified. Now it is known as law of conservation of mass and energy. It states as follows : ***The mass and energy are interconvertible but the total sum of the mass and energy during any physical or chemical change remains constant.***

PROBLEMS ON LAW OF CONSERVATION OF MASS

EXAMPLE. *4.90 g of $KClO_3$ when heated produced 1·92 g of oxygen and the residue (KCl) left behind weighs 2·96 g. Show that these results illustrate the law of conservation of mass.*

Solution. Mass of $KClO_3$ taken = 4·90 g

Total mass of the products ($KCl + O_2$)

= 2.96 + 1.92 = 4.88 g

Difference between the mass of the reactant and the total mass of the products

= 4.90 – 4·88 = 0·02 g.

This small difference may be due to experimental error.

Thus law of conservation of mass holds good within experimental errors.

PROBLEMS FOR PRACTICE

1. What mass of silver nitrate will react with 5.85 g of sodium chloride to produce 14.35 g of silver chloride and 8.5 g of sodium nitrate, if the law of conservation of mass is true ? **[Ans. 17.0 g]**
2. When 4.2 g of $NaHCO_3$ is added to a solution of acetic acid (CH_3COOH) weighing 10.0 g, it is observed that 2.2 g of CO_2 is released into the atmosphere. The residue left behind is found to weigh 12.0 g. Show that these observations are in agreement with the law of conservation of mass. *(N.C.E.R.T.)*
3. If 6.3 g of $NaHCO_3$ are added to 15.0 g of CH_3COOH solution, the residue is found to weigh 18.0 g. What is the mass of CO_2 released in the reaction ? *(N.C.E.R.T.)* **[Ans. 3.3 g]**

HINTS FOR DIFFICULT PROBLEMS

1. $$\underset{x g}{AgNO_3} + \underset{5\cdot85\ g}{NaCl} \longrightarrow \underset{14\cdot35\ g}{AgCl} + \underset{8\cdot5\ g}{NaNO_3}$$

$x + 5\cdot85 = 14\cdot35 + 8\cdot5$

or $x = 17\cdot0\ g$

2. $$\underset{4\cdot2\ g}{NaHCO_3} + \underset{10\cdot0\ g}{CH_3COOH} \longrightarrow \underbrace{CH_3COONa + H_2O}_{Residue\ =\ 12\cdot0\ g} + \underset{2\cdot2\ g}{CO_2\uparrow}$$

Total mass of reactants = 4·2 + 10·0 g = 14·2 g

Total mass of products = 12·0 + 2·2 g = 14·2 g

1.15. Law of Constant Composition or Definite Proportions

This law was discovered by a French chemist *J.L. Proust* in 1799 and deals with the composition of elements present in a given compound. It states that—

A chemical compound is always found to be made up of the same elements combined together in the same fixed proportion by mass.

For example, pure water obtained from *whatever* source (well, river, lake or sea) or any country (India, Russia, America etc.) will always be made up of only hydrogen and oxygen elements combined together in the same fixed ratio of 1 : 8 by mass.

Further, a sample of CO_2 may be prepared in the laboratory by (*a*) heating lime stone ($CaCO_3$), (*b*) by burning coal in air, (*c*) by the action of dilute hydrochloric acid on marble, (*d*) by heating sodium bicarbonate. In each case, it is found that CO_2 is made up of the same elements *i.e.*, carbon and oxygen, combined together in the same fixed ratio of 12 : 32 or 3 : 8 by mass.

(*a*) $CaCO_3 \xrightarrow{\text{Heat}} CaO + CO_2 \uparrow$

(*b*) $C + O_2 \longrightarrow CO_2 \uparrow$

(*c*) $CaCO_3 + 2HCl \longrightarrow CaCl_2 + CO_2 \uparrow + H_2O$

(*d*) $2NaHCO_3 \xrightarrow{\text{Heat}} Na_2CO_3 + CO_2 \uparrow + H_2O$

Limitations of Law of Constant Composition

(1) *The law is not applicable if an element exists in different isotopes which may be involved in the formation of the compound.* For example, in the formation of the compound CO_2, if C–12 isotope combines, the ratio of C : O is 12 : 32. But if C–14 isotope combines, the ratio of C : O is 14 : 32.

(2) *The elements may combine in the same ratio but the compounds formed may be different.* For example, in the compounds, C_2H_5OH and CH_3OCH_3 (both having same molecular formula viz. C_2H_6O) the ratio of C : H : O = 24 : 6 : 16 = 12 : 3 : 8 by mass.

PROBLEMS ON LAW OF CONSTANT COMPOSITION

EXAMPLE 1. *6·488 g of lead combine directly with 1·002 g of oxygen to form lead peroxide (PbO_2). Lead peroxide is also produced by heating lead nitrate and it was found that the percentage of oxygen present in lead peroxide is 13·38 percent. Use these data to illustrate the law of constant composition.*

Solution. Step 1. *To calculate the percentage of oxygen in first experiment.*

Mass of peroxide formed

$= 6 \cdot 488 + 1 \cdot 002 = 7 \cdot 490$ g.

7·490 g of lead peroxide contain 1·002 g of oxygen

∴ 100 g of lead peroxide will contain oxygen

$$= \frac{1 \cdot 002}{7 \cdot 490} \times 100 = 13 \cdot 38 \text{ g}$$

i.e. oxygen present = 13·38%

Step 2. *To compare the percentage of oxygen in both the experiments.*

Percentage of oxygen in PbO_2 in the first experiment = 13·38

Percentage of oxygen in PbO_2 in the second experiment = 13·38

Since the percentage composition of oxygen in both the samples of PbO_2 is identical, the above data illustrate the law of constant composition.

EXAMPLE 2. *Copper oxide was prepared by the following methods :*

(a) In one case, 1·75 g of the metal were dissolved in nitric acid and igniting the residual copper nitrate yielded 2·19 g of copper oxide.

(b) In the second case, 1·14 g of metal dissolved in nitric acid were precipitated as copper hydroxide by adding caustic alkali solution. The precipitated copper hydroxide after washing, drying and heating yielded 1·43 g of copper oxide.

(c) In the third case, 1·45 g of copper when strongly heated in a current of air yielded 1·83 g of copper oxide.

Show that the given data illustrate the law of constant composition.

Solution. Step 1. *In the first experiment.*

2·19 g of copper oxide contained 1·75 g of Cu.

∴ 100 g of copper oxide contained

$$= \frac{1 \cdot 75}{2 \cdot 19} \times 100 = 79 \cdot 91 \text{ g}$$

Step 2. *In the second experiment.*

$1 \cdot 43\ g$ of copper oxide contained $1 \cdot 14\ g$ of copper

$\therefore$ 100 *g* of copper oxide contained

$$= \frac{1 \cdot 14}{1 \cdot 43} \times 100 = 79 \cdot 72\ g.$$

Step 3. *In the third experiment.*

$1 \cdot 83\ g$ of copper oxide contained $1 \cdot 46\ g$ of copper

$\therefore$ 100 *g* of copper oxide contained

$$= \frac{1 \cdot 46}{1 \cdot 83} \times 100 = 79 \cdot 78\ g.$$

Thus the percentage of copper in copper oxide derived from all the three experiments is nearly the same. Hence, the above data illustrate the law of constant composition.

PROBLEMS FOR PRACTICE

1. $2 \cdot 16$ g of copper metal when treated with nitric acid followed by ignition of the nitrate gave $2 \cdot 70$ g of copper oxide. In another experiment $1 \cdot 15$ g of copper oxide upon reduction with hydrogen gave $0 \cdot 92\ g$ of copper. Show that the above data illustrate the Law of Definite Proportions.
2. Silver chloride is prepared by

   (*i*) dissolving $0 \cdot 5$ g of silver wire in nitric acid and adding excess of hydrochloric acid to silver nitrate formed. The silver chloride precipitated is separated, washed and dried. The weight of silver chloride is 0.66 g.

   (*ii*) heating 1 g of silver metal in a current of dry chlorine gas till the metal is completely converted into its chloride. It is found to weigh 1.32 *g*. Illustrate the law of constant composition by the above data.

HINTS FOR DIFFICULT PROBLEMS

1. % of Cu in copper oxide in 1st case

$$= \frac{2 \cdot 16}{2 \cdot 70} \times 100 = 80\%$$

% of oxygen = 20%

% of Cu in copper oxide in 2nd case

$$= \frac{0 \cdot 92}{1 \cdot 15} \times 100 = 80\%$$

% of oxygen = 20%

2. % of Ag in AgCl in 1st case

$$= \frac{0 \cdot 5}{0 \cdot 66} \times 100 = 75 \cdot 76\%$$

% of Cl = 24·24%

% of Ag in AgCl in 2nd case

$$= \frac{1}{1 \cdot 32} \times 100 = 75 \cdot 76\%$$

% of Cl = 24·24%

1.16. Law of Multiple Proportions

This law which was first studied by Dalton in 1804 may be defined as follows :–

When two elements combine to form two or more chemical compounds, then the masses of one of the elements which combine with a fixed mass of the other, bear a simple ratio to one another.

The law of multiple proportions is illustrated by the following examples :

(1) Compounds of Carbon and Oxygen. The element carbon combines with oxygen to form two compounds, namely, *carbon dioxide* and *carbon monoxide*. In *carbon dioxide*, 12 parts by mass of carbon combine with 32 parts by mass of oxygen while in *carbon monoxide*, 12 parts by mass of carbon combine with 16 parts by mass of oxygen. Therefore, the masses of oxygen which combine with a fixed mass of carbon (12 parts) in carbon monoxide and carbon dioxide are 16 and 32 respectively. These masses of oxygen bear a simple ratio of 16 : 32 or 1 : 2 to each other.

(2) Compounds of Sulphur and Oxygen. Like carbon, the element sulphur also forms two oxides *viz. sulphur dioxide* and *sulphur trioxide*. In *sulphur dioxide*, 32 parts by mass of sulphur combine with 32 parts by mass of oxygen but in case of *sulphur trioxide* 32 parts by mass of sulphur combine with 48 parts by mass of oxygen. Therefore, the masses of oxygen which combine with a fixed mass of sulphur (32 parts) in the two oxides are 32 and 48 respectively. These bear a simple ratio of 32 : 48 or 2 : 3 to each other.

(3) Compounds of Nitrogen and Oxygen. The elements nitrogen and oxygen combine to produce as many as 5 oxides of nitrogen, *viz.*, *Nitrous oxide, Nitric oxide, Nitrogen trioxide, Nitrogen tetroxide or peroxide and Nitrogen pentoxide.*

The masses of nitrogen and oxygen which combine with one another are :–

Compound	**Nitrous oxide (N_2O)**	**Nitric oxide (NO)**	**Nitrogen trioxide (N_2O_3)**	**Nitrogen tetraoxide (N_2O_4)**	**Nitrogen pentoxide (N_2O_5)**
Nitrogen	28	14	28	28	28 (parts by mass)
Oxygen	16	16	48	64	80 (parts by mass)

Fixing the mass of nitrogen as 14 parts, the masses of oxygen in these five oxides are 8, 16, 24, 32 and 48 parts respectively. These masses bear a simple ratio of 1 : 2 : 3 : 4 : 5 to one another.

PROBLEMS ON LAW OF MULTIPLE PROPORTIONS

Imp

• EXAMPLE 1. *Carbon is found to form two oxides, which contain 42·9% and 27·3% of carbon respectively. Show that these figures illustrate the law of multiple proportions.* (N.C.E.R.T.)

Solution.

Step 1. *To calculate the percentage composition of carbon and oxygen in each of the two oxides*

	First oxide	Second oxide	
Carbon	42·9 %	27·3 %	(Given)
Oxygen (by difference)	57·1 %	72·7 %	

Step 2. *To calculate the masses of carbon which combine with a fixed mass i.e., one part by mass of oxygen in each of the two oxides.*

In the first oxide, 57·1 parts by mass of oxygen combine with carbon = 42.9 parts.

∴ 1 part by mass of oxygen will combine with

carbon $= \frac{42 \cdot 9}{57 \cdot 1} = 0 \cdot 751.$

In the second oxide. 72.7 parts by mass of oxygen combine with carbon = 27.3 parts.

∴ 1 part by mass of oxygen will combine with

carbon $= \frac{27.3}{72.7} = 0.376$

Step 3. *To compare the masses of carbon which combine with the same mass of oxygen in both the oxides.*

The ratio of the masses of carbon that combine with the same mass of oxygen (1 part) is

0.751 : 0.376 or 2 : 1

Since this is a simple whole number ratio, so the above data illustrate the law of multiple proportions.

EXAMPLE 2. *Two oxides of a certain metal were separately heated in a current of hydrogen until constant weights were obtained. The water produced in each case was carefully collected and weighed. 2 grams of each oxide gave respectively 0·2517 grams and 0·4526 grams of water. Show that these results establish the Law of Multiple Proportions.*

Solution.

Step 1. *To calculate the mass of oxygen in each oxide.*

Here, we are given –

Mass of each oxide = 2.0 g

Mass of water produced in case I = 0·2517 g

Mass of water produced in case II = 0·4526 g

18 g of H_2O ≡ 16 g of oxygen

i.e. 18 g of water contain oxygen = 16 g

∴ 0·2517 g of water contains oxygen

$= \frac{16}{18} \times 0 \cdot 2517 \text{ g} = 0 \cdot 2237 \text{ g}$

and 0·4526 g of water contains oxygen

$= \frac{16}{18} \times 0 \cdot 4526 \text{ g} = 0 \cdot 4023 \text{ g}$

Step 2. *To calculate the mass of oxygen which would combine with 1 g of metal in each oxide.*

In case I. Mass of metal oxide = 2 g

Mass of oxygen = 0·2237 g

∴ Mass of metal = 2 – 0·2237 = 1·7763 g

∴ Mass of oxygen which combines with 1·7763 g of metal = 0·2237 g

∴ Mass of oxygen which combines with 1 g of metal $= \frac{0 \cdot 2237}{1 \cdot 7763} \text{ g} = 0 \cdot 1259 \text{ g}$

In case II. Mass of metal oxide = 2 g

Mass of oxygen = 0·4023 g

$\therefore$ Mass of metal $= 2 - 0 \cdot 4023 = 1 \cdot 5977$ g

Mass of oxygen which combines with 1.5977 g of metal = 0.4023 g

$\therefore$ Mass of oxygen which combines with 1 g of metal $= \frac{0.4023}{1.5977}$ g $= 0.2515$ g

Step 3. *To compare the masses of oxygen which combine with the same mass of metal in the two oxides.*

The masses of oxygen which combine with 1 g of metal in the two oxides are respectively 0.1259 g and 0.2515 g. These masses are in the ratio

0.1259 : 0.2515 or 1 : 2

Since this is a simple ratio, so the above results establish the *Law of Multiple Proportions*.

Imp

EXAMPLE 3. *Two oxides of a metal contain 27.6% and 30.0% of oxygen respectively. If the formula of the first oxide is M_3O_4, find that of the second.*

Solution. In the first oxide, oxygen = 27.6, metal $= 100 - 27.6 = 72.4$ parts by mass.

As the formula of the oxide is M_3O_4, this means 72.4 parts by mass of metal = 3 atoms of metal and 4 atoms of oxygen = 27.6 parts by mass.

In the second oxide, oxygen $= 30 \cdot 0$ parts by mass and metal $= 100 - 30 = 70$ parts by mass.

But $72 \cdot 4$ parts by mass of metal = 3 atoms of metal

$\therefore$ 70 parts by mass of metal

$= \frac{3}{72 \cdot 4} \times 70$ atoms of metal

$= 2 \cdot 90$ atoms of metal

Also, $27 \cdot 6$ parts by mass of oxygen = 4 atoms of oxygen

$\therefore$ 30 parts by mass of oxygen

$= \frac{4}{27 \cdot 6} \times 30$ atoms of oxygen

$= 4 \cdot 35$ atoms of oxygen

Hence, ratio of M : O in the second oxide

$= 2 \cdot 90 : 4 \cdot 35 = 1 : 1 \cdot 5 = 2 : 3$

$\therefore$ Formula of the metal oxide is M_2O_3.

PROBLEMS FOR PRACTICE

1. Two oxides of lead were separately reduced to metallic lead by heating in a current of hydrogen and the following data obtained :

   (*i*) Mass of yellow oxide taken = 3.45 g

   Loss in mass during reduction = 0.24 g

   (*ii*) Mass of brown oxide taken = 1.227 g

   Loss in mass during reduction = 0.16 g

   Show that the above data illustrate the Law of Multiple Proportions.
2. Copper gives two oxides. On heating 1.0 g of each in hydrogen gas, 0.888 g and 0.799 g of the metal are produced. Show that the results agree with the Law of Multiple Proportions.
3. Two oxides of nitrogen contain the following percentage compositions :

   (*i*) Oxide A contains 63.64 % nitrogen and 36.36 % oxygen.

   (*ii*) Oxide B contains 46.67 % nitrogen and 53.33 % oxygen.

   Establish the Law of Multiple Proportions.
4. A metal forms two oxides. One contains 46.67% of the metal and another, 63.94% of the metal. Show that these results are in accordance with the law of multiple proportions.
5. Nitrogen forms five compounds with oxygen in which 1.0 g of nitrogen combines with 0.572, 1.14, 1.73, 2.28 and 2.85 g of oxygen respectively. Show that these figures agree with law of multiple proportions.
6. Elements X and Y form two different compounds. In the first, 0.324 g of X is combined with 0.471 g of Y. In the second, 0.117 g of X is combined with 0.509 g of Y. Show that these data illustrate the Law of Multiple Proportions.
7. If a certain oxide of nitrogen weighing $0 \cdot 11$ g gives 56 ml of nitrogen and another oxide of nitrogen weighing $0 \cdot 15$ g gives the same volume of nitrogen (both at STP), show that these results support the law of multiple proportions.
8. (*i*) 10 g of lead on heating gave $10 \cdot 78$ g of litharge, PbO.

   (*ii*) $9 \cdot 775$ g of red lead (Pb_3O_4) yielded on strong heating $9 \cdot 545$ g of litharge.

   (*iii*) $4 \cdot 87$ g of lead peroxide (PbO_2) gave on heating $4 \cdot 545$ g of litharge.

   Show that these results illustrate the law of multiple proportions.

HINTS FOR DIFFICULT PROBLEMS

1. Calculate the mass of lead combined with 1 g of oxygen in each case which are 13.375 g and 6.669 g *i.e.* in the ratio of 2 : 1.
2. Calculate the masses of copper combined with 1.0 g of oxygen in each case which are 7.93 g and 3.97 g *i.e.* in the ratio 2 : 1.
3. Calculate the mass of oxygen combined with 1.0 g of nitrogen in each case which are 0.571 g and 1.143 g, *i.e.* in the ratio 1 : 2.
4. Masses of metal that combine with 1.0 g of oxygen are 0.875 g and 1.773 g, the ratio is 1:2.
5. Masses of oxygen which combine with 1 g of N are in the ratio of 1 : 2 : 3 : 4 : 5.
6. Calculate the masses of Y which combine with 1 g of X in each case which are 1.454 g and 435 g *i.e.* in the ratio 1 : 3.
7. 22400 cc of N_2 at STP weigh = 28 g. Calculate the mass of nitrogen. Then subtract from the mass of oxide of nitrogen to calculate the mass of oxygen.
8. Using (*i*), calculate the mass of lead present in litharge in (*ii*) and (*iii*). Then calculate the masses of oxygen in all the three cases which combine with 1 g of lead. The ratio comes out to be 3 : 4 : 6.

1.17. Law of Reciprocal proportions

This law was put forward by Richter in 1792. It states as follows :–

The ratio of the masses of two elements A and B which combine separately with a fixed mass of the third element C is either the same or some simple multiple of the ratio of the masses in which A and B combine directly with each other.

This law may be illustrated with the help of the following examples :

(1) *The elements C and O combine separately with the third element H to form CH_4 and H_2O and they combine directly with each other to form CO_2, as shown in Fig. 1.3.*

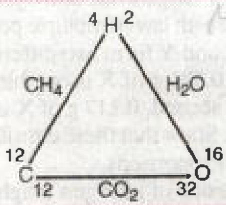


FIGURE 1.3.

In CH_4, 12 parts by mass of carbon combine with 4 parts by mass of hydrogen. In H_2O, 2 parts by mass of hydrogen combine with 16 parts by masses of oxygen. Thus the weight of C and O which combine with fixed mass of hydrogen (say 4 parts by mass) are 12 and 32 *i.e.* they are in the ratio 12 : 32 or 3 : 8.

Now in CO_2, 12 parts by mass of carbon combine directly with 32 parts by mass of oxygen i.e. they combine directly in the ratio 12 : 32 or 3 : 8 which is the same as the first ratio.

(2) *The elements H and O combine separately with the third element S to form H_2S and SO_2 and they combine directly with each other to form H_2O as shown in Fig. 1.4.*

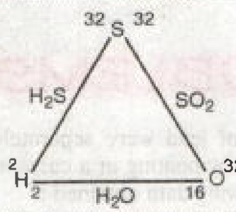


FIGURE 1.4.

As shown in the Fig., the masses of H and O which combine with the fixed mass of S viz 32 parts are 2 and 32 i.e. they are in the ratio 2 : 32 or 1 : 16.

When H and O combine directly to form H_2O, the ratio of their combining masses is 2 : 16 or 1 : 8.

The two ratios are related to each other as

$$\frac{1}{16} : \frac{1}{8} = 1 : 2$$

i.e. they are simple multiple of each other.

EXAMPLE. *Ammonia contains 82·35% of nitrogen and 17·65% of hydrogen. Water contains 88·90% of oxygen and 11·10 % of hydrogen. Nitrogen trioxide contains 63·15 % of oxygen and 36·85% of nitrogen. Show that these data illustrate the law of reciprocal proportions.*

Solution. In NH_3, 17·65 g of H combine with

$N = 82 \cdot 35$ g

$\therefore$ 1 g of H combine with $N = \frac{82 \cdot 35}{17 \cdot 65}$ g

$= 4 \cdot 67$ g

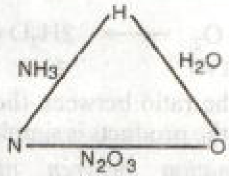


FIGURE 1.5.

In H_2O, 11·10 g of H combine with O

$= 88 \cdot 90$ g

$\therefore$ 1 g of H combine with $O = \frac{88 \cdot 90}{11 \cdot 10}$ g

$= 8 \cdot 01$ g

$\therefore$ Ratio of the masses of N and O which combine with fixed mass (= 1 g) of H

$= 4 \cdot 67 : 8 \cdot 01 = 1 : 1 \cdot 72$

In N_2O_3, ratio of masses of N and O which combine with each other $= 36 \cdot 85 : 63 \cdot 15$

$= 1 : 1 \cdot 71$

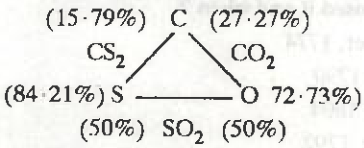
Thus the two ratios are the same. Hence it illustrates the law of reciprocal proportions.

PROBLEMS FOR PRACTICE

1. Carbon dioxide contains 27·27 % of carbon, carbon disulphide contains 15·79 % of carbon and sulphur dioxide contains 50% of sulphur. Are these figures in agreement with the law of reciprocal proportions ? **[Ans. Yes]**
2. Phosphorus trichloride contains 22·57 % of phosphorus, phosphine (PH_3) contains 91·18 % of phosphorus while hydrogen chloride gas contains 97·23 % of chlorine. Prove by calculations, which law is illustrated by these data.
3. 61·8 g of A combine with 80 g of B. 30.9 g of A combine with 106·5 g of C. B and C combine to form compound CB_2. Atomic weights of C and B are respectively 35·5 and 6·6. Show that the law of reciprocal proportions is obeyed.

HINTS FOR DIFFICULT PROBLEMS

1.



1 g C will combine with $S = \frac{84 \cdot 21}{15 \cdot 79} = 5 \cdot 33\,g$

1 g C will combine with $O = \frac{72 \cdot 73}{27 \cdot 27}$

$= 2 \cdot 67$ g

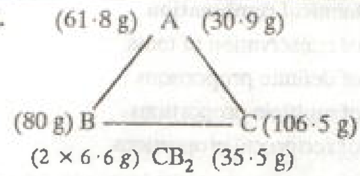
$\therefore$ Ratio of masses of S and O which combine with fixed mass of carbon (viz 1 g)

$= 5 \cdot 33 : 2 \cdot 67$

$= 2 : 1.$

Ratio of masses of S and O which combine directly with each other = 50 : 50 = 1 : 1. Thus the two ratios are simple multiple of each other.

3.



Mass of B combining with 1 g of $A = \frac{80}{61 \cdot 8} = 1 \cdot 29\,g$

Mass of C combining with 1 g of $A = \frac{106 \cdot 5}{30 \cdot 9} = 3 \cdot 45\,g$

Ratio of masses of B and C combining with fixed mass of A = 1·29 : 3·45 = 1 : 2·67 = 3 : 8

Ratio of masses of B and C combining directly with each other

$= 13 \cdot 2 : 35 \cdot 5 = 1 : 2 \cdot 67 = 3 : 8$

Thus the two ratios are same.

1.18. Gay Lussac's Law of Gaseous Volumes

Gay Lussac investigated a large number of chemical reactions occurring in gases. As a result of his experiments, Gay Lussac found that there exists a definite relationship among the volumes of the gaseous reactants and the products. In 1808, he put forward a generalization known as the *Gay Lussac's Law of Combining Volumes*. This may be stated as follows :–

When gases react together, they always do so in volumes which bear a simple ratio to one another and to the volumes of the products, if these are also gases, provided all measurements of volumes are done under similar conditions of temperature and pressure.

Consider, for illustration, the following examples :

(*a*) *Combination between hydrogen and chlorine*. One volume of hydrogen and one volume of chlorine always combine to form two volumes of hydrochloric acid gas.

$$\underset{1\text{ vol.}}{H_2} + \underset{1\text{ vol.}}{Cl_2} \longrightarrow \underset{2\text{ vol.}}{2HCl}$$

The ratio between the volumes of the reactants and the product in this reaction is simple, i.e., 1 : 1 : 2. Hence it illustrates the *Law of Combining Volumes*.

(*b*) *Combination between hydrogen and oxygen*. Two volumes of hydrogen always combine with one volume of oxygen to produce two volumes of steam.

$$\underset{2\text{ vol.}}{2H_2} + \underset{1\text{ vol.}}{O_2} \longrightarrow \underset{2\text{ vol.}}{2H_2O}\text{ (steam)}$$

Here also, the ratio between the volumes of the reactants and the products is simple *i.e.*, 2 : 1 : 2.

(*c*) *Combination between nitrogen and hydrogen*. One volume of nitrogen always combines with three volumes of hydrogen to form two volumes of ammonia.

$$\underset{1\text{ vol.}}{N_2} + \underset{3\text{ vol.}}{3H_2} \longrightarrow \underset{2\text{ vol.}}{2NH_3}$$

This reaction also indicates a simple ratio of 1 : 3 : 2 between the volumes of the reactants and the products.

It may be emphasized here that the volume used in the above reactions are expressed in similar units *i.e., ml or litres*.

ADD TO YOUR KNOWLEDGE

Law of chemical combination	Who proposed it and when ?
(*a*) Law of conservation of mass	Lavoisier, 1774
(*b*) Law of definite proportions	Proust, 1799
(*c*) Law of multiple proportions	Dalton, 1804
(*d*) Law of reciprocal proportions	Richter, 1792

SECTION—V
ATOMS AND MOLECULES— DALTON'S ATOMIC THEORY AND AVOGADRO'S HYPOTHESIS

1.19. Introduction

In Section III, we discussed the classification of matter into Elements, Compounds and Mixtures. However, a number of questions arise, *e.g.*

(*i*) Why is one element different from another ?

(*ii*) Why do elements combine to form compounds ?

(*iii*) Why is a compound different from a mixture ?

To answer the above questions, it is essential to look into the *'structure of matter' i.e.* what are the ultimate building blocks of matter. The concept that matter was made up of small indivisible particles was put forward by Greek philosphers (notably *Democritus*) and these ultimate particles were called *atomos* which means *indivisible*. However, a real good mental picture of matter which could answer the above questions as well as could explain the laws of chemical combination was put forward by John Dalton, an English School teacher in 1808. This is known as 'Dalton's atomic theory.'

1.20. Dalton's Atomic Theory

To describe the structure of matter which could explain the experimental facts known at that time about elements, compounds and mixtures and also the laws of chemical combination, John Dalton in 1808 put forward a theory known as Dalton's atomic theory. The main points of this theory are as follows :--

1. *Matter is made up of extremely small indivisible particles called atoms.*

2. *Atoms of the same element are identical in all respects i.e. size, shape and mass.*

3. *Atoms of different elements have different masses, sizes and also possess different chemical properties.*

4. *Atoms of the same or different elements combine together to form compound atoms (now called as molecules).*

5. *When atoms combine with one another to form compound atoms (molecules), they do so in simple whole number ratios, such as 1 : 1, 2 : 1, 2 : 3 and so on.*

6. *Atoms of two elements may combine in different ratios to form more than one compound. For example, sulphur combines with oxygen to form sulphur dioxide and sulphur trioxide, the combining ratios being 1 : 2 and 1 : 3 respectively.*

7. ***An atom is the smallest particle that takes part in a chemical reaction.*** In other words, whole atoms, rather than fractions of atoms take part in a chemical reaction.

8. *An atom can neither be created nor destroyed.*

Explanation of the Laws of Chemical Combination by Dalton's Atomic Theory.

1. *Law of Conservation of Mass.* Matter is made up of atoms (postulate 1) which can neither be created nor destroyed (postulate 8). Hence matter can neither be created nor destroyed.

2. *Law of Constant Composition.* It follows directly from postulate 5.

3. *Law of Multiple Proportions.* It follows directly from postulate 6.

4. *Law of Reciprocal Proportions.* As atoms combine with each other in simple ratio (postulate 5), therefore all the ratios involved are simple which may be same or some simple multiple of each other.

1.21. Limitations of Dalton's Atomic Theory

Dalton's atomic theory was the first milestone towards the inner structure of matter. It gave a powerful initative to the scientists about the study of matter during the 19th century. It held the ground for about a century. But the brilliant researches conducted in the beginning of 20th century by Sir J.J. Thomson, Lord Rutherford, Neils Bohr and others have revolutionised our knowledge about the structure of atom. The main drawbacks of Dalton's Atomic Theory are :

(*i*) *It could explain the laws of chemical combination by mass but failed to explain the law of gaseous volumes.*

(*ii*) *It could not explain why atoms of different elements have different masses, sizes, valencies etc.*

(*iii*) *Why do atoms of the same or different elements combine at all to form molecules ?*

(*iv*) *What is the nature of binding force between atoms and molecules which accounts for the existence of matter in three states i.e., solids, liquids and gases ?*

(*v*) *It makes no distinction between the ultimate particles of an element or a compound.*

1.22. Modified Dalton's Atomic Theory (Modern Atomic Theory)

As a result of the researches made by various chemists and physicist, Dalton's atomic theory has undergone radical changes. Nevertheless salient points of this theory have been retained since they satisfactorily explain the laws of chemical combination. The main points of the modern theory are :

(*i*) Atom is no longer considered to be indivisible. It has been found that an atom has a complex structure. It is made up of a number of small particles of which the important ones are *electrons, protons and neutrons*.

(*ii*) Atoms of the same element may have different atomic masses. For example, atoms of hydrogen may have atomic masses of 1 amu, 2 amu or 3 amu. Similarly, atoms of chlorine may have atomic masses of 35 amu and 37 amu. *Such atoms of the same element which possess different atomic masses are called* **isotopes**. Thus, atoms of the same element may not be identical in all respects.

(*iii*) Atoms of different elements may have same atomic masses. For example atoms of calcium and argon have the same atomic mass *i.e.*, 40 amu. *Such atoms of the different elements which have the same atomic masses are called* **isobars**. Thus atoms of the different elements may be identical in one or more respects.

***(iv)* The ratio in which the different atoms combine with one another may be fixed and integral but may not always be simple**. For example, the ratio in which the elements C, H and O combine to form a molecule of cane sugar ($C_{12}H_{22}O_{11}$) is 12:22 : 11 *which by no means is simple. However, this ratio is fixed and is also integral.*

***(v)* Atom is the smallest particle that takes part in a chemical reaction**. Though an atom is made up of smaller particles such as electrons, protons and neutrons yet *atom is the smallest particle that takes part in a chemical reaction.*

***(vi)* Atom is no longer indestructible.** By carrying out nuclear reactions, atoms of an element may be changed into another. For example, atoms of nitrogen can be changed into oxygen by bombardment with α-rays. Similarly, uranium ($_{92}U^{235}$) can be converted into plutonium ($_{94}Pu^{239}$) through interaction with neutrons. *The procress of interconversion of elements through changes in atomic nuclei is called* **transmutation.**

Further during these nuclear reactions, a small amount of mass is always converted into energy in accordance with Einstein equation, $E = mc^2$. *Thus, atoms can be interconverted and mass can be changed into energy. In other words, atoms are no longer indestructible.* However, in ordinary *laboratory reactions, atoms remain unaffected, i.e. they can neither be created nor destroyed.*

1.23. Avogadro's Hypothesis/ Law/Principle

According to Gay-Lussac's Law of gaseous volumes, *gases always combine with one another in a simple ratio by volume.* But according to Dalton's Atomic Theory, *elements combine with one another in a simple whole number atomic ratio to form compounds.* Berzelius, a Swedish Chemist, tried to correlate Dalton's Atomic Theory and Gay-Lussac's Law of gaseous volumes. He argued that while *elements combine in a simple ratio by atoms, gases combine in a simple ratio by volume,* there must be some relationship between the volume of a gas and the number of atoms it contains. This led Berzelius to put forward his hypothesis called **Berzelius Hypothesis**. It may be stated as –

Equal volumes of all gases under similar conditions of temperature and pressure contain equal number of atoms.

Let us apply this hypothesis to the gaseous reaction between hydrogen and chlorine to produce hydrogen chloride gas. Experimentally, it has been found that one volume of hydrogen combines with one volume of chlorine to produce two volumes of hydrogen chloride gas.

Hydrogen +	Chlorine $\longrightarrow$	Hydrogen chloride gas
1 vol.	1 vol.	2 vol.
n atoms	*n* atoms	2*n* compound atoms
		(Applying Berzelius Hypothesis)
$\frac{1}{2}$ atom	$\frac{1}{2}$ atom	1 compound atom

This implies that one compound atom of hydrogen chloride gas is made up of $\frac{1}{2}$ atom of hydrogen and $\frac{1}{2}$ atom of chlorine. This is in direct conflict with Dalton's atomic theory which states that atoms are the ultimate particles of elements and are indivisible. This hypothesis was, therefore, rejected.

Avogadro, an Italian scientist, solved this problem which cropped up as a result of Berzelius hypothesis by clearly distinguishing between the two ultimate particles of matter, *i.e.*, an atom and a molecule. According to him,

An atom is the smallest particle of an element which can take part in a chemical reaction. It may or may not be capable of independent existence.

A molecule is the smallest particle of an element or a compound which is capable of independent existence.

Since the smallest particle of a gas which can exist independently is the molecule and not the atom so the volume of a gas must be related to the number of molecules (rather than atoms) present in it. He thus put forwar d his hypothesis known as **Avogadro's hypothesis**. This states that

Equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules.

This hypothesis has been found to explain elegantly all the gaseous reactions and is now widely recognized as a law or a principle known as **Avogadro's Law or Avogadro's principle.**

The above reaction between hydrogen and chlorine can be explained on the basis of Avogadro's Law as follows :

$$\underset{1 \text{ vol.}}{\text{Hydrogen}} + \underset{1 \text{ vol.}}{\text{Chlorine}} \longrightarrow \underset{2 \text{ vol.}}{\text{Hydrogen chloride gas}}$$

(By experiment)

n molecules $\quad n$ molecules $\quad 2n$ molecules

(By Avogadro's Law)

$\frac{1}{2}$ molecule $\quad \frac{1}{2}$ molecule $\quad$ 1 molecule

(By dividing throughout by $2n$)

It implies that one molecule of hydrogen chloride gas is made up of $\frac{1}{2}$ molecule of hydrogen and $\frac{1}{2}$ molecule of chlorine. Since a molecule is made up of two or more atoms so $\frac{1}{2}$ molecule is possible and may contain one or two atoms. Thus, this result does not contradict Dalton's atomic theory.

Applications of Avogadro's Law.

(1) In the calculation of Atomicity of Elementary Gases.

Atomicity of an elementary substance is defined as the number of atoms of the element present in one molecule of the substance, e.g. atomicity of oxygen (O_2) is two while that of ozone (O_3) is three.

Taking the example of oxygen, its atomicity can be calculated as follows :

2 volumes of hydrogen combine with 1 volume of oxygen to form two volumes of water vapours

$$\underset{2 \text{ volumes}}{\text{Hydrogen}} + \underset{1 \text{ volume}}{\text{Oxygen}} \longrightarrow \underset{2 \text{ volumes}}{\text{Water vapours}}$$

Applying Avogadro's hypothesis

$$\underset{2n \text{ molecules}}{\text{Hydrogen}} + \underset{n \text{ molecules}}{\text{Oxygen}} \longrightarrow \underset{2n \text{ molecules}}{\text{Water vapours}}$$

or 1 molecule $\quad \frac{1}{2}$ molecule $\quad$ 1 molecule

Thus 1 molecule of water contains 1/2 molecule of oxygen. But 1 molecule of water contains 1 atom of oxygen. Hence

$\frac{1}{2}$ molecule of oxygen = 1 atom of oxygen

or 1 molecule of oxygen = 2 atoms of oxygen

i.e. atomicity of oxygen = 2.

(2) To find the relationship between molecular mass and vapour density of a gas.

$$\text{Vapour density of a gas} = \frac{\text{Density of the gas}}{\text{Density of hydrogen}}$$

$$= \frac{\text{Mass of certain volume of the gas at S.T.P.}}{\text{Mass of same volume of } H_2 \text{ at S.T.P}}$$

$$= \frac{\text{Mass of } n \text{ molecules of the gas}}{\text{Mass of } n \text{ molecules of } H_2}$$

$$= \frac{\text{Mass of 1 molecule of the gas}}{\text{Mass of 1 molecule of } H_2 \text{(i.e. 2 atoms of H)}}$$

$$= \frac{\text{Molecular mass}}{2}$$

or $\boxed{\text{Molecular mass} = 2 \times \text{Vapour density}}$

Vapour density is also called *'relative density'* of the gas.

(3) To find the relationship between mass and volume of a gas.

Molecular mass = 2 × Vapour density

$$= 2 \times \frac{\text{Mass of certain volume of the gas at S.T.P}}{\text{Mass of same volume of } H_2 \text{ at S.T.P}}$$

$$= 2 \times \frac{\text{Mass of 1L of the gas at S.T.P}}{\text{Mass of 1L of } H_2 \text{ at S.T.P}}$$

$$= 2 \times \frac{\text{Mass of 1 L of the gas at S.T.P}}{0 \cdot 089 \text{ g}}$$

$$= \frac{2}{0 \cdot 089} \times \text{Mass of 1L of the gas at S.T.P}$$

$$= 22 \cdot 4 \times \text{Mass of 1L of the gas at S.T.P.}$$

$$= \text{Mass of } 22 \cdot 4 \text{ L of the gas at S.T.P.}$$

Thus

22·4 litres of any gas at S.T.P. weigh equal to the molecular mass of the gas expressed in grams.* This is called Gram-Molecular Volume (G.M.V) Law

SECTION—VI

ATOMIC AND MOLECULAR MASSES AND MOLE CONCEPT

1.24. Relative Atomic Masses

Having known that matter is made up of atoms, the next immediate interest of the scientists was to determine the masses of the atoms. However as an atom is so small a particle that it cannot be seen or isolated, therefore, it is impossible to determine the actual mass of a single atom by weighing

*Earlier the STP conditions (i.e. Standard temperature and pressure) were taken as 1 atm and 0°C. However, now these are taken as 1 bar and 0°C. Under these conditions, instead of 22·4 L, we have 22·7 L. (1 atm = 1·01 bar). However unless specified as 1 bar and 0°C, STP will be taken as 1 atm and 0°C in this book.

it. Alternatively, the mass of an atom could have been calculated by weighing a large sample of the element and then dividing by the number of atoms contained in it but unfortunately there was no method known to count the number of atoms. The problem was finally solved by Avogadro's hypothesis which states that "Equal volumes of different gases under similar conditions of temperature and pressure contain equal number of molecules." Thus if equal volumes of two different gases are taken under similar conditions of temperature and pressure and then weighed, the ratio of their masses will be equal to the ratio of the masses of their single molecules (because they contained equal number of molecules). For example, taking equal volumes of hydrogen and oxygen, their masses are always found in the ratio of 1 : 16. This means that a molecule of oxygen is 16 times heavier than a molecule of hydrogen. Further as one molecule of hydrogen contains two atoms of hydrogen and one molecule of oxygen contains two atoms of oxygen, it may be interpreted that an atom of oxygen is 16 times heavier than an atom of hydrogen. Thus though the actual masses of the atoms could not be determined but their relative masses could be determined. It follows that if the atomic mass of hydrogen is taken as one, the relative atomic mass of oxygen is 16 (or more accurately it comes out to be 15.88).

In the beginning, the atomic masses of all the elements were obtained by comparing with the mass of hydrogen taken as 1 (because it was the lightest element). But by doing so the atomic masses of most of the elements came out to be fractional. Hence the reference was changed to oxygen taken as 16. However, a still better reference which is now widely accepted, has been found to be carbon taken as 12. On this basis the relative mass of hydrogen comes out to 1.008 and that of oxygen as 15.9994 (or 16). Hence atomic mass of an element may be defined as follows : –

The atomic mass of an element is the number of times an atom of that element is heavier than an atom of carbon taken as 12.

It may be noted that the atomic masses as obtained above are the relative atomic masses and not the actual masses of the atoms. These masses on the atomic mass scale are expressed in terms of *atomic mass units* (abbreviated as amu).

One atomic mass unit (amu) is equal to $\frac{1}{12}$ th of the mass of an atom of carbon-12 isotope.

Thus the atomic mass of hydrogen is 1.008 amu while that of oxygen is 15.9994 amu (or taken as 16 amu).

Prior to 1961, two different standards were used for expressing atomic masses when oxygen was used as the reference. Chemists on their scale (called *chemical scale*) assigned an exact value of 16 to the *average* mass of oxygen atoms as they occur in nature. On the *physical scale*, the isotope $^{16}_{8}O$ was assigned an exact value of 16. As naturally occurring oxygen consists of three isotopes (with mass numbers 16, 17 and 18), the average value for oxygen on physical scale comes out to be 16·0044. To remove this duality, all atomic masses are now expressed on the $^{12}_{6}C$ *scale*, taking mass of this isotope as exactly 12. This scale is called **unified** scale. On this scale, the symbol 'u' is used instead of amu, though the symbol 'amu' is still used quite often in place of 'u'. Thus now we better write are the atomic mass of hydrogen as 1·008 u and that of oxygen as 16 u.

The atomic masses of the elements have been determined accurately during the recent years using an instrument called **"mass spectrometer"**. It is found that in a number of cases, atoms of the same element possess different masses (called isotopes). Thus in such cases, the atomic mass of the element is taken as the average value. For example, ordinary chlorine is a mixture of two isotopes with atomic masses 35 u and 37 u and they are present in the ratio of 3 : 1. Hence the average atomic mass of chlorine would be

$$\frac{35 \times 3 + 37 \times 1}{3 + 1} = 35.5 \text{ u}$$

Evidently, the word 'average' must be included in the definition. Hence atomic mass of an element may more accurately be defined as under:

The atomic mass of an element is the average relative mass of its atoms as compared with an atom of carbon taken as 12.

Alternatively, the average relative atomic mass of an element can be calculated from the 'fractional abundances' of the isotopes of that element.

Fractional abundance *of an isotope is the fraction of the total number of atoms that is comprised of that particular isotope.*

For example, naturally occurring neon consists of three isotopes, the mass numbers of which and their fractional abundances are as follows :

Isotope	**Fractional Abundance**
^{20}Ne	0·9051
^{21}Ne	0·0027
^{22}Ne	0·0922

Average atomic mass of neon

$= 20 \times 0 \cdot 9051 + 21 \times 0 \cdot 0027 + 22 \times 0 \cdot 0922$

$= 20 \cdot 179\, u$

Lastly, it may be pointed out that naturally occurring carbon contains three isotopes of carbon with atomic masses 12, 13 and 14 u. It is the carbon-12 isotope that is taken as a reference on the atomic mass scale.

The atomic masses of a few common elements taking carbon-12 isotope as the reference are given in Table 1.4.

TABLE 1.4. Atomic masses of some common elements (taking C-12 = 12.000)

ELEMENT	SYMBOL	ATOMIC MASS	ELEMENT	SYMBOL	ATOMIC MASS
Aluminium	Al	27·0	Lead	Pb	207·2
Argon	Ar	39·9	Lithium	Li	6·94
Beryllium	Be	9·01	Magnesium	Mg	24·3
Boron	B	10·8	Neon	Ne	20·2
Calcium	Ca	40·1	Nitrogen	N	14·0
Carbon	C	12·0	Oxygen	O	16·0
Chlorine	Cl	35·5	Phosphorus	P	31·0
Copper	Cu	63·5	Potassium	K	39·1
Flourine	F	19·0	Silicon	Si	28·1
Helium	He	4·0	Silver	Ag	107·9
Hydrogen	H	1·008	Sodium	Na	23·0
Iodine	I	126·9	Sulphur	S	32·1
Iron	Fe	55·8	Uranium	U	238·0
			Zinc	Zn	65·4

EXAMPLE. *Use the data given in the following table to calculate the molar mass of naturally occurring argon.*

Isotope	*Isotopic molar mass*	*Abudance*
^{36}Ar	$35 \cdot 96755\ g\ mol^{-1}$	*0·337%*
^{38}Ar	$37 \cdot 96272\ g\ mol^{-1}$	*0·063%*
^{40}Ar	$39 \cdot 9624\ g\ mol^{-1}$	*99·600%*

(N.C.E.R.T.)

Solution. Molar mass of Ar

$= 35 \cdot 96755 \times 0 \cdot 00337 + 37 \cdot 96272 \times 0 \cdot 00063 + 39 \cdot 96924 \times 0 \cdot 99600$

$= \mathbf{39 \cdot 948\ g\ mol^{-1}}$.

Note. Also see Problems 18 page 1/77 and 29 page 1/78 of C.B.S.E. P.M.T. (Special) Problems given at the end of this unit for calculation of average atomic mass.

1.25. Gram Atomic Mass

The atomic mass of an element expressed in grams is called Gram atomic mass.

This amount of the element is also called ***one gram atom.***

e.g. Atomic mass of oxygen = 16 amu

Gram atomic mass of oxygen

(or one gram atom of oxygen) = 16 g

1.26. Molecular Mass

Molecular mass of a substance refers to the relative mass of its molecule and is defined in a manner similar to the atomic mass as follows :

The molecular mass of a substance (element or compound) is the number of times the molecule of the substance is heavier than 1/12 th the mass of an atom of carbon-12 isotope.

or

The molecular mass of a substance is the average relative mass of its molecules as compared with an atom of carbon-12 isotope taken as 12.

The molecular mass of a substance can be calculated by adding the atomic masses of all the atoms present in one molecule of the substance.

e.g. Molecular mass of H_2SO_4

$= 2 \times$ At. mass of H + At. mass of S + $4 \times$ At. mass of O.

$= 2 \times 1.0 + 32.0 + 4 \times 16.0 = 98.0$ u.

1.27. Gram Molecular Mass

The molecular mass of a substance expressed in grams is called it Gram molecular mass.

This amount of the substance is also called ***one gram molecule.***

e.g. Molecular mass of $H_2SO_4 = 98.0$ u

Gram molecular mass of H_2SO_4

(or one gram molecule of H_2SO_4) $= 98.0$ g

1.28. Mole Concept

In section 1.24, we discussed how the relative atomic masses of the elements could be determined. However, the problem of finding the absolute masses of the atoms remained unsolved until it became possible to count the number of atoms or molecules in a definite amount of the substance. It is found that one gram atom of any element contains the same number of atoms and one gram molecule of any substance contains the same number of molecules. This number has been experimentally determined and found to be equal to $6 \cdot 022137 \times 10^{23}$ correct upto seven significant figures. However, the value generally used is $6 \cdot 022 \times 10^{23}$. It is called as *'Avogadro's number'* or *'Avogadro's constant'* in honour of Amedeo Avogadro, a great pioneer in this field. It is usually represented by N. Hence

$$\boxed{\text{Avogadro's Number (N)} = 6 \cdot 022 \times 10^{23}}$$

Avogadro's number ***may be defined as the number of atoms present in one gram atom of the element or the number of molecules present in one gram molecule of the substance.***

The amount of the substance containing Avogadro's number of atoms or molecules is called a *mole* (or simply written as *mol*). *Thus a mole is a chemist's unit of counting particles* such as atoms, molecules, ions, electrons, protons etc. which represents a value of $6 \cdot 022 \times 10^{23}$ just as a dozen for 12, score for 144 etc. are used to count different objects. Since the term mole is applicable to all sorts of particles, while using this term, it is very important to indicate the nature of the particles under consideration *i.e.* whether they are atoms, molecules or something else. For example, it is wrong to speak 'one mole of hydrogen'. We must specify whether we are referring to hydrogen atoms or hydrogen molecules. Thus :

A mole of hydrogen atoms means $6 \cdot 022 \times 10^{23}$ atoms of hydrogen whereas a mole of hydrogen molecules means $6 \cdot 022 \times 10^{23}$ molecules of hydrogen, or $2 \times 6 \cdot 022 \times 10^{23}$ atoms of hydrogen (because each hydrogen molecule contains 2 atoms of hydrogen).

Similarly, a mole of oxygen molecules means $6 \cdot 022 \times 10^{23}$ molecules of oxygen or $2 \times 6 \cdot 022 \times 10^{23}$ atoms of oxygen.

A mole of sand particles means $6 \cdot 022 \times 10^{23}$ sand particles.

A mole of electrons means $6 \cdot 022 \times 10^{23}$ electrons and so on.

Generally, when we speak of 1 mole of hydrogen, it implies 1 mole of hydrogen molecules i.e. the natural form in which it exists. However to avoid confusion, in such cases, the molecular forms are called dihydrogen (H_2) dioxygen (O_2), dinitrogen (N_2) etc.

It is interesting to note that the magnitude of Avogadro's number is very large. Its value may be imagined from the fact that this number is about 10^{14} times the population of the entire world or Avogadro's number of marbles are sufficient to cover the surface of the entire earth with 5 km thick layer. Thus, *mole is not a useful measure for things which are much larger than atoms and molecules in size or mass.* Further, as 1 mole represents a very large number of particles, therefore, being SI unit, it can be used with prefixes. For example,

1 mmol $= 10^{-3}$ mol, 1 μ mol $= 10^{-6}$ mol

1 nmol $= 10^{-9}$ mol etc.

It is usually convenient to express large number of atoms or molecules in terms of moles. For example, suppose the sample of a substance (*e.g.* vitamin C) contains $2 \cdot 1 \times 10^{24}$ atoms of hydrogen. Then number of moles of H-atoms present in the sample= $(2 \cdot 1 \times 10^{24})/(6 \cdot 022 \times 10^{23}) = 3 \cdot 5$ mol.

Further, according to Avogadro's hypothesis 'Equal volumes of different gases under similar conditions of temperature and pressure contain equal number of molecules'. This means that

$6 \cdot 022 \times 10^{23}$ molecules of any gas at STP (*i.e.* standard temperature and pressure *viz* 0°C and atmospheric pressure) must have the same volume. This volume has been experimentally found to be 22·4 litres at STP (0°C, 1 atm or 1·01 bar pressure) and is called **molar volume*** (because $6 \cdot 022 \times 10^{23}$ molecules represent one mole of the gas, as already discussed).

Keeping in view the different aspects discussed above, a mole may be defined in any one of the following three ways :

1st Definition (in terms of mass). *A mole is defined as that amount of the substance which has mass equal to gram atomic mass if the substance is atomic or gram molecular mass if the substance is molecular.*

e.g. 1. Mole of carbon atoms = 12 g
(∵ At. mass of C = 12 u)

1 Mole of sodium atoms = 23 g
(∵ At. mass of Na = 23 u)

1 Mole of oxygen atoms = 16 g
(∵ At. mass of O = 16 u)

1 Mole of oxygen molecules = 32 g
(∵ Mol mass of O_2 = 32 u)

1 Mole of H_2O molecules = 18 g
(∵ Mol mass of H_2O = 18 u)

1 Mole of CO_2 molecules = 44 g
(∵ Mol mass of CO_2 = 44 u)

2nd Definition (in terms of number). *A mole is defined as that amount of the substance which contains Avogadro's number ($6 \cdot 022 \times 10^{23}$) of atoms if the substance is atomic or Avogadro's number ($6 \cdot 022 \times 10^{23}$) of molecules if the substance is molecular.*

1 Mole of carbon atoms
$= 6 \cdot 022 \times 10^{23}$ atoms of carbon

1 Mole of sodium atoms
$= 6 \cdot 022 \times 10^{23}$ atoms of sodium

1 Mole of oxygen atoms
$= 6 \cdot 022 \times 10^{23}$ atoms of oxygen

1 Mole of oxygen molecules
$= 6 \cdot 022 \times 10^{23}$ molecules of oxygen

1 Mole of H_2O molecules
$= 6 \cdot 022 \times 10^{23}$ molecules of H_2O

1 Mole of CO_2 molecules
$= 6 \cdot 022 \times 10^{23}$ molecules of CO_2

3rd Definition (in terms of volume). *In case of gases, a mole is defined as that amount of the gas which has a volume of 22·4 litres at STP.*

e.g. 1 Mole of oxygen gas = 22·4 litres of oxygen at STP

1 Mole of CO_2 gas = 22·4 litres of CO_2 at STP

REMEMBER

In C.G.S. system,

1 g mole = Molecular mass expressed in grams $= 6 \cdot 022 \times 10^{23}$ molecules

In M.K.S. system,

1 kg mole = Molecular mass expressed in kg $= 6 \cdot 022 \times 10^{26}$ molecules

Mole Concept for Ionic Compounds. The formula of an ionic compound does not represent a molecule of that compound but expresses only the relative ratio of their constituent ions. However, the term mole is applicable to these compounds also with the modification that the term **'formula mass'** is used in place of 'molecular mass' and the term 'formula unit' is used in place of 'molecule'. Thus

A mole of an ionic compound is defined as that amount of the substance which has mass equal to gram formula mass (i.e. formula mass expressed in grams) or which contains Avogadro's number ($6 \cdot 022 \times 10^{23}$) of formula units.

e.g. 1 Mole of NaCl = 58·5 g of NaCl
(∵ Formula mass of NaCl = 23+35·5
= 58·5 amu)

Also, 1 Mole of NaCl $= 6 \cdot 022 \times 10^{23}$ formula units of NaCl
$= 6 \cdot 022 \times 10^{23}$ Na^+ ions
and $6 \cdot 022 \times 10^{23}$ Cl^- ions

Similarly, 1 Mole of Na_2CO_3 = 106 g of Na_2CO_3
(∵ Formula mass of Na_2CO_3
$= 2 \times 23 + 12 + 3 \times 16 = 106$ amu)

Also, 1 Mole of Na_2CO_3 $= 6 \cdot 022 \times 10^{23}$ formula units of Na_2CO_3
$= 2 \times 6 \cdot 022 \times 10^{23}$ Na^+ ions
and $6 \cdot 022 \times 10^{23}$ CO_3^{2-} ions

(which are further equivalent to $6 \cdot 022 \times 10^{23}$ carbon atoms and $3 \times 6 \cdot 022 \times 10^{23}$ oxygen atoms).

*In general, *molar volume of a substance is the volume of one mole of that substance.*

SI Definition of Mole. The SI definition of mole is as follows : –

A mole is that amount of the substance which contains as many elementary entities as there are atoms in exactly 0·012 kg (i.e. 12g) of carbon – 12 isotope. The elementary entities must be specified i.e. whether they are atoms, molecules, ions, electrons or any other entity.

Atomic Mass and Molecular Mass in terms of Avogadro's Number and Mole Concept. As one mole of any atomic substance is equal to its gram atomic mass and it contains Avogadro's number ($6 \cdot 022 \times 10^{23}$) of atoms. Hence

Gram atomic mass *of an element may be defined as the mass of Avogadro's number of atoms. Similarly,* **Gram molecular mass** *of a substance may be defined as the mass of Avogadro's number ($6 \cdot 022 \times 10^{23}$) of molecules.*

The mass of 1 mole of any substance is usually called its **molar mass***. Evidently the molar mass is equal to the atomic mass or the molecular mass expressed in grams depending upon whether the substance is atomic or molecular. As molar mass is the mass of Avogadro's number of atoms if the substance is atomic, therefore, dividing the molar mass by Avogadro's number, the absolute mass of a single atom can be obtained. Similarly, the absolute mass of a single molecule can be calculated.

Importance of Avogadro's Number and Mole Concept. Avogadro's number and mole concept help in the chemical calculation in a number of ways as follows : –

1. In the calculation of the actual mass of a single atom of an element or a single molecule of a substance.

2. In the calculation of the number of atoms or molecules in a given mass of the element or the compound.

3. In the calculation of the number of molecules present in a given volume of the gas under given conditions.

4. In the calculation of the sizes of the individual atoms or molecules assuming them to be spherical.

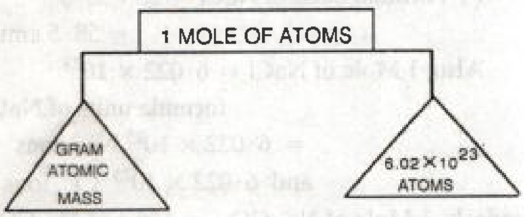
These calculations can be understood by going through the solved examples given below :

PROBLEMS ON AVOGADRO'S NUMBER AND MOLE CONCEPT

Recapitulation of Formulas

(i) In case of atomic substances,

1 Mole = Gram atomic mass or 1 Gram atom of the element



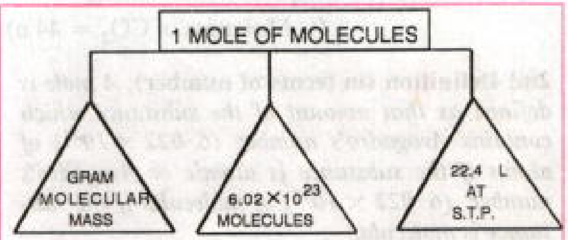
= $6 \cdot 022 \times 10^{23}$ atoms of the element.

(ii) In case of molecular substances,

1 Mole = Gram molecular mass or 1 Gram molecule

= $6 \cdot 022 \times 10^{23}$ molecules of the substance

= $22 \cdot 4$ litres at STP if the substance is a gas.



(iii) In case of ionic compounds,

1 Mole = Gram formula mass of the compound

= $6 \cdot 022 \times 10^{23}$ formula units of the compound

= $6 \cdot 022 \times 10^{23}$ times the number of atoms present in one formula unit of the compound.

(iv) $\dfrac{\text{Mass of substance}}{\text{Molar mass}} = \text{Moles}$

or Mass of substance** = Moles × Molar mass

$\dfrac{\text{Number of entities}}{\text{Avogadro's No.}} = \text{Moles}$

or Number of entities = Moles × Avogadro's No.

*Obviously the units of molar mass are g mol^{-1} or kg mol^{-1}.

**Molar volume of a liquid or a solid can be obtained by dividing the molar mass by density of the substance at the given temperature and pressure.

TYPE I. On the calculation of the actual mass of a single atom or a single molecule

EXAMPLE 1. *Calculate the mass of (i) an atom of silver (ii) a molecule of carbon dioxide.*

Solution. (*i*) 1 mole of Ag atoms = 108 g

($\because$ Atomic mass of silver = 108 u)

$= 6 \cdot 022 \times 10^{23}$ atoms

$6 \cdot 022 \times 10^{23}$ atoms of silver have mass = 108 g

$\therefore$ Mass of one atom of silver

$$= \frac{108}{6 \cdot 022 \times 10^{23}} = \mathbf{1 \cdot 793 \times 10^{-22}\ g}$$

(*ii*) 1 mole of CO_2 = 44 g

($\because$ Molecular mass of CO_2 $= 1 \times 12 + 2 \times 16 = 44$ u)

$= 6 \cdot 022 \times 10^{23}$ molecules

Thus, $6 \cdot 022 \times 10^{23}$ molecules of CO_2 have mass = 44g

$\therefore$ 1 molecule of CO_2 has mass

$$= \frac{44}{6 \cdot 022 \times 10^{23}} = \frac{44 \times 10^{-23}}{6 \cdot 022}\ g$$

$$= \mathbf{7 \cdot 307 \times 10^{-23}\ g}$$

TYPE II. On the calculation of number of atoms or molecules in a given mass of the substance

EXAMPLE 2. *How many atoms and molecules of sulphur are present in 64·0 g of sulphur (S_8) ?*

Solution. Molecular formula of sulphur = S_8

$\therefore$ Molecular mass of sulphur (S_8) $= 32 \times 8$

$= 256 \cdot 0$ u

1 mole of sulphur molecules = 256 g

$= 6 \cdot 022 \times 10^{23}$ molecules of sulphur

Now, 256 g of sulphur contain $6 \cdot 022 \times 10^{23}$ molecules

$\therefore$ 64 g of sulphur will contain

$$= \frac{6 \cdot 022 \times 10^{23} \times 64}{256}$$

$$= \mathbf{1 \cdot 506 \times 10^{23}\ molecules.}$$

1 molecule of sulphur (S_8) contains 8 atoms of sulphur

$\therefore$ $1 \cdot 506 \times 10^{23}$ molecules of sulphur will contain sulphur atoms $= 8 \times 1 \cdot 506 \times 10^{23}$

$= \mathbf{1 \cdot 2048 \times 10^{24}\ atoms.}$

EXAMPLE 3. *Calculate the number of molecules present*

(i) in 34·20 grams of cane sugar ($C_{12}H_{22}O_{11}$)

(ii) in one litre of water assuming that the density of water is 1 g/cm³.

(iii) in one drop of water having mass 0·05 g.

Solution. (*i*) 1 mole of $C_{12}H_{22}O_{11}$ = 342 g

[$\because$ Molecular mass of cane sugar ($C_{12}H_{22}O_{11}$)

$= 12 \times 12 + 22 \times 1 + 11 \times 16 = 342$ amu]

$= 6 \cdot 022 \times 10^{23}$ molecules

Now 342 g of cane sugar contain $6 \cdot 022 \times 10^{23}$ molecules

$\therefore$ 34·2 g of cane sugar will contain

$$= \frac{6 \cdot 022 \times 10^{23}}{342} \times 34 \cdot 2$$

$$= \mathbf{6 \cdot 022 \times 10^{22}\ molecules}$$

(*ii*) 1 mole of water = 18 g = $6 \cdot 022 \times 10^{23}$ molecules

Mass of 1 litre of water = Volume $\times$ density

$= 1000 \times 1 = 1000$ g

Now 18 g of water contains = $6 \cdot 022 \times 10^{23}$ molecules

$\therefore$ 1000 g of water will contain

$$= \frac{6 \cdot 022 \times 10^{23} \times 1000}{18}$$

$$= \mathbf{3 \cdot 346 \times 10^{25}\ molecules}$$

(*iii*) 1 mole of H_2O = 18g = $6 \cdot 022 \times 10^{23}$ molecules

Mass of 1 drop of water = 0·05 g

Now 18 g of H_2O contain = $6 \cdot 022 \times 10^{23}$ molecules

$\therefore$ 0·05 g of H_2O will contain

$$= \frac{6 \cdot 022 \times 10^{23}}{18} \times 0 \cdot 05$$

$$= \mathbf{1 \cdot 673 \times 10^{21}\ molecules}$$

EXAMPLE 4. *Calculate the number of atoms of the constituent elements in 53 g of Na_2CO_3.*

Solution. *By mole concept,*

1 mole of Na_2CO_3

= Gram formula mass of Na_2CO_3

$= 2 \times 23 + 12 + 16 \times 3 = 106$ g

Now, 106 g of Na_2CO_3 = 1 mole

$\therefore$ 53 g of $Na_2CO_3 = \frac{1 \times 53}{106}$

$= 0\cdot5$ mole of Na_2CO_3

But 1 mole of Na_2CO_3 contains 2 moles of Na^+ ion or $2 \times 6\cdot022 \times 10^{23}$ Na^+ ions.

$\therefore$ 0·5 mole of Na_2CO_3 will contain

$2 \times 6\cdot022 \times 10^{23} \times 0\cdot5$ Na^+ ions

$= \mathbf{6\cdot022 \times 10^{23}}$ **Na^+ ions**

Again, 1 mole of Na_2CO_3 contains 1 mole of carbon atoms $= 6\cdot022 \times 10^{23}$ carbon atoms

$\therefore$ 0·5 mole of Na_2CO_3 will contain

$= 6\cdot022 \times 10^{23} \times 0\cdot5$ carbon atoms

$= \mathbf{3\cdot011 \times 10^{23}}$ **carbon atoms**

Further, 1 mole of Na_2CO_3 contains 3 moles of oxygen atoms or $3 \times 6\cdot022 \times 10^{23}$ oxygen atoms

$\therefore$ 0·5 mole of Na_2CO_3 will contain

$3 \times 6\cdot022 \times 10^{23} \times 0\cdot5$ oxygen atoms

$= \mathbf{9\cdot033 \times 10^{23}}$ **oxygen atoms**

TYPE III. On the calculation of number of molecules present in a given volume of a gas under given conditions

EXAMPLE 5. *Calculate the number of molecules present in 350 cm³ of NH_3 gas at 273 K and 2 atmosphere pressure.*

Solution. First of all, we have to determine the volume of the gas at STP.

Given conditions	*At STP*
$V_1 = 350$ cm³	$V_2 = ?$
$T_1 = 273$ K	$T_2 = 273$ K
$P_1 = 2$ atmospheres	$P_2 = 1$ atm

Applying gas equation : $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$

we get $\frac{350 \times 2}{273} = \frac{1 \times V_2}{273}$

or $V_2 = \frac{350 \times 2}{273} \times \frac{273}{1} = 700$ cm³

By mole concept,

1 mole of $NH_3 = 6\cdot022 \times 10^{23}$ molecules

$= 22400$ cm³ at STP

Thus, 22400 cm³ of NH_3 at STP contain $6\cdot022 \times 10^{23}$ molecules

$\therefore$ 700 cm³ of NH_3 at STP will contain

$= \frac{6\cdot022 \times 10^{23}}{22400} \times 700$

$= \mathbf{1\cdot882 \times 10^{22}}$ **molecules**

TYPE IV. On the calculation of sizes of individual atoms or molecules

EXAMPLE 6. *(i) Assuming the density of water to be 1g/cm³, calculate the volume occupied by one molecule of water.*

(ii) Assuming the water molecule to be spherical, calculate the diameter of the water molecule.

(iii) Assuming that oxygen atom occupies half of the volume occupied by the water molecule, calculate approximately the diameter of the oxygen atom.

Solution. (*i*) 1 mole of $H_2O = 18$g $= 18$ cm³

($\because$ density of $H_2O = 1$ g/cm³)

$= 6\cdot022 \times 10^{23}$ molecules of H_2O

Thus $6\cdot022 \times 10^{23}$ molecules of H_2O have volume = 18 cm³

$\therefore$ 1 molecule of H_2O will have volume

$= \frac{18}{6\cdot022 \times 10^{23}}$ cm³

$= \mathbf{2\cdot989 \times 10^{-23}}$ **cm³**

(*ii*) As water molecule is assumed to be spherical, if R is its radius, then its volume will be

$\frac{4}{3} \pi R^3 = 2\cdot989 \times 10^{-23}$ cm³

or $R^3 = 7\cdot133 \times 10^{-24}$

or $R = (7\cdot133)^{1/3} \times 10^{-8}$

$= 1\cdot925 \times 10^{-8}$ cm

Take $n = (7\cdot133)^{1/3}$

$\therefore \log n = \frac{1}{3} \log 7\cdot133 = \frac{1}{3} \times 0\cdot8533$

$= 0\cdot2844$

$n = \text{Antilog } 0\cdot2844 = 1\cdot925$

$\therefore$ Diameter of water molecule

$= 2 \times 1\cdot925 \times 10^{-8}$ cm

$= \mathbf{3\cdot85 \times 10^{-8}}$ **cm**

(*iii*) As oxygen atom occupies half of the volume occupied by water molecule, hence if *r* is the radius of oxygen atom, then

$$\frac{4}{3}\pi r^3 = \frac{1}{2} \times 2\cdot989 \times 10^{-23}\ \text{cm}^3$$

or $r^3 = 3\cdot566 \times 10^{-24}$

which gives $r = 1\cdot528 \times 10^{-8}$ cm

$\therefore$ Diameter of oxygen atom

$= 2 \times 1\cdot528 \times 10^{-8}$ cm

$= \mathbf{3\cdot056 \times 10^{-8}}$ **cm.**

TYPE V. On the conversion of moles into mass, volume or number of atoms/molecules and vice versa

EXAMPLE 7. *Calculate the number of moles in each of the following :–*

(i) 392 grams of sulphuric acid (ii) 44·8 litres of carbon dioxide at STP (iii) $6\cdot022 \times 10^{23}$ molecules of oxygen (iv) 9·0 grams of aluminium (v) 1 metric ton of iron (1 metric ton = 10^3 kg) (vi) 7·9 mg of Ca (vii) 65 μg of carbon.

Solution. (*i*) 1 mole of H_2SO_4 = 98 g.

($\because$ Molecular mass of H_2SO_4 $= 2 \times 1 + 32 + 4 \times 16 = 98$ u)

Thus 98 g of H_2SO_4 = 1 mole of H_2SO_4

$\therefore$ 392 g of $H_2SO_4 = \frac{1}{98} \times 392$

$=$ **4 moles of H_2SO_4.**

(*ii*) 1 mole of CO_2 = 22·4 litres at STP

i.e., 22·4 litres of CO_2 at STP = 1 mole

$\therefore$ 44·8 litres of CO_2 at STP $= \frac{1}{22\cdot4} \times 44\cdot8$

$=$ **2 moles CO_2**

(*iii*) 1 mole of O_2 molecules $= 6\cdot022 \times 10^{22}$ molecules

$6\cdot022 \times 10^{23}$ molecules = **1 mole of oxygen molecules.**

(*iv*) 1 mole of Al = 27 g of Al

($\because$ Atomic mass of aluminium = 27 u)

i.e. 27 g of aluminium = 1 mole of Al

$\therefore$ 9 g of aluminium $= \frac{1}{27} \times 9$

$=$ **0·33 mole of Al**

(*v*) 1 metric ton of Fe $= 10^3$ kg $= 10^6$ g

1 mole of Fe = 56 g of Fe

$\therefore$ 10^6 g of Fe $= \frac{10^6}{56}$ moles

$= 1\cdot786 \times 10^4$ moles.

(*vi*) 7·9 mg of Ca $= 7\cdot9 \times 10^{-3}$ *g* of Ca

$= \frac{7\cdot9 \times 10^{-3}}{40}$ mol (At. mass of Ca = 40 *u*)

(*vii*) 65·5 μg of C $= 65\cdot5 \times 10^{-6}$ *g* of C

$= \frac{65\cdot5 \times 10^{-6}}{12}$ mol $=$ $\mathbf{5\cdot458 \times 10^{-6}}$ **mol**

EXAMPLE 8. *Calculate the mass of (i) 0·1 mole of KNO_3 (ii) 1×10^{23} molecules of methane and (iii) 112 cm^3 of hydrogen at STP.*

Solution. (*i*) 1 mole of KNO_3 = 101 g

($\because$ Formula mass of KNO_3 $= 1 \times 39 + 1 \times 14 + 3 \times 16 = 101$ u)

$\therefore$ 0·1 mole of KNO_3 = 101 × 0·1

= **10·1 g of KNO_3**

(*ii*) 1 mole of CH_4 = 16 g

$= 6\cdot022 \times 10^{23}$ molecules

i.e. $6\cdot022 \times 10^{23}$ molecules of methane have mass = 16 g

$\therefore$ 1×10^{23} molecules of methane will have mass $= \frac{16}{6\cdot022 \times 10^{23}} \times 10^{23} =$ **2·657 g.**

(*iii*) 1 mole of H_2 = 2 g = 22400 cm^3 at STP

i.e. 22400 cm^3 of H_2 at STP have mass = 2 g

$\therefore$ 112 cm^3 of H_2 at STP will have mass

$= \frac{2}{22400} \times 112 =$ **0·01 g.**

EXAMPLE 9. *Arrange the following in order of their increasing masses in grams ?*

(i) One atom of silver, (ii) One gram-atom of nitrogen, (iii) One mole of calcium, (iv) One mole of oxygen molecules (v) 10^{23} atoms of carbon and (vi) One gram of iron.

Solution. (*i*) 1 mole of Ag atoms = 108 g

$= 6\cdot022 \times 10^{23}$ atoms

i.e. mass of $6\cdot022 \times 10^{23}$ atoms of Ag = 108 g

$\therefore$ Mass of 1 atom of Ag $= \frac{108}{6\cdot022 \times 10^{23}}$

$= \mathbf{1\cdot793 \times 10^{-22}}$ **g**

(*ii*) Mass of one gram atom of N = At. mass in grams = **14·0 g**

(*iii*) Mass of one mole of Ca = At. mass in grams = **40·0 g** (*iv*) Mass of one mole of oxygen molecules

$= $ Mol. mass in grams $= 32 \cdot 0$ g

(*v*) 1 mole of C atoms $= 12$ g

$= 6 \cdot 022 \times 10^{23}$ atoms

i.e., Mass of $6 \cdot 022 \times 10^{23}$ atoms of C $= 12 \cdot 0$ g

$\therefore$ Mass of 10^{23} atoms of C

$$= \frac{12}{6 \cdot 022 \times 10^{23}} \times 10^{23} = 1 \cdot 993 \text{ g}$$

(*vi*) Mass of iron $= 1 \cdot 0$ g (Given)

Thus the order of increasing masses is :

One atom of silver < one gram of iron < 10^{23} atoms of carbon < one gram-atom of nitrogen < one mole of oxygen < one mole of calcium.

EXAMPLE 10. *Calculate the volume at STP occupied by (i) 14 g of nitrogen, (ii) 1·5 moles of carbon dioxide and (iii) 10^{21} molecules of oxygen.*

Solution. (*i*) Molecular mass of nitrogen $= 28$ u

1 mole of nitrogen = 28 g = 22.4 litres at STP

i.e. 28 g of nitrogen occupy 22.4 litres at STP

$\therefore$ 14 g of nitrogen will occupy $= \frac{22.4}{28} \times 14$

= 11.2 litres at STP

(*ii*) 1 mole of carbon dioxide = 22.4 litres at STP

$\therefore$ 1.5 moles of carbon dioxide will occupy $\frac{22.4}{1} \times 1.5$ **= 33.6 litres at STP**

(*iii*) 1 mole of O_2 molecules

$= 6.022 \times 10^{23}$ molecules $= 22.4$ litres at STP

i.e 6.022×10^{23} molecules of oxygen occupy 22.4 litres at STP

10^{21} molecules of oxygen will occupy

$= \frac{22.4}{6.022 \times 10^{23}} \times 10^{21}$ litres at STP

$= 3.72 \times 10^{-2}$ litres at STP

$= 3.72 \times 10^{-2} \times 10^{3}$ cm^3 at STP

= 37.2 cm^3 at STP

PROBLEMS FOR PRACTICE

1. (*a*) What is the mass of
   (*i*) 1 mole of water (*ii*) 0.5 mole of CO_2
   (*iii*) 2.5 moles of Cl_2 ?
   [Ans. (*i*) 18.0 g (*ii*) 22.0 g (*iii*) 177.5 g]
   (*b*) How many moles of atoms are contained in :
   (*i*) 9.0 g of Aluminium (*ii*) 0.8 g of Iron ?
   [Ans. (*i*) $\frac{1}{3}$ mole (*ii*) $\frac{1}{70}$ mole]
2. Calculate the number of moles in each of the following amounts of materials
   (*i*) 10.0 g of $CaCO_3$
   (*ii*) 1×10^{23} molecules of CO_2
   [Ans. (*i*) 0.1 mole (*ii*) 0.166 mole]
3. What is the mass in grams of :
   (*i*) 6.022×10^{23} atoms of oxygen ?
   (*ii*) 1.0×10^{23} molecules of H_2S ?
   (*iii*) 6.022×10^{23} molecules of oxygen ?
   (*iv*) 1.5 moles of H_2SO_4 ?
   [Ans. (*i*) 16.0g (*ii*) 5.645 g (*iii*) 32.0 g (*iv*) 147.0 g]
4. Calculate the number of atoms in each of the following
   (*i*) 52 moles of He (*ii*) 52 *u* of He
   (*iii*) 52 g of He (*N.C.E.R.T.*)
   [Ans. (*i*) 3.1314×10^{25} (*ii*) 13 (*iii*) 7.8286×10^{24}]
5. Which of the following weighs most ?
   (*i*) 50 g of iron
   (*ii*) 5 g atoms of nitrogen
   (*iii*) 0.1 g atom of silver
   (*iv*) 1×10^{23} atoms of carbon
   [Ans. (*i*) 50.0 g (*ii*) 70.0 g (*iii*) 10.8 g (*iv*) 1.992 g. Hence 5g atoms of nitrogen weigh most]
6. Calculate the number of molecules present in 22 g of CO_2 ?
   [Ans. 3.011×10^{23} molecules]
7. Calculate the mass of CO_2 which contains the same number of molecules as are contained in 40 g of oxygen. **[Ans. 55 g]**
8. Calculate the mass of Na_2CO_3 which will have the same number of molecules as contained in 12.3 g of $MgSO_4.7H_2O$. **[Ans. 5.3 g]**
9. Calculate the volume occupied by 10^{22} molecules of a gas at 300 K and 760 mm pressure.
   [Ans. 408.9 cm^3]
10. Find the number of atoms of each type present in 3.42 grams of cane sugar ($C_{12}H_{22}O_{11}$).
    [Ans. C = 7.226×10^{22} atoms, H = 1.325×10^{23} atoms]

1 g H   1 g of H = 1 mol = 6.023 × 10^{23}

PRACTICE PROBLEMS CONTD.

11. Calculate the mass of 1 molecule of —
(*i*) oxygen (*ii*) ammonia

[Ans. (*i*) 5.314 × 10^{-23} g (*ii*) 2.823 × 10^{-23} g]

12. (*a*) Calculate the volume occupied at STP by
(*i*) 16.0 g of oxygen (*ii*) 1.5 moles of oxygen and
(*iii*) 6.022 × 10^{23} molecules of carbon dioxide

[Ans. (*i*) 11.20 litres (*ii*) 33.60 litres (*iii*) 22.4 litres]

13. (*i*) How many grams of H_2S are contained in 0.40 mole of H_2S ? **[Ans. 13.6 g]**

(*ii*) How many gram atoms of H and S are contained in 0.40 mole of H_2S ?

[Ans. 0.8 g atoms of H and 0.4 g atoms of S]

(*iii*) How many molecules of H_2S are contained in 0.40 mole of H_2S ? **[Ans. 2.4088×10^{23} molecules]**

(*iv*) How many atoms of H and S are contained in 0.40 mole of H_2S ? **[Ans. 4.8176×10^{23} atoms of H and 2.4088 × 10^{23} atoms of S]**

14. You are supplied with a gas containing 0.32 g of oxygen. Calculate the number of moles and number of molecules present in it.

[Ans. 0.01 mole, 6.022 × 10^{21} molecules]

15. The mass of a litre of oxygen at standard conditions of temperature and pressure is 1.43 g and that of a litre of SO_2 is 2.857 g.

(*i*) How many molecules of each gas are there in this volume ? **[Ans. 2.688 × 10^{22} molecules]**

(*ii*) What is the mass in grams of a single molecule of each gas ? **[Ans. O_2 = 5.320 × 10^{-23} g and SO_2 = 10.629 × 10^{-23} g]**

(*iii*) What are the molecular masses of SO_2 and O_2 respectively ? **[Ans. O_2 = 32.032 u and SO_2 = 63.997 u]**

16. The mass of 350 cm^3 of a diatomic gas at 273 K at 2 atmospheres pressure is one gram. Calculate the mass of one atom of the gas.

[Ans. 2.657 × 10^{-23}g]

17. How many atoms and molecules of phosphorus are present in 124 g of phosphorus (P_4) ?

[Ans. 24.088×10^{23} atoms and 6.022×10^{23} molecules]

18. What is the mass of a water molecule in gram ? How many molecules are present in one drop of pure water which weighs 0.05 g. If the same drop of water evaporates in one hour, calculate the number of molecules leaving the liquid surface per second.

[Ans. 2.989 × 10^{-23} g, 1.673 × 10^{21} molecules and 4.647 × 10^{17} molecules/sec.]

19. What is the mass of carbon present in 0.5 mole of $K_4[Fe(CN)_6]$? **[Ans. 36 g]**

20. The cost of table salt (NaCl) and table sugar ($C_{12}H_{22}O_{11}$) is Rs. 2 per kg and Rs. 6 per kg respectively. Calculate their costs per mole.

[Ans. Salt = 12 p, Sugar = Rs. 2.05 p]

21. Chlorophyll, the green colouring matter of plants responsible for photosynthesis, contains 2.68% of magnesium by weight. Calculate the number of magnesium atoms in 2.0 g of chlorophyll.

(N.C.E.R.T.) **[Ans. 1.345 × 10^{21}]**

22. How much time would it take to distribute one Avogadro's number of wheat grains if 10^{10} grains are distributed each second ?

[Ans. 190,893 years approx.]

23. Calculate the total number of electrons present in 1·4 g of nitrogen gas.

[Ans. 4·214 × 10^{23} electrons]

HINTS FOR DIFFICULT PROBLEMS

4. (*ii*) 4 *u* of He ≡ 1 atom of He

7. 40 g O_2 = 40/32 moles = 1·25 moles

1·25 moles of CO_2 = 1·25 × 44 *g* = 55 *g*

(Equal moles contain equal number of molecules)

8. $12 \cdot 3\, g\, MgSO_4 . 7\,H_2O = \frac{12 \cdot 3}{(24+32+64+126)}$ mol

$= \frac{12 \cdot 3}{246}$ mol = 0·05 mol

0·05 mol Na_2CO_3 = 0·05 × 106 g = 5·3 g

9. Volume occupied by 10^{22} molecules at STP

$$= \frac{22400}{6 \cdot 02 \times 10^{23}} \times 10^{22}\ cm^3 = 372 \cdot 1\ cm^3$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2};\ \frac{760 \times V_1}{300} = \frac{760 \times 372 \cdot 1}{273}$$

or $V_1 = 408 \cdot 9\ cm^3$

HINTS CONTD.

15. 1 L O_2 or $SO_2 = \frac{1}{22 \cdot 4}$ mol

$= \frac{1}{22 \cdot 4} \times 6 \cdot 02 \times 10^{23}$ molecules

$= 2 \cdot 688 \times 10^{22}$ molecules

Mass of 1 molecule of $O_2 = \frac{1 \cdot 43}{2 \cdot 688 \times 10^{22}} g$

$= 5 \cdot 320 \times 10^{-23} g$

Mass of 1 molecule of $SO_2 = \frac{2 \cdot 857}{2 \cdot 688 \times 10^{22}}$

$= 1 \cdot 0629 \times 10^{-23} g$

Gram Molecular mass of O_2

= Mass of 22·4 L at STP

$= 1 \cdot 43 \times 22 \cdot 4 = 32 \cdot 032$ g

Molecular mass of $O_2 = 32 \cdot 032$ u

Similarly, molecular mass of SO_2 can be calculated.

16. (*i*) Find the volume of the gas at STP which is 700 cm^3

(*ii*) Find the mass of 1 mole *i.e.* 22.4 litres of gas at STP, *i.e.* $\frac{22400}{700} = 32$ g

(*ii*) Calculate the mass of one molecule *i.e.* $\frac{32}{6.022 \times 10^{23}}$ g and divide it by 2 to get the mass of one atom of the gas.

18. 1 mol of $H_2O = 18 \text{ g} = 6 \cdot 022 \times 10^{23}$ molecules

∴ Mass of one molecule of H_2O

$= \frac{18}{6 \cdot 022 \times 10^{23}} g = 2 \cdot 989 \times 10^{-23} g$

0·05 g of $H_2O = \frac{0 \cdot 05}{18}$ mol

$= \frac{0 \cdot 05}{18} \times 6 \cdot 022 \times 10^{23}$ molecules

$= 1 \cdot 673 \times 10^{21}$ molecules

No. of molecules leaving per sec $= \frac{1 \cdot 673 \times 10^{21}}{60 \times 60}$

$= 4 \cdot 647 \times 10^{17}$

19. 1 mol of $K_4[Fe(CN)_6]$ contains C = 6 g atom

0·5 mol of $K_4[Fe(CN)_6]$ will contain C

$= 6 \times 0 \cdot 5$ g atom = 3 g atom $= 3 \times 12$ g = 36 g

20. 1 mole of NaCl = 58·5 g

∴ Cost of NaCl per mole $= \frac{2}{1000} \times 58 \cdot 5$ Rs.

$= 0 \cdot 117$ Re $= 11 \cdot 7\, p = 12\, p.$

1 mole of sugar $(C_{12}H_{22}O_{11}) = 342$ g

∴ Cost of sugar per mole $= \frac{6}{1000} \times 342$ Rs.

= Rs. 2·05 *p*.

21. 100 *g* chlorophyll contains Mg = 2·68 g

∴ 2 g chlorophyll will contain Mg $= \frac{2 \cdot 68}{100} \times 2$

$= 0 \cdot 0536$ g

1 mole of Mg $= 24 \text{ g} = 6 \cdot 022 \times 10^{23}$ atoms

∴ $0 \cdot 0536 \text{ g Mg} = \frac{6 \cdot 022 \times 10^{23}}{24} \times 0 \cdot 0536$ atoms

$= 1 \cdot 345 \times 10^{21}$ atoms

22. Time Required $= \frac{6 \cdot 02 \times 10^{23}}{10^{10}} s = 6 \cdot 02 \times 10^{13} s$

$= \frac{6 \cdot 02 \times 10^{13}}{60 \times 60 \times 24 \times 365}$ years

= 190,893 years approx.

23. $1 \cdot 4 \text{ g } N_2 = \frac{1 \cdot 4}{28}$ mol $= 0 \cdot 05$ mol

$= 0 \cdot 05 \times 6 \cdot 02 \times 10^{23}$ molecules

$= 3 \cdot 01 \times 10^{22}$ molecules

$= 3 \cdot 01 \times 10^{22} \times 14$ electrons

$= 4 \cdot 214 \times 10^{23}$ electrons

1.29. Problems Involving Mole Concept in Solutions

A solution is defined as a homogeneous mixture of two or more chemically non-reacting substances, the relative amounts of which can be varied upto a certain limit.

If a solution consists of only two components, it is called a **binary solution**. The component present in smaller amount is called the **solute** while the other present in larger amount is called the **solvent.**

The **concentration** of a solution can be expressed in a number of ways as follows : –

(1) Strength. ***The strength of a solution is defined as the amount of the solute in grams present per litre of the solution (i.e. g/L or g L^{-1}) or it is defined as the amount of solute in grams present in 100 g of the solution (called percent strength by mass).***

(2) Molarity. ***The molarity of a solution is defined as the number of moles of the solute present per litre of the solution. It is represented by the symbol, M.***

(3) Molality. ***The molality of a solution is defined as the number of moles of the solute dissolved in 1000 g of the solvent. It is represented by the symbol, m.***

(4) Normality. ***The normality of a solution is defined as the number of gram equivalents of the solute present/litre of the solution. It is represented by the symbol, N.***

(5) Mole fraction. ***The mole fraction of any component in the solution is equal to the number of moles of that component divided by the total number of moles of all the components. For a solution containing n_2 moles of the solute dissolved in n_1 moles of the solvent,***

Mole fraction of solute in the solution

$$(x_2) = \frac{n_2}{n_1 + n_2}$$

Mole fraction of solvent in the solution

$$(x_1) = \frac{n_1}{n_1 + n_2}$$

$$x_1 + x_2 = 1$$

The equivalent masses of acids, bases and salts are calculated as follows :–

$$\text{Eq. mass of an acid} = \frac{\text{Mol. mass of the acid}}{\text{Basicity}}$$

$$\text{Eq. mass of a base} = \frac{\text{Mol. mass of the base}}{\text{Acidity}}$$

$$\text{Eq. mass of a salt} = \frac{\text{Mol. mass of the salt}}{\text{Total positive valency of metal atoms}}$$

Basicity *is the number of displaceable H^+ ions from one molecule of the acid (e.g. 1 for HCl, 2 for H_2SO_4, 3 for H_3PO_4 etc.).*

Acidity *is the number of displaceable OH^- ions from one molecule of the base (e.g. 1 for NaOH, 2 for $Ca(OH)_2$ etc.).*

Molarity equation. If a solution having molarity M_1 and volume V_1 is diluted to volume V_2 so that the new molarity is M_2, then as the total number of moles in the solution remains the same, we have

$$\boxed{M_1 \times V_1 = M_2 \times V_2}$$

This equation is called molarity equation.

For a balanced chemical equation involving n_1 moles of reactant 1 and n_2 moles of reactant 2,

$$\boxed{\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}}$$

For exact neutralisation of V_a ml of an acid having molarity M_a and basicity n_a by V_b ml of a base having molarity M_b and acidity n_b,

$$\boxed{n_a M_a V_a = n_b M_b V_b}$$

Normality equation. If a solution having normality N_1 and volume V_1 is diluted to volume V_2 so that the new normality is N_2 or if V_1 cc of solution of substance A and normality N_1 react exactly with V_2 cc of solution of substance B and normality N_2, then in the first case as the number of gram equivalents remains the same and in the second case, substances react in equivalent amounts,

$$\boxed{N_1 \times V_1 = N_2 \times V_2}$$

This equation is called normality equation.

REMEMBER

Normality of a solution = No. of g eq. L^{-1}

= No. of milli eq. mL^{-1}

Molarity of a solution = No. of moles L^{-1}

= No. of milli moles mL^{-1}

EXAMPLE 1. *A solution of oxalic acid, $(COOH)_2 . 2H_2O$ is prepared by dissolving 0·63 g of the acid in 250 cm^3 of the solution. Calculate (a) molarity (b) normality of the solution.*

Solution. **(a) Calculation of molarity.**

Molar mass of oxalic acid,

$$\begin{array}{l} COOH \\ | \\ COOH \end{array} .2H_2O = 126 \text{ g mol}^{-1}$$

$\therefore$ 0.63 g of oxalic acid $= \frac{0.63}{126}$ mole of oxalic acid $= 0.005$ mole of oxalic acid

$$\left[\text{Moles} = \frac{\text{Mass in g}}{\text{Molar mass}}\right]$$

Thus 250 cm^3 of the solution contain oxalic acid = 0.005 mole

$\therefore$ 1000 cm^3 of the solution contain oxalic acid $= \frac{0\cdot005}{250} \times 1000 = 0.02$ mole

i.e. Molarity of the solution = **0.02 M**

(*b*) Calculation of normality

$$\text{Eq. mass of oxalic acid} = \frac{\text{Mol. mass of oxalic acid}}{\text{Basicity}}$$

$$= \frac{126}{2} = 63$$

$\therefore$ 0.63 g of oxalic acid $= \frac{0.63}{63}$ g eq. = 0.01 g eq.

$$\left[\text{Gram equivalents} = \frac{\text{Mass in g}}{\text{Eq. mass}}\right]$$

Thus 250 cm^3 of the solution contain oxalic acid = 0.01 g eq

$\therefore$ 1000 cm^3 of the solution contain oxalic acid $= \frac{0.01}{250} \times 1000 = 0.04$ g eq

i.e. Normality of the solution = **0.04 N**

• EXAMPLE 2. *Commercially available concentrated hydrochloric acid contains 38% HCl by mass.*

(a) What is the molarity of this solution ? The density is 1.19 g cm^{-3}.

(b) What volume of concentrated hydrochloric acid is required to make 1.00 L of 0.10 M HCl ?

(N.C.E.R.T.)

Solution. (*a*) Calculation of molarity. 38% HCl by mass means that 38 g of HCl are present in 100 g of the solution.

$$\text{Volume of 100 g of the solution} = \frac{\text{Mass}}{\text{Density}}$$

$$= \frac{100}{1.19} = 84.03 \text{ cm}^3$$

Molar mass of HCl = 36.5 g mol^{-1}

$\therefore$ 38 g HCl $= \frac{38}{36.5}$ moles = 1.04 moles

Thus 84.03 cm^3 of the solution contain HCl = 1.04 moles

$\therefore$ 1000 cm^3 of the solution contain HCl

$$= \frac{1.04}{84.03} \times 1000$$

$$= 12.38 \text{ moles}$$

i.e. Molarity of the solution = **12.38 M**

(*b*) Calculation of volume of conc. HCl for 1.00 L of 0.10 M HCl.

Applying molarity equation, we have

$$\underset{\text{(conc. HCl)}}{M_1 \times V_1} = \underset{(1\cdot0 \text{ L of } 0\cdot10 \text{ M HCl})}{M_2 \times V_2}$$

$$12\cdot38 \times V_1 = 0\cdot10 \times 1\cdot0$$

or $V_1 = \frac{0\cdot1}{12\cdot38}$ L $= \frac{0\cdot1}{12\cdot38} \times 1000$ cm^3

$$= \mathbf{8\cdot1 \text{ cm}^3}$$

EXAMPLE 3. *(a) A sample of NaOH weighing 0.40 g is dissolved in water and the solution is made to 50.0 cm^3 in volumetric flask. What is the molarity of the resulting solution ?*

(b) How many grams of NaOH should be dissolved to make 100 cm^3 of 0·15 M NaOH solution ?

Solution. (*a*) 0.40 g of NaOH $= \frac{0.40}{40}$ moles = 0.01 mole

Thus 50.0 cm^3 of the solution contain NaOH = 0.01 mole

$\therefore$ 1000 cm^3 of the solution contain NaOH

$$= \frac{0.01}{50} \times 1000 = 0.2 \text{ mole}$$

$\therefore$ Molarity of the solution = **0.2 M**

(*b*) 1000 cm^3 of 0.15 M NaOH contain NaOH = 0.15 mole

$\therefore$ 100 cm^3 of 0.15 M NaOH contain NaOH

$$= \frac{0.15}{1000} \times 100 = 0.015 \text{ mole}$$

$$= 0.015 \times 40 \text{ g} = \mathbf{0.6 \text{ g}}$$

PROBLEMS FOR PRACTICE

1. A solution is prepared by dissolving 18.25 g of NaOH in distilled water to give 200 cm^3 of the solution. Calculate the molarity of the solution. **[Ans. 2.28 M]**

2. How many moles and how many grams of sodium chloride (NaCl) are present in 250 cm^3 of a 0.500 M NaCl solution ? **[Ans. 0.125 mole, 7.312 g]**

3. Concentrated aqueous sulphuric acid is 98% H_2SO_4 by mass and has a density of 1.84 g cm^{-3}. What volume of the concentrated acid is required to make 5.0 litre of 0.500 M H_2SO_4 solution ? **[Ans. 136 cm^3]**

4. How many grams of barium chloride ($BaCl_2$) are needed to prepare 100 cm^3 of 0.250 M $BaCl_2$ solution ? **[Ans. 5.20 g]**

5. How many moles and how many grams of sodium chloride (NaCl) are present in 250 mL of a 0·50 M NaCl solution ? **[Ans. 0·125 mol, 7·32 g]**

6. A sample of $NaNO_3$ weighing 0·38 g is placed in a 50·0 mL measuring flask. The flask is then filled with water upto the mark on the neck. What is the molarity of the solution ? *(N.C.E.R.T.)* **[Ans. 0·090 M]**

7. In a reaction vessel, 0·184 g of NaOH is required to be added for completing the reaction. How many milli litres of 0·150 M NaOH solution should be added for this requirement ? *(N.C.E.R.T.)* **[Ans. 30·7 mL]**

HINTS FOR DIFFICULT PROBLEMS

3. 98% H_2SO_4 by mass means 98 g H_2SO_4 are present in 100 g of the solution = 100/1·84 cm^3

$$= 54\cdot35 \text{ cm}^3$$

Molar mass of H_2SO_4 = 98 g mol^{-1}

Hence 98 g H_2SO_4 = 1 mole

∴ Molarity of the given solution

$$= \frac{1}{54\cdot35} \times 1000 = 18\cdot4 \text{ M}$$

Applying $M_1V_1 = M_2V_2$, $18\cdot4 \times V_1 = 5 \times 0\cdot500$

or $V_1 = 0\cdot136 \text{ L} = 136 \text{ cm}^3$

4. Molecular mass of $BaCl_2$ = 137 + 71 = 208 u.

For 1000 cm^3 of 1 M $BaCl_2$ sol., mass of $BaCl_2$ needed = 208 g

∴ For 100 cm^3 of 0·25 M $BaCl_2$, mass needed

$$= \frac{208}{1000} \times 100 \times 0\cdot25 \text{ g}$$

$$= 5\cdot20 \text{ g}$$

6. Formula mass of $NaNO_3$ = 23 + 14 + 48 = 85

$$0\cdot38 \text{ g } NaNO_3 = \frac{0\cdot38}{85} \text{ mol} = 0\cdot0045 \text{ mol}$$

$$\text{Molarity of the solution} = \frac{0\cdot0045}{50} \times 1000 = 0\cdot090 \text{ M.}$$

7. 1000 mL of 0·150 M NaOH contain NaOH

$$= 0\cdot150 \text{ mol} = 0\cdot150 \times 40 = 6 \text{ g}$$

i.e. 6 g in 0·150 M M NaOH solution ≡ 1000 mL

∴ 0·184 g in 0·150 M NaOH solution

$$\equiv \frac{1000}{6} \times 0\cdot184 = 30\cdot7 \text{ mL.}$$

ADD TO YOUR KNOWLEDGE

1. 1 amu (or 1 u) = $\frac{1}{12}$th of mass of an atom of C − 12 = $\frac{1}{N_0}$ g. That is why 1 amu or 1 u is also called **1 Avogram.**

2. The number of molecules present in 1 cm^3 of an ideal gas at STP is called **Loschmidt number.**

$$\text{Its value} = \frac{6\cdot02 \times 10^{23}}{22400} = 2\cdot69 \times 10^{19}.$$

ADD TO YOUR KNOWLEDGE CONTD.

3. As ionic compounds are not molecular, the term '**Formality**' is used in place of molarity to express the concentration of their solution. *Formality is the number of formula weights present in one litre of the solution.*

$$\text{No. of formula weights} = \frac{\text{Weight of the compound in g}}{\text{Formula wt of the compound}}$$

4. Molarity of a solution does not change with temperature because it involves masses of the solute and solvent which do not change with temperature. Normality and molarity change with temperature because they involve volumes which change with temperature.
5. **Standard solution.** A solution whose normality or molarity is known is called a standard solution.

SECTION—VII

PERCENTAGE COMPOSITION AND MOLECULAR FORMULAE

1.30. Calculation of Percentage Composition from Formula

The percentage of any element or constituent in a compound is the number of parts by mass of that element or constituent present in 100 parts by mass of the compound. It can be calculated by the following two steps :

Step 1. Calculate the molecular mass of the compound from its formula by adding the atomic masses of the elements present.

Step 2. Calculate the percentage of the element or the constituent by applying the following relation :

Percentage of the element or constituent

$$= \frac{\text{No. of parts by mass of the element or constituent}}{\text{Mol. mass of the compound}} \times 100$$

SOLVED EXAMPLES

EXAMPLE 1. *Calculate the percentage composition of the various elements in $MgSO_4$.*

Solution. Mol. mass of $MgSO_4$

$$= 24 + 32 + 4 \times 16 = 120$$

$$\% \text{ of Mg} = \frac{\text{No. of parts by mass of Mg}}{\text{Mol. mass of } MgSO_4} \times 100$$

$$= \frac{24}{120} \times 100 = \mathbf{20\%}$$

$$\% \text{ of S} = \frac{\text{No. of parts by mass of S}}{\text{Mo. mass of } MgSO_4} \times 100$$

$$= \frac{32}{120} \times 100 = \mathbf{26 \cdot 67\%}$$

$$\% \text{ of O} = \frac{\text{No. of parts by mass of O}}{\text{Mol. mass of } MgSO_4} \times 100$$

$$= \frac{64}{120} \times 100 = \mathbf{53 \cdot 33\%.}$$

EXAMPLE 2. *Calculate the percentage of water of crystallisation in the sample of blue vitriol ($CuSO_4 . 5\,H_2O$).*

Solution. Mol. mass of $CuSO_4 . 5\,H_2O$

$$= 63 \cdot 5 + 32 + 4 \times 16 + 5 \times 18 = 249 \cdot 5$$

No. of parts by mass of $H_2O = 5 \times 18 = 90$

$$\therefore \quad \% \text{ of } H_2O = \frac{90}{249 \cdot 5} \times 100$$

$$= \mathbf{36 \cdot 07\%.}$$

EXAMPLE 3. *Calculate the percentage of cation in ammonium dichromate.*

Solution. Molecular formula of ammonium dichromate is $(NH_4)_2Cr_2O_7$

Mol. mass of $(NH_4)_2Cr_2O_7$

$$= 2 \times (14 + 4) + 2 \times 52 + 7 \times 16 = 252$$

No. of parts by mass of cation viz NH_4^+

$$= 2 \times (14 + 4) = 36$$

$$\therefore \quad \% \text{ of } NH_4^+ = \frac{36}{252} \times 100 = \frac{100}{7}$$

$$= \mathbf{14 \cdot 29\%.}$$

PROBLEMS FOR PRACTICE

1. Find the percentage composition of potassium chlorate ($KClO_3$).

[Ans. K = 31·84%, Cl = 28·98%, O = 39·18%]

2. Calculate the percentage of (*i*) SO_4^{2-} (*ii*) H_2O in pure crystals of Mohr salt viz. $FeSO_4 . (NH_4)_2SO_4 . 6 H_2O$.

[Ans. SO_4^{2-} = 48·98%, H_2O = 27·55%]

3. Calculate the percentage of water of crystallisation in the sample of washing soda, $Na_2CO_3 . 10 H_2O$.

[Ans. 62·94%]

4. A sample of clay is found to have the formula $Al_2O_3 . K_2O . 6 SiO_2$. Calculate the percentage of alumina (Al_2O_3), potassium oxide (K_2O) and silica (SiO_2) in the sample.

[Ans. Al_2O_3=18·35%, K_2O=16·90%, SiO_2=64·75%]

5. $Fe_2(SO_4)_3$ is used in water and sewage treatment to aid the removal of suspended impurities. Calculate the mass percentage of iron, sulphur and oxygen in this compound. *(N.C.E.R.T.)*

[Ans. Fe = 28%, S = 24%, O = 48%]

HINTS FOR DIFFICULT PROBLEMS

1. Mol. mass of $KClO_3 = 39 + 35{\cdot}5 + 48 = 122{\cdot}5.$

$$\% \text{ of K} = \frac{39}{122{\cdot}5} \times 100, \% \text{ of Cl} = \frac{35{\cdot}5}{122{\cdot}5} \times 100,$$

$$\% \text{ of O} = \frac{48}{122{\cdot}5} \times 100$$

2. Mol. mass of $FeSO_4 . (NH_4)_2SO_4 . 6 H_2O$

$$= 56 + 32 + 64 + 36 + 32 + 64 + 108 = 392$$

$$\% \text{ of } H_2O = \frac{108}{392} \times 100 = 27{\cdot}55\%$$

$$\% \text{ of } SO_4^{2-} = \frac{2\,(32+64)}{392} \times 100 = 48{\cdot}98\%$$

4. Mol. mass of $Al_2O_3 . K_2O . 6 SiO_2$

$$= (2 \times 27 + 3 \times 16) = (2\times39+16) + 6\,(28+2\times16)$$
$$= 102 + 94 + 360 = 556$$

$$\therefore \% \text{ of } Al_2O_3 = \frac{102}{556} \times 100 = 18{\cdot}35\%$$

$$\% \text{ of } K_2O = \frac{94}{556} \times 100 = 16{\cdot}90\%$$

$$\% \text{ of } SiO_2 = \frac{360}{556} \times 100 = 64{\cdot}75\%$$

5. Mol. mass of $Fe_2(SO_4)_3$

$$= 2 \times 56 + (32 + 64) \times 3 = 400$$

$$\% \text{ of Fe} = \frac{2 \times 56}{400} \times 100 = 28\%$$

$$\% \text{ of S} = \frac{32 \times 3}{400} \times 100 = 24\%$$

$$\% \text{ of O} = \frac{4 \times 16 \times 3}{400} \times 100 = 48\%.$$

1.31. Empirical and Molecular Formulae

We know that the formula of water is H_2O while that of sulphuric acid is H_2SO_4. Similarly, formula of carbon dioxide is CO_2, while that of ammonia is NH_3 and so on. Thus *a formula is a symbolic representation of one molecule of the substance which tells the number and kinds of atoms of various elements present in its molecule.* The question is – How are these formulas determined ? The determination of the formula of a substance involves first the determination of its 'Empirical Formula' and then the 'Molecular Formula'. Let us now explain what we mean by these different types of formulas and the method of their calculation.

Empirical Formula. ***The empirical formula of a compound expresses the simplest whole number ratio of the atoms of the various elements present in one molecule of the compound.***

For example, the empirical formula of benzene is CH, that of hydrogen peroxide is HO and that of glucose is CH_2O. This suggests that in the molecule of benzene, one carbon atom is present for every one hydrogen atom ; in the molecule of hydrogen peroxide, one atom of hydrogen is present for every one oxygen atom, and in the molecule of glucose, one atom of carbon is present for every one atom of oxygen and two atoms of hydrogen. Thus, *the empirical formula of a compound represents only the atomic ratio of the various elements present in its molecule.*

Molecular Formula. ***The molecular formula of a compound represents the true formula of its molecule. It expresses the actual number of atoms of various elements present in one molecule of the compound.***

For example, the molecular formula of benzene is C_6H_6, that of hydrogen peroxide is H_2O_2 and that of glucose is $C_6H_{12}O_6$.

This suggests that one molecule of benzene contains six atoms of carbon and six atoms of hydrogen, one molecule of hydrogen peroxide contains two atoms of hydrogen and two atoms of oxygen and one molecule of glucose contains six atoms of carbon, twelve atoms of hydrogen and six atoms of oxygen.

1.32. Relation Between Empirical and Molecular Formulae

The molecular formula of a compound is a simple whole number multiple of its empirical formula. Expressing mathematically,

$$\boxed{\text{Molecular formula} = n \times \text{Empirical formula}}$$

where n is any integer such as 1, 2, 3 ... etc..

When $n = 1$, Molecular Formula = Empirical Formula

When $n = 2$, Molecular Formula = 2 × Empirical Formula and so on.

The value of 'n' can be obtained from the following relation, $n = \dfrac{\text{Molecular mass}}{\text{Empirical formula mass}}$

The molecular mass of a volatile compound can be determined by Victor Meyer's method or by employing the relation

$$\boxed{\text{Molecular mass} = 2 \times \text{Vapour density}}$$

Empirical formula mass can, however, be obtained from its empirical formula simply by adding the atomic masses of the various atoms present in it. Thus, for glucose (E.F. = CH_2O), the empirical formula mass can be calculated as follows :

Empirical formula mass = At. mass of carbon + 2 × At. mass of hydrogen + At. mass of oxygen $= 12.0\ u + 2 \times 1.0\ u + 16 \cdot 0\ u = 30 \cdot 0\ u.$

1.33. Calculation of The Empirical Formula

The empirical formula of a chemical compound can be deduced from a knowledge of the

(*a*) *percentage composition of different elements, and* (*b*) *atomic masses of the elements.*

The following steps are involved in the calculation of the emipirical formula.

Step 1. *To calculate the relative number of atoms or atomic ratio.* Divide the percentage of each element by its atomic mass. This gives the relative number of atoms or the atomic ratio of the various elements present in one molecule of the compound.

$$\text{Atomic ratio} = \frac{\text{Percentage of an element}}{\text{At. mass of the same element}}$$

Step 2. *To calculate the simplest atomic ratio.* Divide the atomic ratio obtained in step 1 by the smallest quotient or the least value from amongst the values obtained for each element. This gives the simplest atomic ratio.

Step 3. *To calculate the simplest whole number ratio.* The simplest atomic ratios as calculated in step 2 are generally whole numbers. If not, then –

(*a*) raise the values to the nearest whole number, or (*b*) multiply all the simplest atomic ratios by a suitable integer.

Step 4. *To write the empirical formula.* Write the symbols of the various elements side by side. Now insert the numerical value of the simplest whole number ratio of each element as obtained in step 3 at the lower right hand corner of each symbol. This gives the *empirical formula* of the compound.

PROBLEMS ON THE CALCULATION OF EMPIRICAL FORMULAS

EXAMPLE 1. *An inorganic salt gave the following percentage composition – Na = 29·11, S = 40·51 and O = 30·38*

Calculate the empirical formula of the salt.

Solution. *Calculation of empirical formula*

Element	Symbol	Percentage of elements	At. mass of elements	Relative no. of atoms $= \frac{\text{Percentage}}{\text{At. mass}}$	Simplest atomic ratio	Simplest whole no. atomic ratio
Sodium	Na	29·11	23	$\frac{29\cdot11}{23} = 1\cdot266$	$\frac{1\cdot266}{1\cdot266} = 1$	2
Sulphur	S	40·51	32	$\frac{40\cdot51}{32} = 1\cdot266$	$\frac{1\cdot266}{1\cdot266} = 1$	2
Oxygen	O	30·38	16	$\frac{30\cdot38}{16} = 1\cdot897$	$\frac{1\cdot89}{1\cdot266} = 1\cdot5$	3

Thus the Empirical Formula is $Na_2S_2O_3$.

EXAMPLE 2. *2·38 g of uranium was heated strongly in a current of air. The resulting oxide weighed 2·806 g. Determine the empirical formula of the oxide. (At. mass U = 238 ; O = 16).*

Solution. **Step 1.** *To calculate the percentage of uranium and oxygen in the oxide.*

2·806 g of the oxide contain uranium = 2·38 g.

$\therefore$ Percentage of uranium $= \frac{2\cdot38}{2\cdot806} \times 100 = 84\cdot82$

Hence, the percentage of oxygen in the oxide = $100\cdot00 - 84\cdot82 = 15\cdot18$.

Step 2. *To calculate the empirical formula*

Element	Symbol	Percentage of elements	At. mass of elements	Relative no. of atoms $= \frac{\text{Percentage}}{\text{At. mass}}$	Simplest atomic ratio	Simplest whole no. atomic ratio
Uranium	U	84·82	238	$\frac{84\cdot82}{238} = 0\cdot3562$	$\frac{0\cdot3562}{0\cdot3562} = 1$	3
Oxygen	O	15·18	16	$\frac{15\cdot18}{16} = 0\cdot94875$	$\frac{0\cdot94875}{0\cdot3562} = 2.666$	8

Hence the empirical formula of the oxide is U_3O_8.

EXAMPLE 3. *A crystalline salt on being rendered anhydrous loses 45·6% of its weight. The percentage composition of the anhydrous salt is*

Aluminium = 10·50% ; Potassium = 15·1% ; Sulphur = 24·96% ; Oxygen = 49·92%.

Find the simplest formula of the anhydrous and crystalline salt.

Solution. **Step 1.** *To calculate the empirical formula of the anhydrous salt*

Element	Symbol	Percentage of elements	At. mass of elements	Relative no. of atoms $= \frac{\text{Percentage}}{\text{At. mass}}$	Simplest atomic ratio	Simplest whole no. atomic ratio
Potassium	K	15·1	39	$\frac{15\cdot10}{39} = 0\cdot39$	$\frac{0\cdot39}{0\cdot39} - 1$	1
Aluminium	Al	10·50	27	$\frac{10\cdot50}{27} = 0\cdot39$	$\frac{0\cdot39}{0\cdot39} = 1$	1
Sulphur	S	24·96	32	$\frac{24\cdot96}{32} = 0\cdot78$	$\frac{0\cdot78}{6\cdot39} = 2$	2
Oxygen	O	49·92	16	$\frac{49\cdot92}{16} = 3\cdot12$	$\frac{3\cdot12}{0\cdot39} = 8$	8

Thus the empirical formula of the anhydrous salt is K Al S_2 O_8

Step 2. *To calculate the empirical formula mass of the anhydrous salt.*

Empirical formula mass of the anhydrous salt ($KAlS_2O_8$)

$= 1 \times 39 \cdot 0 + 1 \times 27.0 + 2 \times 32.0 + 8 \times 16.0$

$= 258.0$ u.

Step 3. *To calculate the empirical formula mass of the hydrated salt.*

Let the empirical formula mass of the hydrated salt $= 100 \cdot 0$ u.

Loss of weight due to dehydration $= 45 \cdot 6\%$

$\therefore$ Empirical formula mass of the anhydrous salt $= 100 - 45 \cdot 6 = 54 \cdot 4$ u.

Now, if the empirical formula mass of the anhydrous salt is $54 \cdot 4$ that of hydrated is $= 100$

$\therefore$ If the empirical formula mass of the anhydrous salt is 258, that of hydrated is

$$= \frac{100}{54 \cdot 4} \times 258 = 474 \cdot 3 \text{ u.}$$

Step 4. *To calculate the number of molecules of water in the hydrated salt.*

Total loss in mass due to dehydration

$= 474 \cdot 3 - 258 \cdot 0 = 216 \cdot 3$ u

Loss in mass due to one molecule of water $= 18 \cdot 0$ u

$\therefore$ No. of molecules of water in the hydrated sample $= \dfrac{216 \cdot 3}{18} = 12$

Step 5. *To calculate the empirical formula of the hydrated salt.*

Empirical formula of the anhydrous salt $= KAlS_2O_8$

No. of molecules of water of crystallization $= 12$

$\therefore$ Empirical formula of the hydrated salt $= KAlS_2O_8 . 12H_2O$

EXAMPLE 4. *The percentage composition of ferrous ammonium sulphate is $14 \cdot 32\%$ Fe^{2+}; $9 \cdot 20\%$ NH_4^+p; $49 \cdot 0\%$ SO_4^{2-} and $27 \cdot 57\%$ H_2O. What is the empirical formula of the compound ?*

Solution. *Calculation of empirical formula.*

Formula of the constituent	Percentage	Mol. mass/At. mass of constituents	Relative no. of constituents $= \dfrac{\text{Percentage}}{\text{Mol. mass}}$	Simplest ratio of constituents
Fe^{2+}	$14 \cdot 32$	56	$\dfrac{14 \cdot 32}{56} = 0 \cdot 2557$	$\dfrac{0 \cdot 2557}{0 \cdot 2557} = 1$
NH_4^+	$9 \cdot 20$	18	$\dfrac{9 \cdot 20}{18} = 0 \cdot 5111$	$\dfrac{0 \cdot 5111}{0 \cdot 2557} = 2$
SO_4^{2-}	$49 \cdot 0$	96	$\dfrac{49 \cdot 0}{96} = 0 \cdot 5104$	$\dfrac{0 \cdot 5104}{0 \cdot 2557} = 2$
H_2O	$27 \cdot 57$	18	$\dfrac{27 \cdot 57}{18} = 1 \cdot 532$	$\dfrac{1 \cdot 532}{0 \cdot 2557} = 6$

Hence the empirical formula of the compound is

$(Fe^{2+})(NH_4^+)_2(SO_4^{2-})_2(H_2O)_6$ or $Fe(NH_4)_2(SO_4)_2 \cdot 6H_2O$.

PROBLEMS FOR PRACTICE

1. An inorganic salt on analysis gave the following percentage composition :

   Pb $= 62 \cdot 6$, N $= 8 \cdot 4$, O $= 29$

   What is empirical formula of the compound ? Also name the compound. (At. mass Pb = 207, N = 14, O = 16). **[Ans. $PbN_2O_6.Pb(NO_3)_2$, Lead nitrate]**

2. An oxide of nitrogen gave the following percentage composition :

   N $= 25 \cdot 94$

   and O $= 74 \cdot 06$

   Calculate the empirical formula of the compound. **[Ans. N_2O_5]**

3. A sample of salt has the following percentage composition :

$Fe = 36{\cdot}76$; $S = 21{\cdot}11$ and $O = 42{\cdot}14$

Calculate the empirical formula of the compound. (At. mass Fe = 56, S = 32 and O = 16)

[Ans. $FeSO_4$]

4. A salt containing water of crystallization gave the following percentage composition :

$Mg = 9{\cdot}76$; $S = 13{\cdot}01$;
$O = 26{\cdot}01$ and $H_2O = 51{\cdot}22$

Calculate the simplest formula. (At. mass of Mg=24) [Ans. $MgSO_4 . 7H_2O$]

5. Calculate the empirical formula of gold chloride which contains 35·1% of chlorine. At. mass of Au = 197. [Ans. $AuCl_3$]

6. Calculate the empirical formula of a mineral having the following composition :

$CaO = 48{\cdot}0\%$; $P_2O_5 = 41{\cdot}3\%$; $CaCl_2 = 10{\cdot}7\%$

[Ans. $9CaO. 3P_2O_5 . CaCl_2$]

HINTS FOR DIFFICULT PROBLEMS

4.

Constituent	*%age*	*Atomic mass/Mol. mass*	*Relative no. of constituents = % / Mol. mass*	*Simplest Ratio*
Mg	9·76	24	9·76/24 = 0·407	1
S	13·01	32	13·01/32 = 0·407	1
O	26·01	16	26·01/16 = 1·626	4
H_2O	51·22	18	51·22/18 = 2·846	7

∴ Empirical Formula = $MgSO_4 . 7\,H_2O$.

6.

Constituent	*%age*	*Mol. mass*	*Relative no. of const. = % / Mol. mass*	*Simplest Ratio*
CaO	48·0	56	48·0/56 = 0·857	9
P_2O_5	41·3	142	41·3/142 = 0·291	3
$CaCl_2$	10·7	111	10·7/111 = 0·096	1

∴ Empirical Formula = $9\,CaO . 3\,P_2O_5 . CaCl_2$.

1.34. Calculation of Molecular Formula

The molecular formula of a compound can be deduced from its

(1) *Empirical formula* and

(2) *Molecular mass*. The determination of molecular formula involves the following steps :

(*a*) *Calculation of the empirical formula from the percentage composition.*

(*b*) *Calculation of empirical formula mass by adding the atomic masses of all the atoms present in the empirical formula.*

(*c*) *Determination of the molecular mass of the compound from the given data.*

(*d*) *Determination of the value of 'n' by using the relation,*

$$n = \frac{\textit{Molecular Mass}}{\textit{Empirical Formula Mass}}$$

(*e*) *Determination of the molecular formula by using the relation*

Molecular Formula = n × Empirical Formula.

PROBLEMS ON THE CALCULATION OF MOLECULAR FORMULAS

EXAMPLE 1. *A compound containing sodium, sulphur, hydrogen and oxygen gave the following results on analysis :—*

Na = 14·28%, S = 9·92%, H = 6·20%

Calculate the molecular formula of the anhydrous compound. If all the atoms of hydrogen in the compound are present in combination with oxygen as water of crystallization, what is the structure of the crystalline salt ? The molecular mass of the crystalline salt is 322.

Solution. Step 1. *To calculate the percentage of oxygen.* The given compound contains oxygen but its percentage is not given. This can, however, be calculated by subtracting the sum of percentages of Na, S and H from 100 as shown below :

Sum of percentage composition of Na, S and H = $14 \cdot 28 + 9 \cdot 92 + 6 \cdot 20 = 30 \cdot 40$

$\therefore$ Percentage of oxygen = $100 \cdot 00 - 30 \cdot 40 = 69 \cdot 60$

Step 2. *To calculate the empirical formula.*

Element	Symbol	Percentage of element	At. mass of element	Relative no. of atoms. $= \frac{\text{Percentage}}{\text{At. mass}}$	Simplest atomic ratio	Simplest whole no. atomic ratio
Sodium	Na	14·28	23	$\frac{14 \cdot 28}{23} = 0 \cdot 62$	$\frac{0 \cdot 62}{0 \cdot 31} = 2$	2
Sulphur	S	9·92	32	$\frac{9 \cdot 92}{32} = 0 \cdot 31$	$\frac{0 \cdot 31}{0 \cdot 31} = 1$	1
Hydrogen	H	6·20	1	$\frac{6 \cdot 2}{1} = 6 \cdot 20$	$\frac{6 \cdot 20}{0 \cdot 31} = 20$	20
Oxygen	O	69·60	16	$\frac{69 \cdot 60}{16} = 4 \cdot 35$	$\frac{4 \cdot 35}{0 \cdot 31} = 14$	14

Hence the empirical formula of the compound is $Na_2SH_{20}O_{14}$

Step 3. *To calculate the empirical formula mass.* Empirical formula mass of the compound ($Na_2SH_{20}O_{14}$)

$= 2 \times 23 \cdot 0 + 1 \times 32 \cdot 0 + 20 \times 1 \cdot 0 + 14 \times 16 \cdot 0$

$= 322$

Step 4. *To calculate the value of 'n'*

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{322}{322} = 1$$

Step 5. *To calculate the molecular formula of the compound.*

Molecular formula = $n \times$ Empirical formula

$= 1 \times Na_2SH_{20}O_{14} = Na_2SH_{20}O_{14}$

Step 6. *To calculate the number of molecules of water of crystallization.* Since all the H-atoms are present in the form of H_2O, 20 atoms of hydrogen would combine with 10 oxygen atoms to produce 10 molecules of H_2O. Thus, the number of molecules of water of crystallization present in the salt $Na_2SH_{20}O_{14}$ is 10.

Step 7. *To determine the structure of the crystalline salt.* Out of the 14 oxygen atoms, 10 are present in form of H_2O. The remaining 4 must be a part of the salt consisting of Na and S. A plausible formula for this salt is Na_2SO_4 and that of crystalline salt will be $Na_2SO_4 \cdot 10H_2O$

EXAMPLE 2. *An organic substance containing carbon, hydrogen and oxygen gave the following percentage composition.*

C = 40·687% ; H = 5·085% and O = 54·228%

The vapour density of the compound is 59. Calculate the molecular formula of the compound.

Solution. Step 1. *To calculate the empirical formula of the compound.*

Element	Symbol	Percentage of element	At. mass of element	Relative no. of atoms. $= \frac{\text{Percentage}}{\text{At. mass}}$	Simplest atomic ratio	Simplest whole no.atomic ratio
Carbon	C	40·687	12	$\frac{40 \cdot 687}{12} = 3 \cdot 390$	$\frac{3 \cdot 390}{3 \cdot 389} = 1$	2
Hydrogen	H	5·085	1	$\frac{5 \cdot 085}{1} = 5 \cdot 085$	$\frac{5 \cdot 085}{3 \cdot 389} = 1 \cdot 5$	3
Oxygen	O	54·228	16	$\frac{54 \cdot 228}{16} = 3 \cdot 389$	$\frac{3 \cdot 389}{3 \cdot 389} = 1$	2

$\therefore$ Empirical Formula is $C_2H_3O_2$

Step 2. *To calculate the empirical formula mass.*

The empirical formula of the compound is $C_2H_3O_2$.

$\therefore$ Empirical formula mass
$= (2 \times 12) + (3 \times 1) + (2 \times 16) = 59.$

Step 3. *To calculate the molecular mass of the salt.*

The vapour density of the compound = 59. *(Given)*

Using the relation between vapour density and molecular mass –

Molecular mass = 2 × vapour density
$= 2 \times 59 = 118.$

Step 4. *To calculate the value of 'n'*

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{118}{59} = 2$$

Step 5. *To calculate the molecular formula of the salt.*

Molecular formula = n × (Empirical formula)
$= 2 \times C_2H_3O_2 = C_4H_6O_4$

Thus the *molecular formula* is $C_4H_6O_4$.

PROBLEMS FOR PRACTICE

1. A crystalline salt when heated becomes anhydrous and loses 51·2% of its weight. The anhydrous salt on analysis gave the following percentage composition :
   Mg = 20·0% ; S = 26·66% and O = 53·33%
   Calculate the molecular formula of the anhydrous salt and the crystalline salt. Molecular mass of the anhydrous salt is 120. [Ans. $MgSO_4 . 7H_2O$]
2. A compound containing carbon, hydrogen and oxygen gave the following analytical data :
   C = 40·0% and H = 6·67%
   Calculate the molecular formula of the compound if its molecular mass is 180. [Ans. $C_6H_{12}O_6$]
3. On analysis, a substance was found to have the following percentage composition :
   K = 31·84, Cl = 28·98 and O = 39·18
   Calculate its molecular formula if its molecular mass is 122·5. [Ans. $KClO_3$]
4. An organic liquid having carbon, hydrogen, nitrogen and oxygen was found to contain
   C=41·37% ; H =5·75% ; N = 16·09% and the rest oxygen. Calculate the molecular formula of the liquid if its V.D. is 43.3. [Ans. $C_3H_5NO_2$]
5. A chemical compound is found to have the following composition :
   C = 19·57% ; Fe = 15·2% ; N = 22·83 % ;
   K = 42·39%
   Calculate the empirical formula of the compound. What will be its molecular formula if the molecular mass of the compound is 368 ? Name the compound and describe the action of hydrogen peroxide on it. [Ans. $K_4Fe(CN)_6$]
6. Butyric acid contains only C, H and O. A 4.24 mg sample of butyric acid is completely burned. It gives 8.45 mg of CO_2 and 3.46 mg of H_2O. The molecular mass of butyric acid was determined by experiment to be 88 amu. What is molecular formula ? *(N.C.E.R.T.)* [Ans. $C_4H_8O_2$]
7. A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3·38 *g* carbon dioxide, 0·690 *g* of water and no other products. A volume of 10·0 L (measured at S.T.P.) of this welding gas is found to weigh 11·6 g. Calculate (*i*) empirical formula (*ii*) molar mass of the gas, and (*iii*) molecular formula. *(N.C.E.R.T.)* [Ans. (*i*) CH (*ii*) 26 (*iii*) C_2H_2]

HINTS FOR DIFFICULT PROBLEMS

1. Calculate the empirical formula of the anhydrous salt. It comes out to be $MgSO_4$. E.F. mass = 120. Mol. mass = 120. Hence molecular formula = $MgSO_4$.

   As crystalline salt on becoming anhydrous loses 51·2% by mass, this means

   48·8 g of anhydrous salt contains H_2O = 51·2 g

   $\therefore$ 120 g of anhydrous salt contains H_2O
   $= \frac{51 \cdot 2}{48 \cdot 8} \times 120 \text{ g} = 126 \text{ g} = \frac{126}{18}$ molecules
   = 7 molecules.

   Hence mol. formula of crystalline salt
   $= MgSO_4 . 7H_2O$.

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5. The molecular formula $K_4FeC_6N_6$ suggests that it is *Potassium ferrocyanide*. It is oxidised by H_2O_2 to potassium ferricyanide.

$$H_2O_2 \longrightarrow H_2O + O$$

$$2K_4[Fe(CN)_6] + O + H_2O \longrightarrow 2K_3[Fe(CN)_6] + 2KOH$$

6. 1 mole CO_2 contains 1 g atom of C

i.e. 44 g CO_2 contains C = 12 g

$\therefore$ 8·45 mg CO_2 will contain C $= \frac{12}{44} \times 8{\cdot}45$ mg

This is present in 4·24 mg of the compound.

$\therefore$ % of C in the compound $= \frac{12}{44} \times \frac{8{\cdot}45}{4{\cdot}24} \times 100$

$= 54{\cdot}4\%$

or use the formula directly

$$\% \text{ of C} = \frac{12}{44} \times \frac{\text{Mass of } CO_2}{\text{Mass of compound}} \times 100$$

Similarly, % of H

$$= \frac{2}{18} \times \frac{\text{Mass of } H_2O}{\text{Mass of compound}} \times 100$$

$$= \frac{2}{18} \times \frac{3{\cdot}46}{4{\cdot}24} \times 100 = 9{\cdot}1\%$$

$\therefore$ % of O = 100—(54·4 + 9·1) = 36·5%

Calculate E.F. If comes out to be C_2H_4O.

E.F. mass = 44 u, Mol mass = 88 u.

Hence n = Mol. mass/E.F mass = 2

$\therefore$ Mol. formula = 2 × E.F. = $C_4H_8O_2$

7. Amount of carbon in 3·38 g CO_2

$$= \frac{12}{44} \times 3{\cdot}38\, g = 0{\cdot}9218\, g$$

Amount of hydrogen in 0·690 g H_2O

$$= \frac{2}{18} \times 0{\cdot}690\, g = 0{\cdot}0767\, g$$

As compound contains only C and H, therefore, total mass of the compound

$= 0{\cdot}9218 + 0{\cdot}0767\, g = 0{\cdot}9985\, g$

% of C in the compound

$$= \frac{0{\cdot}9218}{0{\cdot}9985} \times 100 = 92{\cdot}32$$

$$\% \text{ of H in the compound} = \frac{0{\cdot}0767}{0{\cdot}9985} \times 100 = 7{\cdot}68$$

Now calculate empirical formula. It comes out to be CH.

$$\text{Molar mass} = \frac{11{\cdot}6}{10{\cdot}0} \times 22{\cdot}4 = 26\, g\, mol^{-1}$$

E.F. mass = 13 $\therefore$ n = 2. Hence M.F. = C_2H_2

ADD TO YOUR KNOWLEDGE

Usually no direct method is used for finding the percentage of oxygen in a compound. It is calculated by finding the percentages of all elements except oxygen. Then %age of oxygen = 100 — (Sum of %ages of all other elements present in the compound).

SECTION—VIII
STOICHIOMETRY OF CHEMICAL REACTIONS

1.35. General Information

One of the most important aspects of a chemical equation is that when it is written in the balanced form, it gives quantitative relationships between the various reactants and products in terms of moles, masses, molecules and volumes. This is called *stoichiometry* (Greek word, meaning 'to measure an element'). For example, a balanced chemical equation alongwith the quantitative information conveyed by it is given below :

$$\underset{\substack{1 \text{ Mole} \\ 40 + 12 + 3 \times 16 \\ = 100 \text{ g}}}{CaCO_3} + \underset{\substack{2 \text{ Moles} \\ 2(1 + 35.5) \\ = 73 \text{ g}}}{2HCl} \longrightarrow \underset{\substack{1 \text{ Mole} \\ 40 + 2 \times 35.5 \\ = 111 \text{ g}}}{CaCl_2} + \underset{\substack{1 \text{ Mole} \\ 2 \times 1 + 16 \\ = 18 \text{ g}}}{H_2O} + \underset{\substack{1 \text{ Mole} \\ 12 + 2 \times 16 \\ = 44 \text{ g} \\ \text{or 22.4 litres at STP}}}{CO_2}$$

Thus, (*i*) 1 mole of calcium carbonate reacts with 2 moles of hydrochloric acid to give 1 mole of calcium chloride, 1 mole of water and 1 mole of carbon dioxide.

(*ii*) 100 g of calcium carbonate react with 73 g hydrochloric acid to give 111 g of calcium chloride, 18 g of water and 44 g (or 22·4 litres at STP) of carbon dioxide.

The quantitative information conveyed by a chemical equation helps in a number of calculations. The problems involving these calculations may be classified into the following different types :

(1) *Mass-Mass Relationships i.e.* mass of one of the reactants or products is given and the mass of some other reactant or product is to be calculated.

(2) *Mass-Volume Relationship i.e.* mass/volume of one of the reactants or products is given and the volume/mass of the other is to be calculated.

(3) *Volume-Volume Relationship i.e.* volume of one of the reactants or the products is given and the volume of the other is to be calculated.

The general method of calculations for all the problems of the above types consists of the following steps :

(*i*) Write down the balanced chemical equation.

(*ii*) Write the relative number of moles or the relative masses (gram atomic or molecular masses) of the reactants and the products below their formulae.

(*iii*) In case of a gaseous substance, write down 22·4 litres at STP below the formula in place of 1 mole.

(*iv*) Apply *unitary method* to make the required calculations.

Now we shall take up a few solved examples to illustrate the different types of problems.

TYPE I. Involving Mass-Mass Relationship

EXAMPLE 1. *Calculate the mass of iron which will be converted into its oxide (Fe_3O_4) by the action of 18 g of steam on it.*

Solution. The chemical equation representing the reaction is

$$\underset{\substack{3\times 56\\ =168\text{ g}}}{3Fe} + \underset{\substack{4\times 18\\ =72\text{ g}}}{4H_2O} \longrightarrow Fe_3O_4 + 4H_2$$

Now 72 g of steam react with 168 g of iron

$\therefore$ 18 g of steam will react with

$$\frac{168}{72}\times 18 = 42 \text{ g of iron}$$

Thus the mass of iron required = **42 g.**

EXAMPLE 2. *What mass of slaked lime would be required to decompose completely 4 grams of ammonium chloride and what would be the mass of each product ?*

Solution. The equation representing the decomposition of NH_4Cl by slaked lime, *i.e.*, $Ca(OH)_2$ is,

$$\underset{\substack{40+2(1+16)\\ =74\text{ g}}}{Ca(OH)_2} + \underset{\substack{2(14+4+35\cdot5)\\ =107\text{ g}}}{2NH_4Cl} \longrightarrow \underset{\substack{40+2\times35\cdot5\\ =111\text{ g}}}{CaCl_2} + \underset{\substack{2(14+3\times1)\\ =34\text{ g}}}{2NH_3} + \underset{\substack{2(2\times1+16)\\ =36\text{ g}}}{2H_2O}$$

(*i*) *To calculate the mass of $Ca(OH)_2$ required to decompose 4g of NH_4Cl.*

From the above equation,

107 g of NH_4Cl are decomposed by 74 g of $Ca(OH)_2$

$\therefore$ 4 g of NH_4Cl will be decomposed by

$$\frac{74}{107}\times 4 = 2\cdot766 \text{ g of } Ca(OH)_2$$

Thus the mass of slaked lime required = **2·766 g.**

(*ii*) *To calculate the mass of $CaCl_2$ formed.*

107 g of NH_4Cl when reacted with $Ca(OH)_2$ produce 111 g of $CaCl_2$.

$\therefore$ 4 g of NH_4Cl when reacted with $Ca(OH)_2$ will produce

$$\frac{111}{107}\times 4 = 4.15 \text{ g of } CaCl_2$$

Hence the mass of $CaCl_2$ produced

= **4·15 g.**

(*iii*) *To calculate the mass of NH_3 produced.*

107 g of NH_4Cl when reacted with $Ca(OH)_2$ give 34 g of NH_3.

$\therefore$ 4 g of NH_4Cl when reacted with $Ca(OH)_2$ will produce

$$\frac{34}{107}\times 4 = 1\cdot271 \text{ g of } NH_3$$

Hence, the mass of NH_3 produced

= **1·271 g.**

(*iv*) *To calculate the mass of H_2O formed*

107 g of NH_4Cl when reacted with $Ca(OH)_2$ yield 36 g of H_2O

$\therefore$ 4 g of NH_4Cl when reacted with $Ca(OH)_2$ will yield $= \frac{36}{107} \times 4 = 1{\cdot}3458$ g of H_2O

So the mass of H_2O formed = **1·3458 g.**

EXAMPLE 3. *1·5 g of an impure sample of sodium sulphate dissolved in water was treated with excess of barium chloride solution when 1·74 g of $BaSO_4$ were obtained as dry precipitate. Calculate the percentage purity of the sample.*

Solution.

Given that 1·5 g of impure Na_2SO_4 $\xrightarrow[\text{BaCl}_2\text{ gave}]{\text{Treated with}}$ 1·74 g of $BaSO_4$

The chemical equation representing the reaction is

$$\underset{\substack{2\times23+32+4\times16\\=142\text{ g}}}{Na_2SO_4} + BaCl_2 \longrightarrow \underset{\substack{137+32+4\times16\\=233\text{ g}}}{BaSO_4}$$

Step 1. *To calculate the mass of Na_2SO_4 from 1·74 g of $BaSO_4$.* From the chemical equation –

233 g of $BaSO_4$ are produced from 142 g of Na_2SO_4

$\therefore$ 1·74 g of it would be obtained from $\frac{142}{233} \times 1{\cdot}74 = 1{\cdot}06$ g

The mass of pure Na_2SO_4 present in 1·5 g of impure sample = 1·06 g.

Step 2. *To calculate the percentage purity of impure sample.*

1·5 g of impure sample contains 1·06 g of pure Na_2SO_4

$\therefore$ 100 g of the impure sample will contain

$= \frac{1{\cdot}06}{1{\cdot}5} \times 100 = 70{\cdot}67$ g of pure Na_2SO_4

Percentage purity of impure sample = **70·67**

EXAMPLE 4. *Current market prices of Al, Zn and Fe scraps per kg are Rs. 20, Rs. 16 and Rs. 3 respectively. If H_2 is to be prepared by the reaction of one of these metals with H_2SO_4, which would be the cheapest metal to use ? Which would be most expensive ?*

Solution. The various chemical reactions involved are given below :

(*i*) $\underset{2\times27=54\text{ g}}{2Al} + 3H_2SO_4 \longrightarrow Al_2(SO_4)_3 + \underset{3\times2=6\text{ g}}{3H_2}$

(*ii*) $\underset{65\text{ g}}{Zn} + H_2SO \longrightarrow ZnSO_4 + \underset{2\text{g}}{H_2}$

(*iii*) $\underset{56\text{ g}}{Fe} + H_2SO_4 \longrightarrow FeSO_4 + \underset{2\text{ g}}{H_2}$

Let us suppose that the amount of hydrogen to be prepared = 100 g.

Step 1. *To calculate the cost of preparation of 100 g of H_2 from Al*

6 g of H_2 is prepared from Al = 54 g

$\therefore$ 100 g of H_2 will be obtained from Al

$= \frac{54}{6} \times 100 = 900$ g

Cost of 1000 g of Al = Rs. 20

$\therefore$ Cost of 900 g of Al $= \frac{20}{1000} \times 900$

= Rs. 18

Step 2. *To calculate the cost of preparation of 100 g of H_2 from Zn.*

2 g of H_2 is produced from Zn = 65 g

$\therefore$ 100 g of H_2 will be obtained from Zn

$= \frac{65}{2} \times 100 = 3250$ g

Cost of 1000 g of Zn = Rs. 16

$\therefore$ Cost of 3250 g of Zn $= \frac{16}{1000} \times 3250 =$ Rs. 52.

Step 3. *To calculate the cost of preparation of 100 g of H_2 from Fe.*

2 g of H_2 is produced from Fe = 56 g

$\therefore$ 100 g of H_2 will be obtained from Fe

$= \frac{56}{2} \times 100 = 2800$ g

Cost of 1000 g of Fe = Rs. 3

$\therefore$ Cost of 2800 g of Fe $= \frac{3}{1000} \times 2800$

= Rs. 8.40.

Thus Fe is the cheapest and Zn is the most expensive metal to use for the preparation of H_2.

• **EXAMPLE 5.** *In order to find the strength of a sample of sulphuric acid, 10 g were diluted with water and a piece of marble weighing 7 g placed in it. When all action had ceased, the marble was removed, washed, dried and was found to weigh 2.2 g. What was the percentage strength of sulphuric acid ?*

Solution. Mass of marble taken $= 7 \cdot 0$ g.

Mass of marble left unused $= 2 \cdot 2$ g

$\therefore$ Mass of marble reacted $= 7 \cdot 0 - 2 \cdot 2 = 4 \cdot 8$ g.

The chemical equation involved in the above problem is

$$\underset{\substack{40+12+16\times 3\\=100\text{ g}}}{CaCO_3} + \underset{\substack{2+32+4\times 16\\=98\text{ g}}}{H_2SO_4} \longrightarrow CaSO_4 + H_2O + CO_2$$

Step 1. *To calculate the mass of pure H_2SO_4 required to react with 4·8 g of marble.*

100 g of marble react with $H_2SO_4 = 98$ g

$\therefore$ 4·8 g of marble will react with H_2SO_4

$$= \frac{98}{100} \times 4 \cdot 8 = 4 \cdot 704 \text{ g.}$$

Step 2. *To calculate the strength of sulphuric acid.*

10 g of dil. H_2SO_4 contain pure H_2SO_4 $= 4 \cdot 704$ g.

$\therefore$ 100 g of dil. H_2SO_4 will contain pure

$$H_2SO_4 = \frac{4 \cdot 704}{10} \times 100 = 47 \cdot 04$$

Thus the percentage strength of sulphuric acid $=$ **47·04**.

• **EXAMPLE 6.** *1·84 g of a mixture of $CaCO_3$ and $MgCO_3$ is strongly heated till no further loss of mass takes place. The residue weighs 0·96 g. Calculate the percentage composition of the mixture.*

Solution. Let the mass of $CaCO_3$ in the mixture be x g.

$\therefore$ The mass of $MgCO_3$ in the mixture will be $(1 \cdot 84 - x)$ g.

Step 1. *To calculate the mass of CaO residue from x g calcium carbonate.*

$$\underset{100\text{ g}}{CaCO_3} \longrightarrow \underset{56\text{ g}}{CaO} + CO_2$$

$\therefore$ 100 g of $CaCO_3$ upon decomposition give a residue of 56 g of CaO

$\therefore$ x g of $CaCO_3$ will give

$$\frac{56 \times x}{100} = 0 \cdot 56 \times x \text{ g of CaO}$$

Thus the mass of CaO formed $= 0 \cdot 56\,x$ g

Step 2. *To calculate the mass of MgO residue from $(1 \cdot 84 - x)$ g of magnesium carbonate.*

Using the equation

$$\underset{84\text{ g}}{MgCO_3} \longrightarrow \underset{40\text{ g}}{MgO} + CO_2$$

84 g of $MgCO_3$ upon decomposition yield a residue of 40 g of MgO

$\therefore$ $(1 \cdot 84 - x)$ g of $MgCO_3$ will yield

$$= \frac{40 \times (1 \cdot 84 - x)}{84} \text{ g of MgO}$$

Thus the mass of MgO formed

$$= \frac{40}{84}(1 \cdot 84 - x) \text{ g.}$$

Step 3. *To calculate the masses of $CaCO_3$ and $MgCO_3$ in the mixture.*

Total mass of the residue $= 0 \cdot 96$ g.

Equating the total masses of CaO and MgO from Step 1 and Step 2, we have –

$$0 \cdot 56\,x + \frac{40}{84}(1 \cdot 84 - x) = 0 \cdot 96$$

$$0 \cdot 56\,x \times 84 + 40 \times 1 \cdot 84 - 40\,x = 0 \cdot 96 \times 84$$

or $\quad 47 \cdot 04\,x - 40\,x = 80 \cdot 64 - 73 \cdot 60$

or $\quad 7 \cdot 04\,x = 7 \cdot 04 \quad$ or $\quad x = 1$

Thus the mass of $CaCO_3$ in the mixture $= 1 \cdot 00$ g and the mass of $MgCO_3$ in the mixture $= 1 \cdot 84 - 1 \cdot 00 =$ **0·84** g.

$$\% \text{ of } CaO_3 \text{ in the mixture} = \frac{1}{1 \cdot 84} \times 100 = \mathbf{54 \cdot 35}$$

$$\% \text{ of } MgCO_3 \text{ in the mixture} = 100 - 54 \cdot 35 = \mathbf{45 \cdot 65}$$

EXAMPLE 7. *A solid mixture weighing 5·00 g containing lead nitrate and sodium nitrate was heated below 600°C until the mass of the residue was constant. If the loss of mass is 28%, find the mass of lead nitrate and sodium nitrate in the mixture.*

(At wts. of Pb = 207, Na = 23, N = 14, O = 16)

Solution.
$$\underset{2\times 331 = 662\text{ g}}{2\,Pb(NO_3)_2} \xrightarrow{\Delta} \underset{2\times 223 = 446\text{ g}}{2\,PbO} + 4\,NO_2 \uparrow + O_2 \uparrow$$

$$2\,NaNO_3 \xrightarrow{\Delta} 2\,NaNO_2 + O_2 \uparrow$$
$$\underset{2 \times 85 = 170\,g}{} \qquad \underset{2 \times 69 = 138\,g}{}$$

Suppose $Pb(NO_3)$ in the mixture $= x$ g

Then $NaNO_3$ in the mixture $= (5 - x)\,g$

662 g $Pb(NO_3)_2$ give residue $= 446$ g

x g $Pb(NO_3)_2$ will give residue $= \frac{446}{662} \times x$ g

$= 0 \cdot 674 x$ g

170 g $NaNO_3$ give residue $= 138$ g

$(5 - x)$ g $NaNO_3$ will give residue

$= \frac{138}{170} \times (5 - x)\,g = 0 \cdot 812\,(5 - x)$ g

Actual residue obtained

$= 5 - \frac{28}{100} \times 5 = 3 \cdot 6\,g$

$\therefore \quad 0 \cdot 674 x + 0 \cdot 812\,(5 - x) = 3 \cdot 6$

or $0 \cdot 138 x = 0 \cdot 46$

or $\quad x = 3 \cdot 33$ g

i.e. $Pb(NO_3)_2$ in the mixture $= 3 \cdot 33$ g

$NaNO_3$ in the mixture $= 5 - 3 \cdot 33$

$= 1 \cdot 67$ g

TYPE II. Involving Mass-Volume Relationship

EXAMPLE 8. *What volume of carbon dioxide measured at 27°C and 746·7 mm pressure would be obtained by treating 10·0 g of pure marble with dilute hydrochloric acid ? (Aq. tension at 27°C is 26·7 mm).*

Solution. The chemical equation representing the reaction is –

$$\underset{100\,g}{CaCO_3} + 2HCl \longrightarrow CaCl_2 + H_2O + \underset{\substack{22400\,cm^3 \\ \text{at STP}}}{CO_2}$$

Step 1. *To calculate the volume of CO_2 evolved from 10 g of $CaCO_3$ at STP.*

From the chemical equation,

100 g of marble react to produce 22400 cm^3 of CO_2 at STP

$\therefore$ 10g of it would produce

$$\frac{22400}{100} \times 10 = 2240 \text{ cm}^3$$

This is the volume of CO_2 evolved at STP.

Step 2. *To calculate the volume of CO_2 at 27°C and 746·7 mm pressure*

Initial conditions	*Final conditions*
Volume $V_1 = 2240$ cm^3	$V_2 = ?$ cm^3
Pressure $P_1 = 760$ mm	$P_2 = 746 \cdot 7 - 26 \cdot 7 = 720$ mm
Temperature $T_1 = 273$ K	$T_2 = 27 + 273 = 300$ K

By the gas equation, we have –

$$\frac{V_2 \times 720}{300} = \frac{2240 \times 760}{273}$$

$$V_2 = \frac{2240 \times 760}{273} \times \frac{300}{720} = 2598.3 \text{ cm}^3$$

Thus the volume of carbon dioxide

$= \mathbf{2598.3\ cm^3}$.

EXAMPLE 9. *1·0 g of a mixture of carbonates of calcium and magnesium gave 240 cm^3 of CO_2 at STP. Calculate the percentage composition of the mixture.* **(West Bengal J.E.E. 2003)**

Solution. Mass of mixture of carbonates of Ca and Mg taken $= 1 \cdot 0$ g

Let the mass of $CaCO_3 = x$ g

$\therefore$ Mass of $MgCO_3 = (1 - x)$ g

The chemical equations involved are :

$$\underset{\substack{40+12+3\times16 \\ = 100\,g}}{CaCO_3} \longrightarrow CaO + \underset{\substack{22400\,cm^3 \\ \text{at STP}}}{CO_2} \quad ...(i)$$

$$\underset{\substack{24+12+3\times16 \\ = 84\,g}}{MgCO_3} \longrightarrow MgO + \underset{\substack{22400\,cm^3 \\ \text{at STP}}}{CO_2} \quad ...(ii)$$

Step 1. *To calculate the volume of CO_2 evolved at STP from x g of $CaCO_3$.*

100 g of $CaCO_3$ evolve CO_2 at STP $= 22400$ cm^3

$\therefore \quad x$ g of $CaCO_3$ will evolve CO_2 at STP

$= \frac{22400}{100} \times x \text{ cm}^3 = 224\,x \text{ cm}^3$

Step 2. *To calculate the volume of CO_2 evolved at STP from (1 – x)g of $MgCO_3$.*

84 g of $MgCO_3$ evolve CO_2 at STP $= 22400$ cm^3

$\therefore (1 - x)$ g of $MgCO_3$ will evolve CO_2 at STP

$= \frac{22400}{84} \times (1 - x) \text{ cm}^3 = \frac{800}{3}(1 - x) \text{ cm}^3$

Step 3. *To calculate the value of x*

$\therefore$ Total volume of CO_2 evolved at STP

$$= 224\,x + \frac{800}{3}(1-x)\ \text{cm}^3$$

But total volume of CO_2 evolved at STP

$$= 240\ \text{cm}^3 \quad (Given)$$

$$\therefore \quad 224\,x + \frac{800}{3}(1-x) = 240.$$

or $\quad 672\,x + 800 - 800\,x = 720$

or $\quad 128\,x = 80$

$$\therefore \quad x = \frac{5}{8}$$

Step 4. *To calculate the percentage composition of the mixture.*

$$\therefore \text{ Percentage of } CaCO_3 = \frac{5}{8 \times 1} \times 100$$

$$= \mathbf{62 \cdot 5}$$

$$\therefore \text{ Percentage of } MgCO_3 = 100 - 62\cdot 5$$

$$= \mathbf{37 \cdot 5}$$

TYPE III. Involving Volume-Volume Relationship

EXAMPLE 10. *What volume of oxygen at STP is required to effect complete combustion of 200 cm³ of acetylene and what would be the volume of carbon dioxide formed ?*

Solution. The chemical equation representing the combustion of acetylene is

$$\underset{2\text{ Vol.}}{2\,C_2H_2} + \underset{5\text{ Vol.}}{5\,O_2} \longrightarrow \underset{4\text{ Vol.}}{4\,CO_2} + 2\,H_2O$$

Step 1. *To calculate the volume of O_2 at STP required to effect complete combustion of 200 cm³ of acetylene.*

Applying Gay Lussac's Law of gaseous volumes,

2 Vol. of C_2H_2 require O_2 for complete combustion = 5 Vol.

$\therefore$ 200 cm³ of C_2H_2 will require O_2 for complete combustion $= \frac{5}{2} \times 200 = 500$ cm³ at STP.

Thus the volume of O_2 required

= 500 cm³ at STP.

Step 2. *To calculate the volume of CO_2 produced at STP.*

Applying Gay Lussac's Law of gaseous volumes,

2 Vol. of C_2H_2 produce CO_2 = 4 Vol.

$\therefore$ 200 cm³ of C_2H_2 at STP will produce CO_2

$$= \frac{4}{2} \times 200 = 400\ \text{cm}^3 \text{ at STP}$$

Thus the volume of CO_2 produced

= 400 cm³ at STP.

PROBLEMS FOR PRACTICE

(A) Involving Mass-Mass Relationship

1. In the commercial manufacture of nitric acid, how many moles of NO_2 produce 7·33 mol of HNO_3 in the reaction :

   $3\,NO_2\,(g) + H_2O\,(l) \longrightarrow 2\,HNO_3\,(aq) + NO\,(g)$?

   (N.C.E.R.T.) [Ans. **10·995 moles**]

2. How much of Fe can be theoretically obtained by the reduction of 1 kg of Fe_2O_3 ? [Ans. **700 g**]

3. Calculate the mass of 60% H_2SO_4 required to decompose 50 g of chalk (calcium carbonate).

   [Ans. **81.67 g**]

4. Which is cheaper : 40% hydrochloric acid at the rate of 50 paise per kilogram or 80% H_2SO_4 at the rate of 25 paise per kilogram to completely neutralize 7 kg of caustic potash ?

   [Ans. **Cost of HCl = Rs. 5.70, Cost of H_2SO_4 = Rs. 1.91**]

5. Excess of $AgNO_3$ solution was added to 2.2 g of commercial sample of common salt dissolved in water. The mass of dried precipitate of silver chloride was 2.11 g. Calculate the percent purity of common salt.

   [Ans. **39.10**]

6. A sample of dolomite contained 45% of $CaCO_3$, 40% of $MgCO_3$ and 15% clay. Calculate the mass of sulphuric acid of 30% strength required to react completely with 10 g of the sample.

   [Ans. **30.27 g**]

PRACTICE PROBLEMS CONTD.

7. Calculate the mass of graphite that must be burnt to produce 13.2 g of CO_2. [Ans. 3.6 g]

8. One gram of a mixture of potassium and sodium chlorides on treatment with excess of silver nitrate gave 2 g of AgCl. What was the composition of the two salts in the original mixture ? [Ans. NaCl = 14%, KCl = 86%]

(B) Involving Mass-Volume Relationship

9. What volume of oxygen at 18°C and 750 mm pressure can be obtained from 10 g of potassium chlorate ? [Ans. 2.963 litres]

10. What mass of iodine is liberated from a solution of potassium iodide when 1 litre of chlorine gas at 10° C and 750 mm pressure is passed through it ? [Ans. 10.78 g]

11. 1.4 g of a sample of chalk ($CaCO_3$) containing clay as impurity were treated with excess of dilute hydrochloric acid. Volume of CO_2 evolved when measured at 15°C and 768 mm pressure was 282 cm^3. Calculate the percentage purity of the sample. [Ans. 86.1%]

12. How much marble of 96.5% purity would be required to prepare 10 litres of carbon dioxide at STP when the marble is acted upon by dilute hydrochloric acid ? [Ans. 46.26 g]

13. Calculate the volume of SO_2 at STP obtained by burning 500 g of S containing 4% sand by weight. [Ans. 336 litres]

14. 2.50 g sample of sodium bicarbonate when strongly heated gave 300 cm^3 of CO_2 measured at 27° and 760 mm pressure. Calculate the percentage purity of the sample. [Ans. 81·9%]

15. 10 ml of liquid carbon disulphide (sp. gravity 2·63) is burnt in oxygen. Find the volume of the resulting gases measured at S.T.P. [Ans. 23·25 L]

16. The drain cleaner, Drainex contains small bits of aluminium which react with caustic soda to produce hydrogen. What volume of hydrogen at 20°C and one bar will be released when 0·15 g of aluminium reacts ? *(N.C.E.R.T)* [Ans. 203·0 ml]

(C) Involving Volume-Volume Relationship

17. 5.6 litres of methane (CH_4) gas are ignited in oxygen gas. Calculate the number of moles of CO_2 formed. [Ans. 0.25 mole]

18. Calculate the volume of air containing 21% oxygen by volume at STP, required in order to convert 294 cm^3 of sulphur dioxide to sulphur trioxide under the same conditions. [Ans. 700 cm^3]

19. What volume of a solution of HCl containing 73 g/litre would suffice for the exact neutralization of NaOH obtained by allowing 0.46 g of metallic sodium to react with water. [Ans. 10.0 cm^3]

20. Find out the volume of Cl_2 at STP produced by the action of 100 cm^3 of 0.2 N HCl on excess of MnO_2. [Ans. 112 cm^3]

HINTS FOR DIFFICULT PROBLEMS

1. 2 moles of HNO_3 are produced from 3 moles of NO_2 ∴ 7·33 moles of HNO_3 will be produced from

$NO_2 = \frac{3}{2} \times 7{\cdot}33 = 10{\cdot}995$ moles.

2. $\underset{\substack{2 \times 56 + 3 \times 16 \\ = 160\ g}}{Fe_2O_3} \equiv \underset{\substack{2 \times 56 \\ = 112\ g}}{2\,Fe}$.

3. $\underset{100\ g}{CaCO_3} + \underset{98\ g}{H_2SO_4} \longrightarrow CaSO_4 + H_2O + CO_2$

50 g chalk will require pure H_2SO_4 = 49 g

∴ 60% H_2SO_4 required $= \frac{100}{60} \times 49 = 81{\cdot}67$ g

4. $\underset{56\ kg}{KOH} + \underset{36\cdot5\ kg}{HCl} \longrightarrow KCl + H_2O$

Pure HCl and required for 7 kg KOH $= \frac{36{\cdot}5}{56} \times 7$

$= 4{\cdot}5625$ kg

40% HCl acid required $= \frac{100}{40} \times 4{\cdot}5625$

$= 11{\cdot}40625$ kg

Cost = 11·40625 × 50 p = Rs. 5·70 p.

$\underset{2 \times 56\ kg}{2\,KOH} + \underset{98\ kg}{H_2SO_4} \longrightarrow K_2SO_4 + 2\,H_2O$

Proceed as above to calculate the cost of 80% H_2SO_4.

9. $\underset{\substack{2\,(39 + 35{\cdot}5 + 48) \\ = 245\ g}}{2\,KClO_3} \longrightarrow 2\,KCl + \underset{3 \times 22\cdot4\ L\ at\ STP}{3\,O_2}$

10. $2\,KI + \underset{22\cdot4\ L\ at\ STP}{Cl_2} \longrightarrow 2\,KCl + \underset{254\ g}{I_2}$

First convert the given volume to volume at S.T.P.

HINTS CONTD.

11. $\underset{100\text{ g}}{CaCO_3} + 2HCl \longrightarrow CaCl_2 + H_2O + \underset{22400\text{ c.c. at S.T.P.}}{CO_2}$

First convert the given volume to volume at S.T.P. Calculate the mass of $CaCO_3$ from which this volume of CO_2 at S.T.P. is obtained. Then calculate %age purity.

12. $\underset{100\text{ g}}{CaCO_3} + 2HCl \longrightarrow CaCl_2 + H_2O + \underset{22\cdot4\text{ L at STP}}{CO_2}$

10 L of CO_2 at STP will be obtained from pure

$CaCO_3 = \frac{100}{22\cdot4} \times 10 = 44\cdot64$ g

Impure marble required $= \frac{100}{96\cdot5} \times 44\cdot64$
$= 46\cdot26$ g

15. 10 ml of $CS_2 = 10 \times 2\cdot63$ g $= 26\cdot3$ g

$$\underset{76\text{ g}}{CS_2} + 3O_2 \longrightarrow \underset{\substack{3\text{ moles}\\=3\times22\cdot4\text{ L at STP}}}{CO_2 + 2SO_2}$$

16. $\underset{2\times27\text{ g}}{2Al} + 2NaOH + 2H_2O \longrightarrow 2NaAlO_2 + \underset{\substack{3\times22400\text{ ml}\\\text{at StP}}}{3H_2}$

$\therefore$ H_2 produced at STP from 0·15 g Al

$= \frac{3\times22400}{54} \times 0\cdot15$ ml $= 186\cdot7$ ml

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

(STP conditions) (Required conditions)

i.e. $\frac{1\text{ atm} \times 186\cdot7\text{ ml}}{273\text{ K}} = \frac{0\cdot987\text{ atm} \times V_2}{293\text{ K}}$

(1 bar = 0·987 atm)

or $V_2 = 203\cdot0$ ml

18. $2SO_2 + O_2 \longrightarrow SO_3$.

O_2 required $= \frac{1}{2} \times 294 = 147$ cc.

Air required $= \frac{100}{21} \times 147 = 700$ cc.

19. $\underset{\substack{2\times23\text{ g}\\=46\text{ g}}}{2Na} + 2H_2O \longrightarrow \underset{\substack{2\times40\text{ g}\\=80\text{ g}}}{2NaOH} + H_2$

NaOH produced from 0·46 g Na = 0·80 g

$\underset{40\text{ g}}{NaOH} + \underset{36\cdot5\text{ g}}{HCl} \longrightarrow NaCl + H_2O$

HCl required for exact neutralisation of 0·80 g

NaOH $= \frac{36\cdot5}{40} \times 0\cdot80$ g $= 0\cdot73$ g

73 g HCl are present in 1000 cm^3

$\therefore$ 0·73 g HCl will be present in 10 cm^3

20. $MnO_2 + \underbrace{4HCl}_{4\times36\cdot5\text{ g}} \longrightarrow MnCl_2 + 2H_2O + \underbrace{Cl_2}_{22400\text{ cm}^3\text{ at STP}}$

HCl present in 100 cm^3 of 0·2 N HCl

$= \frac{0\cdot2}{1000} \times 100$ g eq. $= 0\cdot02$ g eq.

$= 0\cdot02 \times 36\cdot5$ g $= 0\cdot73$ g

1.36. Stoichiometry of Reactions in Solution

The general method of calculations is same as discussed in section 1.35. However, the only additional calculation involved is as follows :

If volume of a solution of known molarity or normality is given, calculate the amount in grams or moles present in it keeping in mind that 1000 cc of 1 M solution contain 1 mol of the solute or 1000 cc of 1 N solution contain 1 gram equivalent of the solute. Conversely, from the balanced equation, the amount of the required substance in moles that will react can be calculated and knowing the volume of the solution of that substance (as given in the problem), molarity of the solution of that substance can be calculated or if molarity of the solution of that substance is given instead of volume, the volume of the solution can be calculated.

A few examples solved below will help to understand the calculations more clearly.

EXAMPLE 1. ***Calcium carbonate reacts with aqueous HCl according to the reaction***

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l).$$

What mass of $CaCO_3$ is required to react completely with 25 mL of 0·75 M HCl ? **(N.C.E.R.T.)**

Solution. Step 1. ***To calculate mass of HCl in 25 mL of 0·75 M HCl***

1000 mL of 0·75 M HCl contain HCl

$= 0\cdot75$ mol

$= 0\cdot75 \times 36\cdot5\,g = 24\cdot375\,g$

$\therefore$ 25 mL of 0·75 HCl will contain HCl

$= \frac{24 \cdot 375}{1000} \times 25\ g = 0 \cdot 6844\ g.$

Step 2. *To calculate mass of $CaCO_3$ reacting completely with $0 \cdot 9125$ g of HCl*

$$CaCO_3\,(s) + 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)$$

2 mol of HCl *i.e.* $2 \times 36 \cdot 5\ g = 73$ g HCl react completely with $CaCO_3 = 1$ mol $= 100$ g

$\therefore$ $0 \cdot 6844$ g HCl will react completely with

$CaCO_3 = \frac{100}{73} \times 0 \cdot 6844\ g = \mathbf{0 \cdot 938\ g}.$

EXAMPLE 2. *Calculate the volume of hydrogen liberated at STP when 500 cm^3 of $0 \cdot 5$ N sulphuric acid reacts with excess of zinc. (H = 1, O = 16, S = 32).*

Solution. The chemical equation representing the reaction is :

$$\underset{=98 \cdot 0\ g}{Zn + H_2SO_4} \longrightarrow ZnSO_4 + \underset{\substack{=22400\ cm^3 \\ \text{at STP}}}{H_2}$$

Step 1. *To calculate the amount of H_2SO_4 in 500 cm^3 of $0 \cdot 5$ N H_2SO_4 solution.*

Strength/litre of H_2SO_4 solution

$=$ Normality $\times$ Eq. mass

$= 0 \cdot 5 \times 49 = 24 \cdot 5$ g/litre

Now, 1000 cm^3 of the acid solution contain $24 \cdot 5$ g of pure H_2SO_4

$\therefore$ 500 cm^3 of the acid solution will contain

$24 \cdot 5 \times \frac{500}{1000} = 12 \cdot 25$ g of pure H_2SO_4

Step 2. *To calculate the volume of H_2 liberated at STP*

98 g of H_2SO_4 react to liberate 22400 cm^3 of H_2 at STP

$\therefore$ $12 \cdot 25$ g of H_2SO_4 would liberate H_2 at STP

$= \frac{22400}{98} \times 12 \cdot 25\ cm^3 = 2800\ cm^3$

Thus the volume of hydrogen liberated at STP = **2800 cm^3**

EXAMPLE 3. *Calculate the volume of $0 \cdot 05$ M $KMnO_4$ solution required to oxidise completely $2 \cdot 70$ g of oxalic acid ($H_2C_2O_4$) in acidic medium.*

Solution. The equation representing the chemical change is –

$$2KMnO_4 + 3H_2SO_4 \longrightarrow K_2SO_4 + 2MnSO_4 + 3H_2O + 5\,(O)$$

$$[H_2C_2O_4 + (O) \longrightarrow 2CO_2 + H_2O\,] \times 5$$

$$\underset{2\ \text{moles}}{2KMnO_4} + 3H_2SO_4 + \underset{5\ \text{moles}}{5H_2C_2O_4} \longrightarrow K_2SO_4 + 2MnSO_4 + 10CO_2 + 8H_2O$$

Step 1. *To calculate the number of moles of $KMnO_4$ required to completely oxidise $2 \cdot 70$ g of $H_2C_2O_4$ in acidic medium.*

Molar mass of $H_2C_2O_4$

$= 2 \times 1 + 2 \times 12 + 4 \times 16 = 90 \cdot 0\ g\ mol^{-1}$

$\therefore$ No. of moles of $H_2C_2O_4$ contained in $2 \cdot 70$ g of it $= \frac{\text{Mass in grams}}{\text{Mol. mass}} = \frac{2 \cdot 70}{90} = 0 \cdot 03$

From the above equation,

5 moles of $H_2C_2O_4$ are oxidised by 2 moles of $KMnO_4$

$\therefore$ $0 \cdot 03$ mole of $H_2C_2O_4$ will be oxidised by

$\frac{2 \times 0 \cdot 03}{5} = 0 \cdot 012$ mole of $KMnO_4$

Step 2. *To calculate the volume of $0 \cdot 05$ M $KMnO_4$ solution.*

Now, $0 \cdot 05$ mole of $KMnO_4$ are contained in 1000 cm^3 of the solution.

$\therefore$ $0 \cdot 012$ mole of $KMnO_4$ wil be contained in

$\frac{1000}{0 \cdot 05} \times 0 \cdot 012 = 240\ cm^3$ of solution.

Thus the required volume of $0 \cdot 05$ M $KMnO_4$ solution = **240 cm^3**.

PROBLEMS FOR PRACTICE

1. 5.0 g of marble was added to 7.5 g dilute hydrochloric acid. After the reaction was over, it was found that 0.5 g of marble was left unused. Calculate the percentage strength of hydrochloric acid. What volume of CO_2 measured at STP will be evolved in the above reaction ?

[Ans. **43.8% and 1008 cm^3**]

PRACTICE PROBLEMS CONTD.

2. Calculate the volume of $1 \cdot 00$ mol L^{-1} aqueous sodium hydroxide that is neutralized by 200 mL of $2 \cdot 00$ mol L^{-1} aqueous hydrochloric acid and mass of sodium chloride produced. (N.C.E.R.T.)

[Ans. 400 mL, 23·4 g]

3. Bromine is prepared commercially by the reaction

$2Br^-(aq) + Cl_2(aq) \longrightarrow 2Cl^-(aq) + Br_2(aq)$

Suppose we have $50 \cdot 0$ mL of $0 \cdot 060$ M solution of NaBr. What volume of $0 \cdot 050$ M solution of Cl_2 is needed to react completely with the Br^- ?

[Ans. 30 mL Cl_2 solution]

HINTS FOR DIFFICULT PROBLEMS

1. $\underset{100\ g}{CaCO_3} + \underset{\substack{2 \times 36 \cdot 5\ g \\ = 73\ g}}{2HCl} \longrightarrow CaCl_2 + H_2O + \underset{\substack{22400\ cm^3 \\ \text{at STP}}}{CO_2}$

Marble reacted $= 5 - 0 \cdot 5\ g = 4 \cdot 5\ g$

HCl reacted with $4 \cdot 5$ g marble $= \frac{73}{100} \times 4 \cdot 5\ g$

$= 3 \cdot 285\ g$

% strength $= \frac{3 \cdot 285}{7 \cdot 5} \times 100 = 43 \cdot 8\%$

CO_2 evolved at S.T.P. $= \frac{22400}{100} \times 4 \cdot 5\ cm^3$

$= 1008\ cm^3$

2. $\underset{(NaOH)}{M_1V_1} = \underset{(HCl)}{M_2V_2}$ *i.e.* $1 \cdot 0 \times V_1 = 2 \cdot 0 \times 200$

or $V_1 = 400$ mL

$NaOH + HCl \longrightarrow NaCl + H_2O$

200 mL of $2 \cdot 0$ M HCl contain HCl

$= \frac{2 \cdot 0}{1000} \times 200 = 0 \cdot 4$ mol

1 mol of HCl produces NaCl = 1 mol

$\therefore$ $0 \cdot 4$ mol of HCl will produce NaCl $= 0 \cdot 4$ mol

$= 0 \cdot 4 \times 58 \cdot 5\ g = 23 \cdot 4\ g.$

3. 50 mL of $0 \cdot 060$ M NaBr contain NaBr

$= \frac{0 \cdot 060}{1000} \times 50$ mol $= 0 \cdot 003$ mol

2 mol of Br^- react with Cl_2 = 1 mol

$\therefore$ $0 \cdot 03$ mol of Br^- will react with Cl_2

$= \frac{1}{2} \times 0 \cdot 03 = 0 \cdot 0015$ mol

$0 \cdot 05$ mol of Cl_2 solution are present in 1000 mL of Cl_2 solution

$\therefore$ $0 \cdot 015$ mol of Cl_2 will be present in Cl_2 solution

$= \frac{1000}{0 \cdot 05} \times 0 \cdot 0015 = 30$ mL.

1.37. Limiting Reactant

Quite often, one of the reactants is present in larger amount than the other as required according to the balanced equation. The amount of the product formed then depends upon the reactant which has reacted completely. This reactant is called the **limiting reactant** or **limiting reagent.** The reactant which is not consumed completely in the reaction is called **excess reactant** as the excess of this reactant is left unreacted.

SOLVED EXAMPLES

EXAMPLE 1. *Zinc and hydrochloric acid react according to the reaction*

$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$

If $0 \cdot 30$ mol Zn are added to hydrochloric acid containing $0 \cdot 52$ mol of HCl, how many moles of H_2 are produced ? **(N.C.E.R.T.)**

Solution. The reaction is

$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$

Thus, 1 mol of zinc reacts with 2 moles of HCl.

$\therefore$ $0 \cdot 30$ mol of zinc will react with HCl $= 2 \times 0 \cdot 30 = 0 \cdot 60$ mol

But we have only $0 \cdot 52$ mol of HCl. Hence zinc cannot react completely and hence is not a limiting reactant.

Again, 2 moles of HCl react with zinc = 1 mol

$\therefore$ $0 \cdot 52$ mol of HCl will react with zinc $= \frac{1}{2} \times 0 \cdot 52 = 0 \cdot 26$ mol

As we have $0 \cdot 30$ mol of zinc, therefore, HCl will react completely *i.e. HCl is the limiting reactant.*

2 moles of HCl produce $H_2 = 1$ mol

$\therefore$ $0 \cdot 52$ mol of HCl will produce H_2

$= \frac{1}{2} \times 0 \cdot 52 \text{ mol} = \mathbf{0 \cdot 26 \text{ mol}}$

EXAMPLE 2. *3.0 g of H_2 react with 29.0 g of O_2 to form H_2O.*

(i) Which is the limiting reactant ?

(ii) Calculate the maximum amount of H_2O that can be formed.

(iii) Calculate the amount of the reactant left unreacted.

Molecular mass of H_2 = 2.016.

Solution.
$$\underset{\substack{2 \times 2 \cdot 016 \\ = 4 \cdot 032 \text{ g}}}{2H_2} + \underset{32 \text{ g}}{O_2} \longrightarrow \underset{\substack{2 \times (2 \cdot 016 + 16) \\ = 36 \cdot 032}}{2H_2O}$$

$$3 \text{ g of } H_2 \text{ require } O_2 = \frac{32}{4 \cdot 032} \times 3 = 23 \cdot 8 \text{ g}$$

Thus O_2 (29 g) is present in excess. Hence **H_2 is the limiting reactant.**

$$H_2O \text{ formed} = \frac{36 \cdot 032}{4 \cdot 032} \times 3 \text{ g} = \mathbf{26 \cdot 8 \text{ g}}$$

$$O_2 \text{ left unreacted} = 29 - 23 \cdot 8 = \mathbf{5 \cdot 2 \text{ g}}$$

EXAMPLE 3. *One litre of oxygen at STP is made to react with three litres of carbon monoxide at STP. Calculate the mass of each substance found after the reaction. Which one is the limiting reactant ?*

Solution. The chemical equation representing the reaction is

$$\underset{2 \text{ vol.}}{2CO} + \underset{1 \text{ vol.}}{O_2} \longrightarrow \underset{2 \text{ vol.}}{2CO_2}$$

Step 1. *To calculate the volume and mass of CO (at STP) left unused after the reaction.*

Applying Gay Lussac's Law of Gaseous Volumes,

1 Vol. of O_2 reacts with CO = 2 Vol.

$\therefore$ 1 litre of O_2 will react with CO = 2 litres at STP

But the volume of CO taken = 3 litre at STP

$\therefore$ Volume of CO (at STP) left unused

$= 3 - 2 = 1$ litre.

Now by mole concept,

Mass of $22 \cdot 4$ litres of CO at STP

= Molecular mass in grams

$= 12 + 16 = 28$ g

$$\therefore \text{Mass of 1 litre of CO at STP} = \frac{28}{22 \cdot 4} \times 1 = 1 \cdot 25 \, g$$

Thus the mass of CO left unused = **$1 \cdot 25$ g**

Step 2. *To calculate the volume and mass of CO_2 formed from 1 litre of O_2 at STP.*

Applying Gay Lussac's Law of Gaseous Volumes,

1 Vol. of O_2 produces CO_2 = 2 Vol.

$\therefore$ 1 litre of O_2 will produce CO_2 = 2 litres at STP

By mole concept,

Mass of $22 \cdot 4$ litres of CO_2 at STP

= Molecular mass in grams

$= 12 + 2 \times 16 = 44$ g

$\therefore$ Mass of 2 litres of CO_2 at STP

$$= \frac{44}{22 \cdot 4} \times 2 = 3 \cdot 928 \, g$$

Thus the mass of CO_2 produced = **$3 \cdot 928$ g.**

Step 3. As oxygen has been completely used up, hence *oxygen is the limiting reactant.*

PROBLEMS FOR PRACTICE

1. 500 cm^3 of 0.250 M Na_2SO_4 solution added to an aqueous solution of 15.00 g of $BaCl_2$ resulted in the formation of a white precipitate of $BaSO_4$. How many moles and how many grams of $BaSO_4$ are formed ? [Ans. **0.072 mole, 16.776 g**]
2. If $20 \cdot 0$ g of $CaCO_3$ is treated with $20 \cdot 0$ g of HCl, how many grams of CO_2 will be produced ? [Ans. **$8 \cdot 80$ g**]

HINTS FOR DIFFICULT PROBLEMS

1. 500 cm^3 of 0·25 M $Na_2SO_4 = \frac{0\cdot25}{1000} \times 500$

$= 0\cdot125$ mole

15 g $BaCl_2 = \frac{15}{208}$ mole $= 0\cdot072$ mole

$Na_2SO_4 + BaCl_2 \longrightarrow BaSO_4 + 2\,NaCl$

Evidently, $BaCl_2$ will be the limiting reactant.

$BaSO_4$ formed $= 0\cdot072$ mole $= 0\cdot072 \times 233$ g

$= 16\cdot776$ g

2.
$$\underset{\substack{40+12+3\times16 \\ =100\,g}}{CaCO_3} + \underset{\substack{2(1+35\cdot5) \\ =73\,g}}{2\,HCl} \longrightarrow CaCl_2 + H_2O + \underset{\substack{12+2\times16 \\ 44\,g}}{CO_2}.$$

Here $CaCO_3$ will be the limiting reactant.

ADD TO YOUR KNOWLEDGE

1. **Percent yield.** The actual yield of a product in any reaction is usually less than the theoretical yield (calculated from the balanced chemical equation) because of certain side reactions taking place or less ideal conditions than required. Hence we have Percent yield $= \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$

2. **Eudiometry.** This is a method for finding the molecular formula of a gaseous hydrocarbon. The apparatus is shown in the Fig 1.6. The method involves the following steps :

(*i*) A known volume of the gaseous hydrocarbon is mixed with an excess (known or unknown volume) of oxygen or air in the eudiometer over mercury.

(*ii*) The mixture is exploded by electric spark and then cooled so that water vapour condense to liquid whose volume is negligible

$$\underset{1\,cc}{C_xH_y(g)} + \underset{\left(x+\frac{y}{4}\right)cc}{\left(x+\frac{y}{4}\right) O_2(g)} \longrightarrow \underset{x\,cc}{x\,CO_2(g)} + \frac{y}{2}H_2O(l)$$

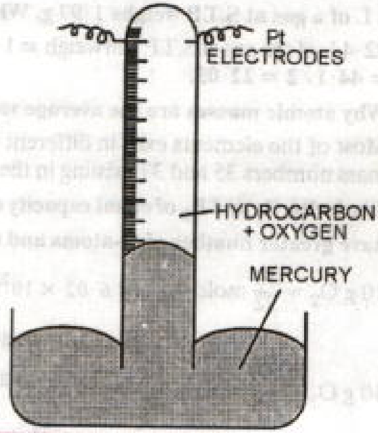

FIGURE 1.6. Eudiometry.

(*iii*) KOH is introduced which absorbs CO_2 and only unused O_2 is left. Thus decrease in volume on adding KOH = Volume of CO_2 produced

Values of x and y are then calculated from the following data :

(*i*) Volume of O_2 used per cc of hydrocarbon

$= \left(x+\frac{y}{4}\right) cc$

(*ii*) Volume of CO_2 produced $= x\,cc$

(*iii*) Contraction on explosion and cooling

$= \left[1+\left(x+\frac{y}{4}\right)\right]-x = \left(\frac{1+y}{4}\right) cc$

Conceptual Questions

Q. 1. What physical quantities are represented by the following units and what are their most common names ?

(i) kg ms^{-2} (ii) kg $m^2 s^{-2}$ (iii) dm^3

Ans. (*i*) Force (newton) (*ii*) Work (joule) (*iii*) Volume

Q. 2. Rewrite the following after required corrections :

(i) The length of a rod is 10 cms. (ii) The work done by a system is 10 Joules.

Ans. (*i*) The length of a rod is 10 cm (*s* is not used).

(*ii*) The work done by a system is 10 joules (small letter is used in place of capital letter).

Q. 3. Classify the following substances into elements, compounds and mixtures :

(i) Milk (ii) 22-carat gold (iii) Iodized table salt (iv) Diamond (v) Smoke (vi) Steel (vii) Brass (viii) Dry ice (ix) Mercury (x) Air (xi) Aerated drinks (xii) Glucose (xiii) Petrol/Diesel/Kerosene oil (xiv) Steam (xv) Cloud.

Ans. Elements — (*iv*), (*ix*)

Compounds — (*viii*), (*xii*), (*xiv*), (*xv*)

Mixtures— (*i*), (*ii*), (*iii*), (*v*), (*vi*), (*vii*), (*x*), (*xi*), (*xiii*).

Q. 4. Why air is sometimes considered as a heterogeneous mixture ?

Ans. This is due to the presence of dust particles which form a separate phase.

Q. 5. Why Law of conservation of mass should better be called as Law of conservation of mass and energy ?

Ans. In nuclear reactions, it is observed that the mass of the products is less than the mass of the reactants. The difference of mass, called the mass defect, is converted into energy according to Einstein equation, $E = \Delta m c^2$. Hence we better call it as a law of conservation of mass and energy.

Q. 6. Is the law of constant composition true for all types of compounds ? Explain why or why not.

Ans. No, law of constant composition is not true for all types of compounds. It is true only for the compounds obtained from one isotope. For example, carbon exists in two common isotopes, ^{12}C and ^{14}C. When it forms CO_2 from ^{12}C, the ratio of masses is 12 : 32 = 3 : 8 but from ^{14}C, the ratio will be 14 : 32 = 7 : 16 which is not same as in the first case.

Q. 7. 1 L of a gas at S.T.P. weighs 1·97 g. What is the vapour density of the gas ?

Ans. 22·4 L of the gas at S.T.P. will weigh = 1·97 × 22·4 = 44·1 g *i.e.* molecular mass = 44·1. Hence vapour density = 44·1/2 = 22·05.

Q. 8. Why atomic masses are the average values ?

Ans. Most of the elements exist in different isotopes *i.e.* atoms with different masses *e.g.* Cl has two isotopes with mass numbers 35 and 37 existing in the ratio 3 : 1. Hence average value is taken.

Q. 9. Two bulbs B_1 and B_2 of equal capacity contain 10g oxygen (O_2) and ozone (O_3) respectively. Which bulb will have greater number of O-atoms and which will have greater number of molecules ?

Ans. $10 \text{ g } O_2 = \frac{10}{32} \text{ mole} = \frac{10}{32} \times 6 \cdot 02 \times 10^{23}$ molecules

$= 1 \cdot 88 \times 10^{23}$ molecules $= 2 \times 1 \cdot 88 \times 10^{23}$ atoms $= 3 \cdot 76 \times 10^{23}$ atoms

$10 \text{ g } O_3 = \frac{10}{48} \text{ mole} = \frac{10}{48} \times 6 \cdot 02 \times 10^{23}$ molecules

$= 1 \cdot 254 \times 10^{23}$ molecules $= 3 \times 1 \cdot 254 \times 10^{23}$ atoms $= 3 \cdot 76 \times 10^{23}$ atoms

Thus both contain the same number of atoms but bulb B_1 contains more number of molecules.

Q. 10. Determine the equivalent weight of each of the following compounds assuming the formula weights of these compounds are *x*, *y* and *z* respectively.

(i) Na_2SO_4 (ii) $Na_3PO_4 . 12 H_2O$ (iii) $Ca_3(PO_4)_2$ ***(M.L.N.R. Allahabad 1991)***

Ans. Eq wt = Mol. wt/Total positive valency of metal atoms

(*i*) $x/2$ (*ii*) $y/3$ (*iii*) $z/6$.

Q. 11. Why are the atomic masses of most of the elements fractional ?

Ans. This is because atomic masses are the relative masses of atoms as compared with an atom of C—12 isotope taken as 12.

Q. 12. Write the formulae and names of three compounds containing same percentage composition of C, H and O.

Ans. Compounds with the same percentage composition of C, H and O will have the same empirical formula. The compounds with the empirical formula CH_2O can be

HCHO	CH_3COOH	$C_6H_{12}O_6$
Formaldehyde	Acetic acid	Glucose

Very Short Answer Questions CARRYING 1 MARK

Q. 1. What is the number of significant figures in $1 \cdot 050 \times 10^4$?

Ans. four

Q. 2. What is the S.I. unit of density ?

Ans. kg m^{-3}

Q. 3. What is AZT ? To which use is it being put ?

Ans. Azidothymidine used for AIDS victims.

Q. 4. What is the law called which deals with the ratios of the volumes of the gaseous reactants and products?

Ans. Gay Lussac's law of gaseous volumes

Q. 5. What is one a.m.u. or one 'u' ?

Ans. 1 a.m.u. or 1 u = $\frac{1}{12}$th of the mass of an atom of carbon—12.

Q. 6. Which isotope of carbon is used for getting relative atomic masses ?

Ans. C—12

Q. 7. Write down the empirical formula of acetic acid.

Ans. CH_2O.

Q. 8. What is the S.I. unit of molarity ?

Ans. mol dm^{-3}.

Short Answer Questions CARRYING 2 or 3 MARKS

Sec. 1.1. to 1.3.
1. Chemistry can prove to be a blessing or a curse depends upon the uses to which it is put. Comment.

Sec. 1.4. to 1.6.
2. Briefly explain the difference between precision and accuracy.
3. What do you mean by 'significant figures' ?
4. Under what conditions the zeros in a number are significant ?
5. After rounding off, what will be the value of
   (*i*) 1·235 (*ii*) 1·225 ?
6. What does symbol SI signify ? Name the seven basic SI units ?
7. What do you understand by unit conversion factor? How does it help to convert height in feet to height in metres ?

Sec. 1.8. to 1.12.
8. Define Element, Compound and Mixture.
9. Give examples of homogeneous mixtures in different physical states (two each).
10. Give three main points of difference between a compound and a mixture.
11. Classify the following substances into elements, compounds and mixtures. Further separate the mixtures into homogeneous or heterogeneous :

(*i*) air (*ii*) milk (*iii*) graphite (*iv*) diamond (*v*) gasoline (*vi*) tap water (*vii*) distilled water (*viii*) oxygen (*ix*) one rupee coin (*x*) 22 carat gold (*xi*) steel (*xii*) iron (*xiii*) sodium chloride (*xiv*) iodized table salt.

Sec. 1.14. to 1.18.

12. Give one experiment involving a chemical reaction to prove that the law of conservation of mass is true.
13. Copper oxide obtained by heating copper carbonate or copper nitrate contains copper and oxygen in the same ratio by mass. Which law is illustrated by this observation ? State the law.
14. Nitrogen forms a number of oxides. Write their formulae and give their names. Giving appropriate calculations, explain the law that follows from it.
15. N and O combine with H to form NH_3 and H_2O and they combine with each other to form NO_2. Which law is illustrated ? Explain briefly.
16. Define Gay Lussac's law of gaseous volumes. Explain with one suitable example.

Sec. 1.20. to 1.22.

17. What are the postulates of Dalton's Atomic Theory ? How do the laws of chemical combination follow from it ?
18. What are the postulates of Modern Atomic Theory?

Sec. 1.24. to 1.27.

19. Which isotope is used as a reference on the atomic scale ? What is one amu or one 'u' ?
20. Why atomic mass is an average value ? Explain with a suitable example.

Sec. 1.28.

21. How is mole related to
    (*a*) number of atoms/molecules
    (*b*) mass of the substance
    (*c*) volume of the gaseous substance ?
22. What is the SI definition of mole ?
23. Comment on the following :—
    '1 mole of hydrogen'

Sec. 1.29.

24. What is the difference between
    (*i*) Normality and Molarity ? (*ii*) Molarity and Molality ?

Sec. 1.30. to 1.34.

25. Define Empirical formula and Molecular formula. What is the relationship between them ?
26. Write the empirical formulae of the following :
    (*i*) N_2O_4 (*ii*) C_6H_6 (*iii*) $C_6H_{12}O_6$ (*iv*) H_2O_2 (*v*) H_2O (*vi*) Na_2CO_3 (*vii*) CH_3COOH

Sec. 1.37.

27. What is a limiting reactant ? Explain with a suitable example.

Long Answer Questions CARRYING 5 or more MARKS

Sec. 1.1. to 1.3.

1. Briefly explain the importance of studying chemistry.
2. What do you understand by 'significant figures' ? What are the rules for determining the number of significant figures? Illustrate with suitable examples.
3. (*a*) What do you understand by 'Rounding off' ?
   (*b*) What are the rules for determining the number of significant figures in answers involving
   (*i*) addition or subtraction
   (*ii*) multiplication or division
   (*iii*) calculation involving a number of steps
   Explain each case with a suitable example.

Sec. 1.4. to 1.6.

4. What is the latest and the most scientific system of units ? What are the seven basic units in this system ? What are derived units ? Give the derived units of the following physical quantities :
   Area, Volume, Density, Speed, Acceleration, Force, Pressure and Energy (or Work).
5. What do you understand by Dimensional Analysis ? Explain with a suitable example.

Sec. 1.7. to 1.8. **6.** Define 'matter'. Briefly describe the physical as well as chemical classification of matter.

Sec. 1.9. to 1.12. **7.** Define Element, Compound and Mixture. What are homogeneous and heterogeneous mixtures ? Give one example of each of the homogeneous mixture in the gaseous phase, liquid phase and solid phase.

8. List the main points of difference between a compound and a mixture.

Sec. 1.14. **9.** State and explain the 'Law of Conservation of Mass'.

Sec. 1.15. **10.** Define 'Law of Constant Composition or Definite Proportions'. Explain with a suitable example.

Sec. 1.16. **11.** State the Law of Multiple Proportions. Explain with two suitable examples.

Sec. 1.17. **12.** State and explain Law of Reciprocal Proportions.

Sec. 1.18. **13.** What is Gay Lussac's Law of Gaseous Volumes ? Explain with two suitable examples.

Sec. 1.20. to 1.22. **14.** What are the main postulates of Dalton's Atomic Theory ? What were its limitations? How has the theory been modi- fied ?

Sec. 1.23. to 1.27. **15.** Define the following :

(*i*) Atom (*ii*) Molecule (*iii*) Atomic mass (*iv*) Gram atomic mass
(*v*) Molecular mass (*vi*) Gram molecular mass.

Give suitable examples in each case.

16. Define Avogadro's law. Taking a suitable example prove that it is not in contradiction with Dalton's Atomic Theory.

Sec. 1.28. **17.** Define 'Avogadro's Number' and 'Mole'. What is their importance ?

Sec. 1.29. **18.** (*a*) Define the terms Solution, Solute and Solvent.

(*b*) Define the different methods used for expressing the concentration of a solution.

Sec. 1.30. to 1.34. **19.** What are Empirical and Molecular Formulae? How are they related to each other?

COMPETITION FOCUS

ADDITIONAL USEFUL INFORMATION

1. Average Deviation. The simplest way to express precision is in terms of average deviation. For this purpose, first the average of the different measurements is calculated. Then the deviation of each measurement from the average is calculated, taking all deviations as positive. The average of these deviations is called average deviation.

2. Relative average deviation (r.a.d.). Precision is more frequently expressed in terms of relative average deviation which is equal to average deviation divided by average value of different measurements.

Example.

Electronic charge (esu) of individual measurements	*Individual deviation from the average*
$4 \cdot 80 \times 10^{-10}$	$0 \cdot 01 \times 10^{-10}$
$4 \cdot 79 \times 10^{-10}$	$0 \cdot 00 \times 10^{-10}$
$4 \cdot 81 \times 10^{-10}$	$0 \cdot 02 \times 10^{-10}$
$4 \cdot 76 \times 10^{-10}$	$0 \cdot 03 \times 10^{-10}$
Total = $19 \cdot 16 \times 10^{-10}$	Total = $0 \cdot 06 \times 10^{-10}$

$$\text{Average} = \frac{19 \cdot 16 \times 10^{-10}}{4} = 4 \cdot 79 \times 10^{-10}\, esu,$$

$$\text{Average dev.} = \frac{0 \cdot 06 \times 10^{-10}}{4} = 0 \cdot 02 \times 10^{-10}\, esu$$

$\therefore$ Result will be reported as $4 \cdot 79 \times 10^{-10} \pm 0 \cdot 02 \times 10^{-10}\, esu$

$$\text{Relative average deviation } (r.a.d) = \frac{0 \cdot 02 \times 10^{-10}\, esu}{4 \cdot 79 \times 10^{-10}\, esu} = 0 \cdot 0042 \text{ or } 0 \cdot 42\% \text{ or } 4200 \text{ parts per million (ppm)}$$

3. Relative error. Accuracy is expressed in terms of the absolute error or the relative error.

$$\text{Absolute error} = \text{Experimentally determined value—Accepted value}$$

$$\text{Relative error} = \frac{\text{Absolute error}}{\text{Accepted value}}$$

Example. Experimentally determined value of electronic charge = $4 \cdot 79 \times 10^{-10}$ *e.s.u.*

Accepted value of electronic charge = $4 \cdot 80 \times 10^{-10}$ *e.s.u.*

$$\text{Error (or Absolute error)} = 4 \cdot 79 \times 10^{-10} - 4 \cdot 80 \times 10^{-10} = -0 \cdot 01 \times 10^{-10}\, e.s.u.$$

$$\text{Relative error} = \frac{-0 \cdot 01 \times 10^{-10}}{4 \cdot 80 \times 10^{-10}} \times 100 = -0 \cdot 2\%$$

ADDITIONAL USEFUL INFORMATION contd.

4. Equivalent mass. The equivalent mass of a substance is defined as the number of parts by mass of that substance which combine with or displace directly or indirectly 1·008 parts by mass of hydrogen or 8 parts by mass of oxygen or 35·5 parts by mass of chlorine.

5. Calculation of equivalent mass

(*i*) Eq. mass of an element $= \dfrac{\text{Atomic mass of the element}}{\text{Valency of the element}}$

e.g. Eq. mass of Cu in CuO $= \dfrac{63 \cdot 6}{2} = 31 \cdot 8$ (Valency of Cu=2)

Eq. mass of Cu in $Cu_2O = \dfrac{63 \cdot 6}{1} = 63 \cdot 6$ (Valency of Cu=1)

(*ii*) Eq. mass of an ion $= \dfrac{\text{Formula mass of the ion}}{\text{Charge on the ion}}$

e.g. Eq. mass of CO_3^{2-} ion $= \dfrac{12 + 3 \times 16}{2} = \dfrac{60}{2} = 30$

(*iii*) Eq. mass of an oxidizing/reducing agent $= \dfrac{\text{Molecular mass}}{\text{No. of electrons lost or gained by one molecule}}$

(*Discussed in unit 9*)

(*iv*) Eq. mass of an acid $= \dfrac{\text{Molecular mass of the acid}}{\text{Basicity of the acid}}$ (*Already discussed*)

(*v*) Eq. mass of a base $= \dfrac{\text{Molecular mass of the base}}{\text{Acidity of the base}}$ (*Already discussed*)

6. Determination of equivalent mass

(*i*) *Hydrogen displacement method*. Calculate the mass of the metal which displaces 1·008 parts by mass of hydrogen.

(*ii*) *Oxide formation or reduction of the oxide method*. Calculate the mass of the metal which combines with or displaces 8 parts by mass of oxygen.

(*iii*) *Chloride formation method*. Calculate the mass of the element which combines with or displaces 35·5 parts by mass of chlorine.

(*iv*) *Metal displacement method*. When a more electropositive metal is added to displace a less electropositive metal from its salt (*e.g.* $Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$),

$$\frac{\text{Mass of metal added}}{\text{Mass of metal displaced}} = \frac{\text{Eq. mass of metal added}}{\text{Eq. mass of metal displaced}}$$

(*v*) *Double decomposition method*. For a reaction of the type

$$AB + CB \longrightarrow AD \downarrow + BC$$

$$(e.g.\ AgNO_3 + NaCl \longrightarrow AgCl \downarrow + NaNO_3)$$

$$\frac{\text{Mass of AB taken}}{\text{Mass of AD formed}} = \frac{\text{Eq. mass of AB}}{\text{Eq. mass of AD}} = \frac{\text{Eq. mass of A+Eq. mass of B}}{\text{Eq. mass of A+Eq mass of D}}$$

(*vi*) *Electrolytic method*. (Faraday's second law). When the same quantity of electricity flows through solutions of different electrolytes, $\dfrac{\text{Mass of X deposited}}{\text{Mass of Y deposited}} = \dfrac{\text{Eq. mass of X}}{\text{Eq. mass of Y}}$

(*vii*) *Neutralization method*. To calculate the equivalent mass of an acid, calculate the mass of the acid neutralized by 1000 cc of 1N base solution (which contains 1g eq). Likewise equivalent mass of a base can be calculated.

(*viii*) *Silver salt method* (for organic acids only). A known mass of silver salt of the organic acid is ignited to give a residue of Ag, then $\dfrac{\text{Eq. mass of RCOOAg}}{\text{Eq. mass of Ag (108)}} = \dfrac{\text{Mass of silver salt}}{\text{Mass of silver}}$

Eq. mass of acid (RCOOH) = Eq. mass of RCOOAg—107.

ADDITIONAL USEFUL INFORMATION contd.

(*ix*) *Conversion method.* When one compound of a metal is converted into another compound of the same metal (*e.g.* metal carbonate ⟶ metal oxide), then

$$\frac{\text{Mass of compound I}}{\text{Mass of compound II}} = \frac{\text{Eq. mass of metal + Eq. mass of anion of compound I}}{\text{Eq. mass of metal + Eq. mass of anion of compound II}}$$

7. Methods of determining atomic mass

(*i*) *By application of the relation* At. mass = Eq. mass × Valency

Knowing approx. atomic mass and exact equivalent mass, first valency is calculated and then the exact atomic mass.

(*ii*) *Dulong and Petit's method.* For solid elements (except Be, B, C and Si), according to *Dulong and Petit's law*

At. mass × Specific heat = 6·4 approx.

$$\therefore \quad \text{Approx. atomic mass} = \frac{6 \cdot 4}{\text{Sp. heat}}$$

$$\text{Exact atomic mass} = \text{Eq. wt.} \times \text{Valency}$$

where $\text{Valency} = \dfrac{\text{Approx. atomic mass}}{\text{Eq. mass}}$ (Take nearest whole no.)

(*iii*) *From ratio of heat capacities.* For gases, we have

Ratio, $\gamma = C_p / C_v$	*1·66*	*1·40*	*1·30*	*1·26*
Atomicity	1	2	3	4

Molecular mass can be found from vapour density (Mol. mass = 2 × V.D)

$$\therefore \quad \text{Atomic mass of the gaseous element} = \frac{\text{Mol. mass}}{\text{Atomicity}}$$

(*iv*) *Volatile chloride method.* (From vapour density measurements for elements forming volatile chlorides)

Molecular mass of the chloride = 2 × V.D.

If x is he valency of the element (M), then formula of its chloride will be MCl_x. Hence

$$\text{Mol. mass of the chloride } MCl_x = \text{At. mass of M} + x \times 35 \cdot 5$$

$$= \text{Eq. mass of M} \times \text{Valency of M} + x \times 35 \cdot 5$$

$$= E \times x + x \times 35 \cdot 5 = x(E + 35 \cdot 5)$$

$$\therefore \quad x(E + 35 \cdot 5) = 2 \times V.D.$$

$$\text{or} \quad x = \frac{2 \times V.D}{E + 35 \cdot 5}$$

Knowing the equivalent mass E of the element, the value of x can be calculated. Then atomic mass = Eq. mass × Valency.

(*v*) *Isomorphism method.* Compounds having similar molecular formulae and identical crystal structure are called isomorphous. The method is based upon the fact that elements in isomorphous compounds have same valencies, *e.g.*

(*a*) K_2SO_4, K_2CrO_4, K_2SeO_4 are isomorphous. Hence valency of S, Cr and Se = 6.

(*b*) $ZnSO_4.7\,H_2O$, $FeSO_4.7\,H_2O$, $MgSO_4.7H_2O$ are isomorphous. Hence valency of Zn, Fe, Mg=2.

Knowing valency, atomic mass = Eq. mass × Valency.

(*vi*) *From percentage of an element in a given compound.* For example, potassium chromate (K_2CrO_4) is found to contain 26·78% of Cr. Its atomic mass can be calculated as follows :

Suppose atomic mass of Cr = a

$$\therefore \quad \text{Mol. mass of } K_2CrO_4 = 2 \times 39 + a + 4 \times 16 = 142 + a$$

$$\therefore \quad \% \text{ of Cr in } K_2CrO_4 = \frac{a \times 100}{142 + a} = 26 \cdot 78$$

which on solving gives $a = 52 \cdot 0$.

ADDITIONAL USEFUL INFORMATION contd.

8. Determination of Avogadro's number. The simplest method is by studying electrolysis of acidulated water. The reaction taking place at the cathode is $2H^+ + 2e^- \longrightarrow H_2$

This means that charge carried by 2 moles of electrons produces 1 mole *i.e.* 2·016 g of H_2 gas.

Thus experimentally the amount of electricity required to produce 2·016 g of H_2 is determined. It is found to be 193000 coulombs. As charge carried by one electron = $1 \cdot 602 \times 10^{-19}$ coulombs, therefore if N is the number of electrons present in 1 mole, then

$$2\,N \times 1 \cdot 602 \times 10^{-19} = 193000 \quad \text{or} \quad N = \frac{193000}{2 \times 1 \cdot 602 \times 10^{-19}} = 6 \cdot 02 \times 10^{23}.$$

C.B.S.E.-P.M.T. (MAINS) SPECIAL

A. SUBJECTIVE QUESTIONS

Q. 1. What is the difference between the following ?

(i) $2 \cdot 5 \times 10^3$ g and $2 \cdot 50 \times 10^3$ g

(ii) 160 cm and 160·0 cm.

Ans. (*i*) $2 \cdot 5 \times 10^3$ has two significant figures while $2 \cdot 50 \times 10^3$ has three significant figures. Hence $2 \cdot 50 \times 10^3$ represents greater accuracy than $2 \cdot 5 \times 10^3$.

(*ii*) 160 has three significant figures while 160·0 has four significant figures. Hence 160·0 represents greater accuracy.

Q. 2. Which of the following has largest number of oxygen atoms ?

1·0 g of O atoms, 1·0 g of O_2, 1·0 g of ozone (O_3)

Justify your answer.

Ans. 1·0 g of O atoms = $\frac{1}{16}$ g atoms of O

$$= \frac{1}{16} \times 6 \cdot 02 \times 10^{23} \text{ atoms}$$

$$= \frac{6 \cdot 02 \times 10^{23}}{16} \text{ atoms}$$

1·0 g of O_2 = $\frac{1}{32}$ mol of O_2 = $\frac{1}{32} \times 6 \cdot 02 \times 10^{23}$ molecules = $2 \times \frac{1}{32} \times 6 \cdot 02 \times 10^{23}$ atoms

$$= \frac{6 \cdot 02 \times 10^{23}}{16} \text{ atoms}$$

1·0 g of O_3 = $\frac{1}{48}$ mol of O_3 = $\frac{1}{48} \times 6 \cdot 02 \times 10^{23}$ molecules = $3 \times \frac{1}{48} \times 6 \cdot 02 \times 10^{23}$ atoms

$$= \frac{6 \cdot 02 \times 10^{23}}{16} \text{ atoms}$$

Thus, all of them contain equal number of atoms.

Q. 3. What is the difference in the molar volume of a gas if S.T.P. conditions are

(i) 1 atm, 0°C (ii) 1 bar, 0°C ?

Ans. (*i*) If S.T.P. conditions are 1 atm, 0°C, molar volume = 22400 cm^3

(*ii*) If S.T.P. conditions are 1 bar, 0°C, molar volume = 22700 cm^3

Q. 4. Why molality is preferred over molarity in expressing the concentration of a solution ?

Ans. Molality is the number of moles of the solute present in 1 kg of the solvent whereas molarity is the number of moles of the solute present per litre of the solution. Thus, molality involves only masses which do not change with temperature whereas molarity involves volume which changes with temperature. Hence molality is preferred over molarity.

Q. 5. Taking N_2 and O_2 as main components of the air (79% N_2, 21% O_2 by volume) what is the molecular mass of air ? How has it been arrived at ?

Ans. For a mixture of gases, the average molecular mass is taken.

Average molecular mass $= \Sigma x_i M_i = x_{N_2} M_{N_2} + x_{O_2} M_{O_2}$ where x_{N_2} and x_{O_2} are mole fractions of N_2 and O_2 and M_{N_2} and M_{O_2} are their molecular masses. As equal volumes contain equal number of moles, therefore, their molar ratio is same as ratio of their volumes.

$\therefore x_{N_2} = 0 \cdot 79$, $x_{O_2} = 0 \cdot 21$. Also $M_{N_2} = 28\,u$, $M_{O_2} = 32\,u$

$\therefore$ Average molecular mass

$$= 0 \cdot 79 \times 28 + 0 \cdot 21 \times 32\,u$$

$$= 22 \cdot 12 + 6 \cdot 72\,u = \mathbf{28 \cdot 84\,u}$$

Q. 6. What is the difference between the mass of a molecule and gram molecular mass ?

Ans. Mass of a molecule is the actual mass of a single molecule expressed in grams whereas gram molecular mass is the mass in gram of Avogadro's number of molecules.

Q. 7. In the combustion of methane, what is the limiting reactant and why ?

Ans. Methane is the limiting reactant because the other reactant is oxygen of the air which is always present in excess. Thus, the amounts of CO_2 and H_2O formed will depend upon the amount of CH_4 burnt.

Q. 8. What is the equivalent weight of $KH(IO_3)_2$ as an oxidant in presence of 4·0 (N) HCl when ICl becomes the reduced form ?

(K=39·0, I=127·0). ***(West Bengal J.E.E. 2001)***

Ans. In $KH(IO_3)_2$, IO_3 is present as IO_3^-. Oxidation state of I will be

$$x - 6 = -1$$

or $$x = +5.$$

$$\overset{+5}{KH(IO_3)_2} \longrightarrow \overset{+1}{2\,ICl}$$

Decrease in oxidation state = +10—2 = 8

∴ Equivalent wt. of $KH(IO_3)_2$

$$= \frac{\text{Mol. wt.}}{8} = \frac{39 + 1 + 2\,(127 + 48)}{8} = 48{\cdot}75$$

Q. 9. What is kg-mole ? Find out the total number of electrons in a kg- mole of O_2.

(West Bengal J.E.E. 2004)

Ans. One kg-mole (or simply called one kilomole (kmol) is the molecular mass of the substance expressed in kilograms. In CGS system, one gram mole of a substance contains Avogadro's number of particles ($6{\cdot}022 \times 10^{23}$). In MKS system, one kg-mole of a substance contains Avogadro's number of particles but its value is $6 \cdot 022 \times 10^{26}$. Thus 1 kg-mole of O_2 will contain $6{\cdot}022 \times 10^{26}$ molecules and hence $16 \times 6{\cdot}022 \times 10^{26}$ electrons.

B. PROBLEMS

Problem 1. **Calculate the percentage of the naturally occurring isotopes ^{35}Cl and ^{37}Cl that accounts for the atomic mass of chlorine taken as 35·45.**

Solution. Suppose ^{35}Cl present $= x\%$. Then ^{37}Cl present $= (100 - x)\%$

∴ Average atomic mass

$$= \frac{x \times 35 + (100 - x) \times 37}{100} = 35{\cdot}45 \quad (Given)$$

or $35x + 3700 - 37x = 3545$

or $2x = 155$ or $x = 77{\cdot}5\%$

This $^{35}Cl = 77{\cdot}5\%$ and $^{37}Cl = 100 - 77{\cdot}5 = 22{\cdot}5\%$.

Problem 2. **Calculate the temperature at which the value in degrees celsius and degrees Fahernheit is same.**

Solution. The reactionship between 0°C and °F is

$$°C = \frac{5}{9}(F - 32)$$

If $x\,°C = x\,°F$, then $x = \frac{5}{9}(x - 32)$

or $9x = 5x - 160$ or $4x = -160$

or $x = -40°C$

Thus $-40°C = -40°F$.

Problem 3. **Convert 22·4 L into cubic metres.**

Solution. $$22{\cdot}4\,L = 22{\cdot}4\ L \times \frac{1000\ \cancel{cm^3}}{1\ L} \times \frac{1\ m}{100\ \cancel{cm}} \times \frac{1\ m}{100\ \cancel{cm}} \times \frac{1\ m}{100\ \cancel{cm}}$$

$$= \frac{22{\cdot}4 \times 1000}{100 \times 100 \times 100}\,m^3 = 22{\cdot}4 \times 10^{-3}\,m^3.$$

Problem 4. **Calculate the molar mass of water if it contains 50% heavy water (D_2O).**

Solution. As water contains 50% D_2O, this means that it contains $\frac{1}{2}$ mole of H_2O and $\frac{1}{2}$ mole of D_2O. Mass of $\frac{1}{2}$ mole of $H_2O = \frac{1}{2} \times 18 = 9\,g$. Mass of $\frac{1}{2}$ mole of $D_2O = \frac{1}{2} \times (2 \times 2 + 16) = 10\,g$. Hence molar mass of the given sample of water $= 9 + 10 = 19\ g\ mol^{-1}$.

Problem 5. **The average molar mass of a mixture of methane (CH_4) and ethene (C_2H_4) present in the ratio of a : b is found to be 20·0 $g\ mol^{-1}$. If the ratio were reversed, what would be the molar mass of the mixture ?**

Solution. Molar mass of $CH_4 = 16\,g\,mol^{-1}$

Molar mass of $C_2H_4 = 28\,g\,mol^{-1}$

When they are present in the ratio $a : b$, their average molar mass

$$= \frac{a \times 16 + b \times 28}{a + b} = 20\,g\,mol^{-1} \quad \text{(Given)}$$

i.e. $16a + 28b = 20(a + b)$

or $4a + 7b = 5(a + b)$ or $a = 2b$

or $$\frac{a}{b} = \frac{2}{1} = 2 : 1$$

If the ratio is reversed, now the ratio $a : b = 1 : 2$

∴ Average molar mass $= \frac{1 \times 16 + 2 \times 28}{1 + 2}$

$$= \frac{16+56}{3} = \frac{72}{3} = 24\ g\ mol^{-1}$$

Problem 6. **20·0 mL of a mixture of oxygen (O_2) and ozone (O_3) was heated till ozone was completely decomposed. The mixture on cooling was found to expand to 21 mL. Calculate the percentage of ozone by volume in the mixture.**

Solution. Decomposition of ozone takes place as follows :

$$\underset{2\text{ volume}}{2\,O_3} \longrightarrow \underset{3\text{ volume}}{3\,O_2}$$

Suppose ozone in the mixture = x mL. Then $O_2 = (20 - x)\,mL$

2 mL of ozone on decomposition give O_2 = 3 mL

∴ x mL of ozone on decomposition will give $O_2 = \frac{3}{2}x = 1{\cdot}5\,x\,mL$

∴ Total volume of mixture after decomposition

$$= 1{\cdot}5\,x + (20 - x)\,mL$$
$$= 20 + 0{\cdot}5\,x$$

∴ $20 + 0{\cdot}5\,x = 21$ *(Given)*

or $0{\cdot}5\,x = 1$ or $x = 2\,mL$

∴ Percentage of ozone in the mixture

$$= \frac{2}{20} \times 100 = \mathbf{10\%}.$$

Problem 7. **Calculate the atomicity of mercury molecules from the following data :**

(a) 10·0 g of mercury combine with 0·8 g of oxygen to form an oxide.

(b) 500 mL of mercury vapour at S.T.P. weigh = 4·465 g

(c) Specific heat of mercury is 0·033.

Solution. Calculation of equivalent mass

0·89 g of oxygen combine with Hg = 10·0 g

∴ 8 g of oxygen will combine with Hg

$$= \frac{10}{0{\cdot}8} \times 8 = 100\,g$$

∴ Equivalent mass of mercury = 100

Calculation of molar mass

500 mL of mercury vapour at S.T.P. weigh = 4·465 g

∴ 22400 mL of mercury vapour at S.T.P. will weigh

$$= \frac{4{\cdot}465}{500} \times 22400\,g$$
$$= 200\,g$$

∴ Molar mass of mercury = 200 g mol^{-1}

Calculation of valency. By Dulong and Petit's law

$$\text{Approx. atomic mass of mercury} = \frac{6{\cdot}4}{\text{Sp. Heat}}$$

$$= \frac{6{\cdot}4}{0{\cdot}033} = 193{\cdot}9$$

$$\therefore\ \text{Valency of mercury} = \frac{\text{Approx. atomic mass}}{\text{Valency}}$$

$$= \frac{193{\cdot}9}{100} \approx 2 \text{ (as valency is a whole no.)}$$

∴ Actual atomic mass = Eq. mass × Valency

$$= 100 \times 2 = 200$$

Calculation of atomicity

$$\text{Atomicity} = \frac{\text{Mol. mass}}{\text{At. mass}} = \frac{200}{200} = 1$$

Thus, mercury molecules are **monoatomic.**

Problem 8. **Insulin contains 3·4% sulphur. Calculate the minimum molecular mass of insulin.**

Solution. Minimum molecular mass of insulin will be the mass containing at least one atom of sulphur.

One atom of S = 32 a.m.u.

Now, 3·4 a.m.u. of sulphur is present in 100 a.m.u. of insulin.

∴ 32 a.m.u. of sulphur will be present in insulin

$$= \frac{100}{3{\cdot}4} \times 32$$
$$= 941{\cdot}2 \text{ a.m.u.}$$

Hence minimum molecular mass of insulin **= 941·2 a.m.u.**

Problem 9. **How many molecules approximately do you expect to be present in (*i*) a small sugar crystal which weighs 10 mg (*ii*) one drop of water with 0·05 cc volume ?**

Solution. (*i*) 10 mg sugar ($C_{12}H_{22}O_{11}$)

$$= 0{\cdot}01\,g = \frac{0{\cdot}01}{342} \text{ mole} = 2{\cdot}92 \times 10^{-5} \text{ mole}$$

$$= (2{\cdot}92 \times 10^{-5}) \times (6{\cdot}02 \times 10^{23})$$

$$= \mathbf{1{\cdot}76 \times 10^{19}\ molecules.}$$

(*ii*) 0·05 cc water = 0·05 g = $\frac{0{\cdot}05}{18}$ mole

$$= 2{\cdot}78 \times 10^{-3} \text{ mole}$$
$$= (2{\cdot}78 \times 10^{-3}) \times (6{\cdot}02 \times 10^{23}) \text{ molecules}$$
$$= \mathbf{1{\cdot}67 \times 10^{21}\ molecules.}$$

Problem 10. **$9{\cdot}7 \times 10^{17}$ atoms of iron weigh as much as 1 cc of H_2 at S.T.P. What is the atomic mass of iron ?**

Solution. Mass of 1 cc of H_2 at S.T.P.

$$= \frac{2{\cdot}016}{22400} = 9{\cdot}0 \times 10^{-5}\,g$$

Mass of $6{\cdot}02 \times 10^{23}$ atoms of Fe

$$= \frac{9{\cdot}0 \times 10^{-5} \times 6{\cdot}02 \times 10^{23}}{9{\cdot}7 \times 10^{17}} = 55{\cdot}9\,g$$

∴ Atomic mass of Fe = **55·9 u**

Problem 11. **20 cc of 1 N HCl, 30 cc of 0·5 N H_2SO_4 and 50 cc of 0·2 N HNO_3 are mixed together. What will be the normality of the final solution ?**

Solution. 20 cc of 1 N HCl contains

$$= \frac{1}{1000} \times 20 = 0 \cdot 020 \text{ g eq.}$$

30 cc of 0·5 N H_2SO_4 contains

$$= \frac{0 \cdot 5}{1000} \times 30 = 0 \cdot 015 \text{ g eq}$$

50 cc of 0·2 N HNO_3 contains

$$= \frac{0 \cdot 2}{1000} \times 50 = 0 \cdot 01 \text{ g eq.}$$

Total no. of gram equivalents

$= 0 \cdot 02 + 0 \cdot 015 + 0 \cdot 01$

$= 0 \cdot 045$

Total volume = 20 + 30 + 50 = 100 cc

∴ Normality of the final solution

$$= \frac{0 \cdot 045}{100} \times 1000 = 0 \cdot 45 \text{ N}$$

Problem 12. **Crystalline magnesium sulphate on heating becomes anhydrous and loses 51·2% of its mass. Calculate the number of water molecules present.**

Solution. 100 parts by mass of the hydrated salt contain 51·2 parts by mass of water. This means that (100—51·2) = 48·8 parts by mass of the anhydrous salt are associated with water = 51·2 parts by mass.

Molecular mass of anhydrous $MgSO_4$

$= 24 + 32 + 64 = 120$

∴ 120 parts of anhydrous $MgSO_4$ will be associated

with $H_2O = \frac{51 \cdot 2}{48 \cdot 8} \times 120 = 125 \cdot 9$ parts

$$= \frac{125 \cdot 9}{18} \text{ molecules} = 7 \text{ molecules.}$$

Problem 13. **1·615 g of anhydrous $ZnSO_4$ was placed in moist air. After a few days its mass was found to be 2·875 g. What is the molecular formula of the hydrated salt ? (At. masses : Zn = 65·5, S = 32, O = 16).**

Solution. Molecular mass of anhydrous $ZnSO_4$

$= 65 \cdot 5 + 32 + 64 = 161 \cdot 5$

1·615 g anhydrous $ZnSO_4$ combine with H_2O

$= 2 \cdot 875 - 1 \cdot 615 = 1 \cdot 260$ g

∴ 161·5 g anhydrous $ZnSO_4$ will combine with

$H_2O = 126 \text{ g} = \frac{126}{18} = 7$ moles

Hence the formula of the hydrated salt is $ZnSO_4 . 7H_2O$.

Problem 14. **4 g carbon were heated with 8 g of sulphur. How much carbon disulphide (CS_2) will be formed when the reaction is complete. What will be its percentage purity ?**

Solution.

$$\underset{12\text{ g}}{C} + \underset{\substack{2 \times 32 \\ = 64\text{ g}}}{2\,S} \longrightarrow \underset{\substack{12 + 2 \times 32 \\ = 76\text{ g}}}{CS_2}$$

Obviously sulphur will be the limiting reactant.

8 g sulphur will produce $CS_2 = \frac{76}{64} \times 8 = 9 \cdot 5$ g

Carbon reacted $= \frac{12}{64} \times 8 = 1 \cdot 5$ g

Carbon left = 4 — 1·5 = 2·5 g

Total mass of products = 9·5 + 2·5

= 12 g

∴ % purity of CS_2 in the product $= \frac{9 \cdot 5}{12} \times 100$

= 79·2%.

Problem 15. **How many years would it take to spend Avogadro's number of rupees at the rate of 10 lakh rupees per second ?** ***(M.L.N.R. Allahabad 1990)***

Solution. $\frac{6 \cdot 02 \times 10^{23}}{10,00,000} s = 6 \cdot 02 \times 10^{17} s$

$$= \frac{6 \cdot 02 \times 10^{17}}{60 \times 60 \times 24 \times 365} = 1 \cdot 9089 \times 10^{10} \text{ years}$$

Problem 16. **Gastric juice contains about 3·0 *g* of HCl per litre. If a person produces about 2·5 litre of gastric juice per day, how many antacid tablets each containing 400 mg of $Al(OH)_3$ are needed to neutralize all the HCl produced in one day.** ***(I.S.M. Dhanbad 1991)***

Solution. HCl is neutralized by $Al(OH)_3$ according to the equation

$$\underset{\substack{27 + 3 \times 17 \\ = 78\,g}}{Al(OH)_3} + \underset{\substack{3 \times 36 \cdot 5 \\ = 109 \cdot 5\,g}}{3\,HCl} \longrightarrow AlCl_3 + 3\,H_2O$$

HCl produced from 2·5 litre of gastric juice $= 3 \cdot 0 \times 2 \cdot 5 = 7 \cdot 5\,g$

109·5 *g* HCl is neutralized by $Al(OH)_3 = 78\,g$

∴ 7·5 *g* HCl will be neutralized by $Al(OH)_3$

$= \frac{78}{109 \cdot 5} \times 7 \cdot 5\,g = 5 \cdot 342\,g = 5342$ mg

∴ Number of tablets required $= \frac{5342}{400} = 13 \cdot 35$ *i.e.* **14 tablets.**

Problem 17. **What volume at N.T.P. of ammonia gas will be required to be passed into 30 mL of N H_2SO_4 solution to bring down the acid normality to 0·2 N ?** ***(M.L.N.R. 1991)***

Solution. Normality is the number of milliequivalents (Meq) per *m*L

$\therefore$ Meq of 1 N H_2SO_4 present originally in 30 mL $= 30 \times 1 = 30$

Meq of 0·2 N H_2SO_4 in 30 mL $= 30 \times 0 \cdot 2 = 6$

$\therefore$ Meq of H_2SO_4 required to be neutralized by $NH_3 = 30 - 6 = 24$

$\therefore$ Meq of NH_3 required to be passed = 24

($\because$ Acid and base react with each other in equivalent amounts)

But 1 mole of NH_3 = 1 g eq *i.e.* 1000 Meq of NH_3 = 22400 mL at N.T.P.

$\therefore$ 24 Meq of $NH_3 = \frac{22400}{1000} \times 24 = \mathbf{537 \cdot 6\ mL}$

Problem 18. **Chlorine has two isotopes of atomic mass units 34·97 and 36·97. The relative abundances of these two isotopes are 0·755 and 0·245 respectively. Find the average atomic mass of chlorine.**

(B.I.T. Ranchi 1991)

Solution. Average atomic mass

$$= \frac{34 \cdot 97 \times 0 \cdot 755 + 36 \cdot 97 \times 0 \cdot 245}{0 \cdot 755 + 0 \cdot 245} = \mathbf{35 \cdot 46}$$

Problem 19. **A gas mixture of 3·0 litres of propane and butane on complete combustion at 25°C produced 10 litres of CO_2. Find out the composition of the gas mixture.** ***(M.L.N.R. Allahabad 1992)***

Solution. $C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$

$C_4H_{10} + 6\frac{1}{2}O_2 \longrightarrow 4CO_2 + 5H_2O$

Suppose propane = x L.

Then butane = $(3 - x)$ L

1 L C_3H_8 gives 3L of CO_2 and 1 L C_4H_{10} gives 4 L of CO_2.

Hence CO_2 produced $= 3x + 4(3 - x)$

$= 12 - x$

$\therefore\ 12 - x = 10$ *(Given)*

Hence $x = 2$ L

i.e. propane = **2L** and butane = **1L**.

Problem 20. **The vapour density of a mixture of NO_2 and N_2O_4 is 38·3 at 26·7°C. Calculate the number of moles of NO_2 in 100 g of the mixture.**

(M.L.N.R. 1993)

Solution. Suppose NO_2 present in 100 g of the mixture = x g

Then N_2O_4 present in the mixture $= (100 - x)\,g$

Molar mass of $NO_2 = 14 + 32 = 46\ g\ mol^{-1}$

Molar mass of $N_2O_4 = 92\ g\ mol^{-1}$

Molar mass of mixture $= 2 \times$ V.D.

$= 2 \times 38 \cdot 3 = 76 \cdot 6\ g\ mol^{-1}$

Expressing all quantities in terms of moles, we should have

$$\frac{x}{46} + \frac{100 - x}{92} = \frac{100}{76 \cdot 6}$$

or $92x + 4600 - 46x = 5524 \cdot 8$

or $46x = 924 \cdot 8$ or $x = 20 \cdot 1\ g$

$\therefore$ No. of moles of NO_2 in the mixture $= \frac{20 \cdot 1}{46} = \mathbf{0 \cdot 437}$

Problem 21. ***Naturally occurring Boron consists of two isotopes whose atomic weights are 10·01 and 11·01. The atomic weight of natural Boron is 10·81. Calculate the percentage of each isotope in natural Boron.*** ***(M.L.N.R. Allahabad 1994)***

Solution. Suppose the percentage of isotope with atomic weight $10 \cdot 01 = x$

Then percentage of isotope with atomic weight $11 \cdot 01 = 100 - x$

$\therefore$ Average atomic weight

$$= \frac{10 \cdot 01x + (100 - x) \times 11 \cdot 01}{100}$$

$$= \frac{10 \cdot 01x + 1101 - 11 \cdot 01x}{100} = \frac{1101 - x}{100}$$

$\therefore\ \frac{1101 - x}{100} = 10 \cdot 81$ or $1101 - x = 1081$

or $x = 1101 - 1081 = 20$

$\therefore$ % age of isotope with atomic weight 10·01

$= \mathbf{20\ \%}$

% age of isotope with atomic weight 11·01 = **80 %**

Problem 22. **The mass of one litre sample of ozonised oxygen at N.T.P. was found to be 1·5 g. When 100 mL of this mixture at N.T.P. were treated with turpentine oil, the volume was reduced to 90 mL. Calculate the molecular mass of ozone.** ***(M.L.N.R. 1996)***

Solution. As ozone is absorbed by turpentine oil, therefore volume of ozone in 100 mL of the mixture $= 100 - 90 = 10\ mL$.

$\therefore$ O_2 in the mixture $= 100 - 10 = 90\ mL$

As 1 L of the mixture weigh = 1·5 g, therefore, average molar mass of the mixture (mass of 22·4 L at S.T.P.) $= 1 \cdot 5 \times 22 \cdot 4 = 33 \cdot 6\ g\ mol^{-1}$

Ratio of ozone : oxygen in the mixture = 10 : 90

If m is the molecular mass of ozone, then

Average mol. mass of mixture

$$= \frac{10 \times m + 90 \times 32}{100} = 33 \cdot 6 \quad \textit{(calculated above)}$$

or $m + 288 = 336$ or $m = \mathbf{48}$

Problem 23. **One gram of a metal hydroxide on heating gives 5·6 g of metal oxide. Calculate the equivalent weight of the metal.** ***(Bihar C.E.C.E. 2000)***

Solution. $\frac{\text{Eq. wt. of metal hydroxide}}{\text{Eq. wt. of metal oxide}} = \frac{\text{Wt. of metal hydroxide}}{\text{Wt. of metal oxide}}$

$$\frac{E + 17}{E + 8} = \frac{7 \cdot 4}{5 \cdot 6}$$

or $(E + 17) \times 5 \cdot 6 = (E + 8) \times 7 \cdot 4$

or $1 \cdot 8\, E = 36$ or $E = 20.$

Problem 24. **Calculate the molarity of water in pure water.** ***(Bihar C.E.C.E. 2000)***

Solution. 1 L of pure water = 1000 cm^3 = 1000 *g* (assuming density = 1 $g\ cm^{-3}$)

$\therefore$ No. of moles in 1 L of pure water $= \frac{1000}{18} = 55 \cdot 55.$

Problem 25. **How many molecules are there in 10 litres of a gas at a pressure of 75 cm at 27 degree celsius ?** ***(Bihar C.E.C.E. 2000)***

Solution. PV = *n* RT.

$$n = \frac{PV}{RT} = \frac{\frac{750}{760} \times 10}{0 \cdot 0821 \times 300} = 0 \cdot 4 \text{ mole}$$

No. of molecules present in 0·4 mole

$= (6 \cdot 023 \times 10^{23}) \times 0 \cdot 4 = 2 \cdot 409 \times 10^{23}.$

Problem 26. **Potassium bromide, KBr contains 32·9% by mass of potassium. If 6·40 g of bromine reacts with 3·60 g of potassium, calculate the number of moles of potassium which combine with bromine to form KBr.** ***(N.C.E.R.T.)***

Solution. In KBr, 32·9 *g* of K are combined with 67·1 *g* of Br

$\therefore$ 3·6 *g* of K will combine with Br_2

$= \frac{67 \cdot 1}{32 \cdot 9} \times 3 \cdot 6 = 7 \cdot 34\ g$

or 6·4 *g* or Br_2 will combine with K $= \frac{32 \cdot 9}{67 \cdot 1} \times 6 \cdot 4 = 3 \cdot 14\ g$

Thus Br_2 is the limiting reactant

K reacted = 3·14 *g*

$\therefore$ KBr formed $= \frac{100}{32 \cdot 9} \times 3 \cdot 14 = 9 \cdot 54\ g$

$= \frac{9 \cdot 54}{39 + 80}$ mol $= \frac{9 \cdot 54}{119}$ mol $= 0 \cdot 0802$ mol

Problem 27. **Two acids H_2SO_4 and H_3PO_4 are neutralized separately by the same amount of an alkali when sulphate and dihydrogen orthophosphate are formed respectively. Find the ratio of the masses of H_2SO_4 and H_3PO_4 [P = 31].** ***(West Bengal J.E.E. 2003)***

Solution. 1 *g* eq. of alkali (NaOH) will neutralize 1 *g* eq. of H_2SO_4 and 1 *g* eq. of H_3PO_4. For the given neutralization reactions

$(H_2SO_4 + 2\,NaOH \longrightarrow Na_2SO_4 + 2\,H_2O,$

$H_3PO_4 + NaOH \longrightarrow NaH_2PO_4 + H_2O),$

equivalent mass of H_2SO_4

$= \frac{\text{Mol. mass}}{2} = \frac{98}{2} = 49$

and equivalent mass of $H_3PO_4 = \frac{\text{Mol. mass}}{1} = 98.$

Hence the ratio of masses of H_2SO_4 and H_3PO_4 = 49 : 48 = 1 : 2.

Problem 28. **10 ml of HCl solution gave 0·1435 *g* of AgCl when treated with excess of $AgNO_3$. Find the normality of the acid solution [Ag = 108]** ***(West Bengal J.E.E. 2003)***

Solution. Cl present in 0·1435 *g* AgCl

$= \frac{35 \cdot 5}{143 \cdot 5} \times 0 \cdot 1435\ g = 0 \cdot 0355\ g$

HCl containing 0·0355 *g* Cl

$= \frac{36 \cdot 5}{35 \cdot 5} \times 0 \cdot 0355\ g = 0 \cdot 0365\ g = \frac{0 \cdot 0365}{36 \cdot 5}\ g$ eq.

$= 0 \cdot 001\ g$ eq.

i.e. 10 ml of HCl contain 0·001 *g* eq.

$\therefore$ 1000 ml of HCl will contain = 0·1 *g* eq. *i.e.* normality = **0·1 N**

Problem 29. **To account for atomic mass of nitrogen as 14·0067, what should be the ratio of ^{15}N and ^{14}N atoms in natural nitrogen ? (atomic mass of ^{14}N = 14·00307 *u* and ^{15}N = 15·001 *u*)** ***(Bihar C.E.C.E. 2003)***

Solution. Suppose %age of $^{14}N = x$.

Then %age of $^{15}N = 100 - x$

$\therefore$ According to the definition of average atomic mass,

$$14 \cdot 0067 = \frac{x \times 14 \cdot 00307 + (100 - x) \times 15 \cdot 001}{100}$$

or $1400 \cdot 67 = 1500 \cdot 1 + x\,(14 \cdot 00307 - 15 \cdot 001)$

$= 1500 \cdot 1 - 0 \cdot 99793\,x$

or $x = \frac{99 \cdot 43}{0 \cdot 99793} = 99 \cdot 636$

i.e. %age of $^{14}N = 99 \cdot 636$

and %age of $^{15}N = 100 - 99 \cdot 636 = 0 \cdot 364$

$\therefore$ Ratio of atoms of ^{15}N and ^{14}N

$= \frac{0 \cdot 364}{99 \cdot 636} = \frac{1}{273 \cdot 73} = \mathbf{1 : 273 \cdot 73}$

I.I.T. (MAINS) SPECIAL

Problem 1. **1·0 g an alloy of aluminium and magnesium when treated with excess of dilute HCl form magnesium chloride and aluminium chloride and hydrogen collected over mercury at 0°C has a volume of 1·20 L at 0·92 atmospheric pressure. Calculate the composition of the alloy.**

Solution. Suppose Al in the alloy = x g

Then Mg in the alloy = $(1-x)$ g

Al and Mg in the alloy will react with HCl acid as follows :

$$\underset{2\times 27 = 54\,g}{2\,Al} + 6\,HCl \longrightarrow 2\,AlCl_3 + \underset{3\times 22\cdot 4\text{ L at S.T.P.}}{3\,H_2}$$

$$\underset{248}{Mg} + 2\,HCl \longrightarrow MgCl_2 + \underset{22\cdot 4\text{ L at S.T.P.}}{H_2}$$

H_2 produced from x g of Al $= \frac{3 \times 22\cdot 4}{54} \times x$

$= \frac{22\cdot 4\,x}{18}$ L at S.T.P.

H_2 produced from $(1-x)$ g of Mg

$= \frac{22\cdot 4}{24} \times (1-x)$ L at S.T.P.

$\therefore$ Total H_2 produced at S.T.P.

$= \frac{22\cdot 4\,x}{18} + \frac{22\cdot 4\,(1-x)}{24}$ L at S.T.P.

Let us now convert the actual volume of H_2 produced to volume at S.T.P.

$P_1 = 0\cdot 92$ atm $\quad P_2 = 1$ atm

$V_1 = 1\cdot 20$ L $\quad V_2 = ?$

$T_1 = 273$ K $\quad T_2 = 273$ K

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad \therefore \quad \frac{0\cdot 92 \times 1\cdot 20}{273} = \frac{1 \times V_2}{273}$$

or $\quad V_2 = 1\cdot 104$ L

$$\therefore \quad \frac{22\cdot 4\,x}{18} + \frac{22\cdot 4\,(1-x)}{m} = 1\cdot 104$$

or $\quad 4 \times 22\cdot 4\,x + 3 \times 22\cdot 4\,(1-x) = 1\cdot 104 \times 72$

or $\quad 89\cdot 6\,x + 67\cdot 2 - 67\cdot 2\,x = 79\cdot 488$

or $\quad 22\cdot 4\,x = 12\cdot 288 \quad$ or $\quad x = 0\cdot 5486$ g

$\therefore$ % of Al = 54·86

and % of Mg = 100 − 54·86 = 49·14.

Problem 2. **A solid mixture (5.0 g) consisting of lead nitrate and sodium nitrate was heated below 600°C until the weight of the residue was constant. If the loss in weight is 28·0 percent, find the amount of lead nitrate and sodium nitrate in the mixture.** ***(I.I.T. 1990)***

Solution. At 600°C lead nitrate and sodium nitrate decompose as follows :

$$2Pb(NO_3)_2 \longrightarrow 2PbO + 4NO_2 + O_2$$

$$2NaNO_3 \longrightarrow 2NaNO_2 + O_2$$

Suppose $Pb(NO_3)_2$ in the mixture = x g

Then $NaNO_3$ in the mixture = $(5-x)$ g

1 mole of $Pb(NO_3)_2$ gives 1 mole of residue of PbO

i.e. $207 + 62 \times 2 = 331$ g $Pb(NO_3)_2$ give residue of PbO $= 207 + 16 = 223$ g

$\therefore$ x g $Pb(NO_3)_2$ will give residue $= \frac{223}{331} \times x$ g

1 mole of $NaNO_3$ gives 1 mole of residue of $NaNO_2$

i.e. $23 + 14 + 48 = 85$ g $NaNO_3$ give residue of $NaNO_2 = 23 + 14 + 32 = 69$ g

$\therefore$ $(5-x)$ g $NaNO_3$ give residue $= \frac{69}{85} \times (5-x)$ g

Total residue $= \frac{223\,x}{331} + \frac{69\,(5-x)}{85}$ g

Loss in weight (given) = 28% of 5g

$= \frac{28}{100} \times 5 = 1\cdot 4$ g

$\therefore$ Residue left $= 5 - 1\cdot 4 = 3\cdot 6$ g

Hence $\frac{223\,x}{331} + \frac{69\,(5-x)}{85} = 3\cdot 6$

On solving, we get $x = 3\cdot 32$ g

$Pb(NO_3)_2 = 3\cdot 32$ g, $NaNO_3 = 5 - 3\cdot 32 = 1\cdot 68$ g.

Problem 3. **A mixture of formic acid and oxalic acid is heated with conc H_2SO_4. The gas produced is collected and on its treatment with KOH solution, the volume of the gas decreased by 1/6 th. Calculate the molar ratio of the two acids in the original mixture.** ***(Roorkee 1990)***

Solution. Suppose no. of moles of formic acid is x and that of oxalic acid is y. Then

$$\underset{x\text{ mole}}{HCOOH} \xrightarrow[H_2SO_4]{\text{Conc}} H_2O\,(l) + \underset{x\text{ mole}}{CO\,(g)}$$

$$\underset{y\text{ mole}}{\begin{array}{c}COOH\\ |\\ COOH\end{array}} \longrightarrow \underset{y\text{ mole}}{CO\,(g)} + \underset{y\text{ mole}}{CO_2\,(g)} + H_2O\,(l)$$

Total no. of moles of gaseous product $= x + 2y$

As only CO_2 is absorbed by KOH, therefore fraction of $CO_2 = \frac{y}{x+2y} = \frac{1}{6}$ or $6y = x + 2y$

or $4y = x$ or $\frac{x}{y} = \frac{4}{1}$

Hence molar ratio of HCOOH to $(COOH)_2$ is 4:1.

Problem 4. **Calculate the number of oxalic acid molecules in 100 ml of 0·02 N oxalic acid solution.** ***(Roorkee 1991)***

Solution. 100 ml of 0·02 N oxalic acid solution

$= 0{\cdot}002$ g eq $= 0{\cdot}001$ mole $= 10^{-3}$ mole

$= \mathbf{6{\cdot}02 \times 10^{20}}$ **molecules.**

Problem 5. **A 2·0 g sample of a mixture containing sodium carbonate, sodium bicarbonate and sodium sulphate is gently heated till the evolution of CO_2 ceases. The volume of CO_2 at 750 mm Hg pressure and at 298 K is measured to be 123·9 ml. A 1·5g of the same sample requires 150 ml of M/10 HCl for complete neutralization. Calculate the percentage composition of the components of the mixture.** ***(I.I.T. 1992)***

Solution. Suppose $Na_2CO_3 = x\,g$, $NaHCO_3 = y\,g$.

Then $Na_2SO_4 = 2 - (x + y)\,g$.

On heating only $NaHCO_3$ will decompose to give CO_2 as follows :

$$\underset{2\,(84)\,g}{2NaHCO_3} \longrightarrow Na_2CO_3 + H_2O + \underset{\substack{22400\text{ ml}\\ \text{at STP}}}{CO_2}$$

y g $NaHCO_3$ will give $CO_2 = \frac{22400}{168} \times y$ ml at STP

Actual CO_2 produced at STP may be calculated as follows :

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}, \frac{760 \times V_1}{273} = \frac{750 \times 123{\cdot}9}{298},$$

$V_1 = 112{\cdot}0$ ml

Hence $\frac{22400}{160}y = 112$ or $y = 0{\cdot}84\,g$

$1{\cdot}5\,g$ of the mixture requires M/10 HCl = 150 ml

$\therefore$ $2{\cdot}0\,g$ of the mixture will require M/10 HCl

$= \frac{150}{1{\cdot}5} \times 2{\cdot}0 = 200$ ml

$= 0{\cdot}02$ mole HCl

$$\underset{106\text{ g}}{Na_2CO_3} + \underset{2\text{ moles}}{2HCl} \longrightarrow 2NaCl + H_2O + CO_2$$

$$\underset{84\text{ g}}{NaHCO_3} + \underset{1\text{ mole}}{HCl} \longrightarrow NaCl + H_2O + CO_2$$

$x\,g$ Na_2CO_3 require HCl $= \frac{2}{106} \times x$ moles

$0{\cdot}84\,g$ $NaHCO_3$ require HCl $= \frac{1}{84} \times 0{\cdot}84$ mole

$= 0{\cdot}01$ mole

Hence $\frac{2x}{106} + 0{\cdot}01 = 0{\cdot}02$ or $x = 0{\cdot}53\,g$

% of $Na_2CO_3 = \frac{0{\cdot}53}{2} \times 100 = \mathbf{26{\cdot}5\%}$

% of $NaHCO_3 = \frac{0{\cdot}84}{2} \times 100 = \mathbf{42{\cdot}0\%}$

% of $Na_2SO_4 = 100 - (26{\cdot}5 + 42{\cdot}0) = \mathbf{31{\cdot}5\%}$

Problem 6. **Upon mixing 45·0 ml of 0·25 M lead nitrate solution with 25·0 ml of 0·10 M chromic sulphate solution, percipitation of lead sulphate takes place. How many moles of lead sulphate are formed ? Also calculate the molar concentration of the species left behind in the final solution. Assume that lead sulphate is completely insoluble (At. wt. of Pb = 207·2).** ***(I.I.T. 1993)***

Solution.
$$\underset{3\text{ moles}}{3Pb\,(NO_3)_2} + \underset{1\text{ mole}}{Cr_2\,(SO_4)_3} \longrightarrow \underset{3\text{ moles}}{3Pb\,SO_4} + 2\,Cr\,(NO_3)_3$$

45·0 ml of 0·25 M $Pb\,(NO_3)_2$ solution

$= \frac{0{\cdot}25}{1000} \times 45$ mole of $Pb\,(NO_3)_2$

$= 0{\cdot}01125$ mole

25·0 ml of 0·10 $MCr_2(SO_4)_3$ solution

$= \frac{0{\cdot}1}{1000} \times 25$ mole of $Cr_2\,(SO_4)_3 = 0.0025$ mole

Evidently $Cr_2\,(SO_4)_3$ will be the limiting reactant.

0·0025 mole $Cr_2\,(SO_4)_3$ will react with

$Pb\,(NO_3)_2 = 3 \times 0{\cdot}0025$ mole $= 0{\cdot}0075$ mole

$\therefore$ $Pb\,(NO_3)_2$ left in the solution

$= 0{\cdot}01125 - 0{\cdot}0075$ mole $= 0{\cdot}00375$ mole

Total volume of the solution = 45 + 25 = 70 ml

$\therefore$ Molar conc of $Pb\,(NO_3)_2$ left in the sol.

$= \frac{0{\cdot}00375}{70} \times 1000$ M $= \mathbf{0{\cdot}05357}$ **M**

$PbSO_4$ formed $= 3 \times 0{\cdot}0025 = \mathbf{0{\cdot}0075}$ **mol**

Problem 7. **A mixture of ethane (C_2H_6) and ethene (C_2H_4) occupies 40 litres at 1·00 atm and at 400K. The mixture reacts completely with 130 g of O_2 to produce CO_2 and H_2O. Assuming ideal gas behaviour, calculate the mole fractions of C_2H_4 and C_2H_6 in the mixture.** ***(I.I.T. 1995)***

Solution. Applying the ideal gas equation,

$PV = n\,RT$

$1\text{ atm} \times 40\text{ L} = n \times 0{\cdot}0821\text{ L atm K}^{-1}\text{ mol}^{-1} \times 400\text{ K}$

or $n = \frac{40}{0{\cdot}0821 \times 400} = 1{\cdot}218$ mole

(where n = total no. of moles of C_2H_6 and C_2H_4 in the mixture)

Suppose the no. of moles of C_2H_6 in the mixture $= x$

Then the no. of moles of C_2H_4 in the mixture

$$= 1 \cdot 218 - x$$

Also $130 \text{ g } O_2 = \frac{130}{32}$ moles of $O_2 = 4 \cdot 0625$ moles

The reactions for complete combustion of C_2H_6 and C_2H_4 are

$$2C_2H_6 + 7\,O_2 \longrightarrow 4CO_2 + 6H_2O \quad ...(i)$$

$$\text{and } C_2H_4 + 3O_2 \longrightarrow 2CO_2 + 2H_2O \quad ...(ii)$$

From (*i*), no. of moles of O_2 required for complete combustion of x moles of $C_2H_6 = \frac{7}{2} \times x = 3 \cdot 5x$

From (*ii*), no. of moles of O_2 required for complete combustion of $(1 \cdot 218 - x)$ moles of C_2H_4

$$= 3\,(1 \cdot 218 - x)$$

Hence $3 \cdot 5x + 3\,(1 \cdot 218 - x) = 4 \cdot 0625$

or $0 \cdot 5x = 4 \cdot 0625 - 3 \cdot 654 = 0 \cdot 4085$

or $x = 0 \cdot 8170$ mole

$\therefore$ Mole fraction of $C_2H_6 = \frac{0 \cdot 817}{1 \cdot 218} = 0 \cdot 67$

and Mole fraction of $C_2H_4 = 1 - 0 \cdot 67 = 0 \cdot 33$

Problem 8. **A mixture of 20 mL of CO, CH_4 and N_2 was burnt in excess of O_2 resulting in the reduction of 13 mL of volume. The residual gas was then treated with KOH solution to show a contraction of 14 mL in volume. Calculate volume of CO, CH_4 and N_2 in the mixture. All measurements are made at constant pressure and temperature.** ***(I.I.T. 1995)***

Solution. Suppose volume of CO = a mL, $CH_4 = b$ mL and $N_2 = c$ mL. Then

$$a + b + c = 20\,mL \quad ...(i)$$

The combustion reactions will be

$$CO + \frac{1}{2}O_2 \longrightarrow CO_2$$

$$CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O\,(l)$$

$$N_2 + O_2 \longrightarrow \text{No reaction}$$

Thus, a mL of CO will produce $CO_2 = a$ mL

b mL of CH_4 will produce $CO_2 = b$ mL

N_2 will remain as such *i.e.* $= c$ mL

As CO_2 is absorbed by KOH, decrease in volume on treating with KOH will be

$$= a + b = 14\,mL\,(Given) \quad ...(ii)$$

The first given decrease is due to O_2 consumed.

a mL of CO will consume $O_2 = \frac{a}{2}$ mL

b mL of CH_4 will consume $O_2 = 2\,b$ mL.

$\therefore$ O_2 consumed $= \frac{a}{2} + 2\,b = 13$ mL (*Given*)...(*iii*)

From eqns. (*i*) and (*ii*), $c = 20 - 14 = 6$ mL

From Eqns. (*ii*) and (*iii*),

$$\frac{a}{2} + 2\,(14 - a) = 13 \text{ or } \frac{3}{2}a = 15 \text{ or } a = 10\,mL$$

$\therefore$ From eqn. (*ii*), $10 + b = 14$ or $b = 4$ mL

CO = 10 mL, CH_4 = 4 mL, N_2 = 6 mL

Problem 9. **One litre of a mixture of O_2 and O_3 at NTP was allowed to react with an excess of acidified solution of KI. The iodine liberated required 40 ml of M/10 sodium thiosulphate solution for titration. What is the weight percent of ozone in the mixture ? Ultraviolet radiations of wavelength 300 nm can decompose ozone. Assuming that one photon can decompose one ozone molecule, how many photons would have been required for the complete decomposition of ozone in the original mixture ?** ***(I.I.T. 1997)***

Solution. $O_3 \longrightarrow O_2 + O$

$$2\,KI + H_2SO_4 + (O) \longrightarrow K_2SO_4 + H_2O + I_2$$

$$2\,Na_2SO_3 + I_2 \longrightarrow Na_2S_4O_6 + 2\,NaI$$

$$2\,Na_2S_2O_3 + 2\,KI + O_3 + H_2SO_4 \longrightarrow O_2 + K_2SO_4 + H_2O + Na_2S_4O_6 + 2\,NaI$$

2 moles of $Na_2S_2O_3$ react with one mole of O_3

No. of moles of $Na_2S_2O_3$ in 40 ml of $\frac{M}{10}$ $Na_2S_2O_3$

$$= \frac{1}{10} \times \frac{1}{1000} \times 40 = 4 \times 10^{-3} \text{ mole}$$

$\therefore$ O_3 reacted (present) $= 2 \times 10^{-3}$ mole

Volume of $O_3 = 2 \times 10^{-3} \times 22 \cdot 4$ L $= 0 \cdot 0448$ L

$\therefore$ Volume of $O_2 = 1 - 0 \cdot 0448 = 0 \cdot 9552$ L

Weight of $O_3 = 2 \times 10^{-3} \times 48 = 0 \cdot 096$ g

Weight of $O_2 = \frac{0 \cdot 9552}{22 \cdot 4} \times 32 = 1 \cdot 3646$ g

$\therefore$ Weight of mixture $= 0 \cdot 096 + 1 \cdot 3646 = 1 \cdot 4606$ g

$\therefore$ % of O_3 by weight $= \frac{0 \cdot 096}{1 \cdot 4606} \times 100 = $ **6 ·573%**

No. of molecules in 2×10^{-3} mole of O_3

$$= 2 \times 10^{-3} \times 6 \cdot 022 \times 10^{23} = 1 \cdot 2044 \times 10^{21}$$

$\therefore$ No. of photons required $= $ **$1 \cdot 2044 \times 10^{21}$**

Problem 10. **1·2 g mixture of Na_2CO_3 and K_2CO_3 was dissolved in water to form 100 cm^3 of a solution. 20 cm^3 of this solution required 40 cm^3 of 0·1 N HCl for neutralisation. Calculate the weight of Na_2CO_3 and K_2CO_3 in the mixture.** ***(Roorkee 1997)***

Solution. Suppose weight of Na_2CO_3 in the mixture $= x\,g$

$\therefore$ Weight of K_2CO_3 in the mixture $= (1 \cdot 2 - x)\,g$

$$\text{Eq. wt. of } Na_2CO_3 = \frac{46+12+148}{2} = 53,$$

$$\text{Eq. wt. of } K_2CO_3 = \frac{78+12+48}{2} = 69$$

No. of g eq. of Na_2CO_3 and K_2CO_3 in the mixture

$$= \frac{x}{53} + \frac{1 \cdot 2 - x}{69}$$

40 cc of 0·1 N HCl contain g eq. of HCl

$$= \frac{0 \cdot 1}{1000} \times 40 = 4 \times 10^{-3}$$

Thus 20 cc of the mixture sol neutralize HCl

$$= 4 \times 10^{-3}\,g\text{ eq}$$

$\therefore$ 100 cc of the mixture sol. will neutralize HCl

$$= \frac{4 \times 10^{-3}}{20} \times 100 = 2 \times 10^{-2}\,g\text{ eq}$$

$$= 0 \cdot 02\,g\text{ eq.}$$

As substances react in equivalent amounts,

$$\frac{x}{53} + \frac{1 \cdot 2 - x}{69} = 0 \cdot 02$$

or $\quad 69x + 63 \cdot 6 - 53x = 0 \cdot 02 \times 53 \times 69$

$$= 73 \cdot 14$$

or $\quad 16x = 9 \cdot 54$

or $\quad x = 0 \cdot 596\,g$

Thus Na_2CO_3 = **0·596 g**

and $K_2CO_3 = 1 \cdot 2 - 0 \cdot 596 =$ **0·604 g**

Problem 11. **A sample of magnesium was burnt in air to give a mixture of MgO and Mg_3N_2. The ash was dissolved in 60 meq of HCl and the resulting solution back titrated with NaOH. 12 meq of NaOH were required to reach the end point. An excess of NaOH was then added and the solution distilled. The ammonia released was then trapped in 10 meq of second acid solution. Back titration of this solution required 6 meq of the base. Calculate the percentage of magnesium burnt to the nitride.** ***(Roorkee 1998)***

Solution. $MgO + 2\,HCl \longrightarrow MgCl_2 + H_2O$ ...(*i*)

$Mg_3N_2 + 8\,HCl \longrightarrow 3\,MgCl_2 + 2\,NH_4Cl$ ...(*ii*)

12 meq of NaOH ≡ 12 meq of HCl

i.e. HCl left unreacted = 12 meq

HCl used up by MgO and Mg_3N_2

= 60—12 = 48 meq = 48 millimoles

Suppose in the mixture, there are x millimoles of MgO and y millimoles of Mg_3N_2

Then $2x + 8y = 48$

or $\quad x + 4y = 24$

Further,

$$NH_4Cl + NaOH \longrightarrow NaCl + H_2O + NH_3$$

Acid used up by NH_3 = 10 — 6 = 4 meq.

$\therefore$ NH_3 produced = 4 meq = 4 millimoles

or NH_4Cl formed in reaction (*ii*) = 4 millimoles

This will be formed from Mg_3N_2 = 2 millimoles *i.e.* $y = 2$.

Hence $x + 4 \times 2 = 24$ or $x = 16$

$$2\,Mg + O_2 \longrightarrow 2\,MgO$$

$$3\,Mg + N_2 \longrightarrow Mg_3N_2$$

16 millimoles of MgO are obtained from Mg

= 16 millimoles

2 millimoles of Mg_3N_2 are obtained from Mg

= 6 millimoles

Total millimoles of Mg = 16 + 6 = 22

Hence Mg converted to $Mg_3N_2 = \frac{6}{22} \times 100$

= 27·27%.

Problem 12. **How many millilitres of 0·5 M H_2SO_4 are needed to dissolve 0·5 g of copper (II) carbonate ?**

(At. mass : H = 1, C = 12, O = 16, S = 32, Cu = 63·5) ***(I.I.T. 1999)***

Solution. $CuCO_3 + H_2SO_4$

$$\longrightarrow CuSO_4 + CO_2 + H_2O$$

1 mole $CuCO_3$ = 63·5 + 12 + 48 = 123·5 g require H_2SO_4 = 1 mole

$\therefore$ 0·5 g $CuCO_3$ will require H_2SO_4

$$= \frac{1}{123 \cdot 5} \times 0 \cdot 5 \text{ mole} = \frac{1}{247} \text{ mole}$$

0·5 mole of 0·5 M H_2SO_4 are present in 1000 ml

$\therefore \frac{1}{247}$ mole of 0·5 M H_2SO_4 will be present in

$$\frac{1000}{0 \cdot 5} \times \frac{1}{247} \text{ ml} = \mathbf{8 \cdot 1\ ml.}$$

Problem 13. **A plant virus is found to consist of uniform cylindrical particles of 150 Å in diameter and 5000 Å long. The specific volume of the virus is 0·75**

cm^3/g. If the virus is considered to be a simple particle, find its molecular weight. (*I.I.T. 1999*)

Solution. Volume of one virus particle $= \pi r^2 h$

$$= \frac{22}{7} \times \left(\frac{150}{2} \times 10^{-8}\ cm\right)^2 \times (5000 \times 10^{-8}\ cm)$$

$$= 8 \cdot 839 \times 10^{-17}\ cm^3$$

Mass of one virus particle

$$= 8 \cdot 839 \times 10^{-17}\ cm^3 \times \frac{1\ g}{0 \cdot 75\ cm^3}$$

$$= 11 \cdot 785 \times 10^{-17}\ g$$

$\therefore$ Molar mass (*i.e.* mass of Avogadro's no. of particles)

$$= (11 \cdot 785 \times 10^{-17}\ g) \times (6 \cdot 023 \times 10^{23})$$

$$= \mathbf{7 \cdot 098 \times 10^7\ g\ mol^{-1}}.$$

Problem 14. **The formula weight of an acid is 82·0 in a titration. 100 cm^3 of a solution of this acid containing 39·0 *g* of the acid per litre were completely neutralised by 95·0 cm^3 of aqueous NaOH containing 40·0 *g* of NaOH per litre. What is the basicity of the acid ?** (*Roorkee 2000*)

Solution. Suppose the equivalent weight of the acid = E

$\therefore$ Normality of the acid solution $= \frac{39}{E}$

Eq. wt. of NaOH = 40

$\therefore$ Normality of NaOH solution $= \frac{40}{40}\ N = 1\ N$

Applying $\underset{(Acid)}{N_1V_1} = \underset{(NaOH)}{N_2V_2}$

$$\frac{39}{E} \times 100 = 1 \times 95$$

or $\quad E = 41 \cdot 0$

$\therefore$ Basicity of the acid $= \frac{\text{Formula wt.}}{\text{Eq. wt.}} = \frac{82 \cdot 0}{41 \cdot 0} = 2.$

Problem 15. **Calculate the molarity of water if its density is 1000 kg/m^3** (*I.I.T. 2003*)

Solution. Molarity of water means number of moles of water in 1 litre of water

1 L of water = 1000 cm^3 = 1000 *g*

$(\because d = 1\ kg/m^3 = 1\ g/cm^3)$

$$1000\ g\ H_2O = \frac{1000}{18}\ \text{moles} = \mathbf{55 \cdot 56\ moles.}$$

Problem 16. **A mixture of FeO and Fe_3O_4 was heated in air to a constant mass. It was found to gain 10% in its mass. Calculate the percentage composition of the original mixture.** (*West Bengal J.E.E. 2004*)

Solution. On heating in air, both FeO and Fe_3O_4 are oxidized to Fe_2O_3 as follows :

$$\underset{\substack{2(56+16)\\=144\ g}}{2\ FeO} + \frac{1}{2}O_2 \longrightarrow \underset{\substack{2\times 56+3\times 16\\=160\ g}}{Fe_2O_3},$$

$$\underset{\substack{2(3\times 56+4\times 16)\\=464\ g}}{2\ Fe_3O_4} + \frac{1}{2}O_2 \longrightarrow \underset{\substack{3\times 160\\=480\ g}}{3\ Fe_2O_3}$$

Suppose FeO in the mixture = *x*%

Then Fe_3O_4 in the mixture = (100 − *x*) %

Fe_2O_3 produced from *x g* of FeO $= \frac{160}{144} \times x\ g$

Fe_2O_3 produced from (100 − *x*) *g* of Fe_3O_4

$$= \frac{480}{464} \times (100 - x)\ g$$

Total Fe_2O_3 produced from 100 *g* of the mixture = 100 + 10 = 110 *g*

$$\therefore \quad \frac{160x}{144} + \frac{480(100-x)}{464} = 110$$

or $\quad \frac{10x}{9} + \frac{30(100-x)}{29} = 110$

or $\quad 290x + 270(100-x) = 110 \times 9 \times 29$

or $\quad 20x = 1710$

or $\quad x = 85 \cdot 5$

$\therefore$ FeO present in the mixture = **85·5%**

and $\quad Fe_3O_4 = 100 - 85 \cdot 5 = \mathbf{14 \cdot 5\%}.$

MULTIPLE CHOICE QUESTIONS

For CBSE– PMT (Preliminary), IIT Screening, AIEEE, AIIMS, AFMC, DPMT, CPMT, BHU and All Other Competitive Examinations

1. The number of significant figures in 0·050 is
(*a*) 1 (*b*) 2
(*c*) 3 (*d*) 4.

2. Which out of the following is not a homogeneous mixture ?
(*a*) Air (*b*) Brass
(*c*) Solution of sugar in water
(*d*) Smoke.

3. A compound has the empirical formula $C_2H_3O_2$. Its vapour density is 59. Its molecular formula will be
(*a*) $C_2H_3O_2$ (*b*) $C_4H_6O_4$
(*c*) $C_6H_9O_6$ (*d*) $C_8H_{12}O_8$.

4. The number of molecules present in 8 g of oxygen gas are
(*a*) $6{\cdot}022 \times 10^{23}$ (*b*) $3{\cdot}011 \times 10^{23}$
(*c*) $12{\cdot}044 \times 10^{23}$ (*d*) $1{\cdot}55 \times 10^{23}$.

5. 112 cm^3 of hydrogen gas at STP contain
(*a*) 0·005 mole (*b*) 0·01 mole
(*c*) 0·02 g (*d*) $3{\cdot}011 \times 10^{22}$ molecules.

6. 10 g $CaCO_3$ on reaction with 0·1 M HCl acid will produce CO_2
(*a*) 1120 cm^3 (*b*) 2240 cm^3
(*c*) 112 cm^3 (*d*) 224 cm^3.

7. One fermi is
(*a*) 10^{-13} cm (*b*) 10^{-15} cm
(*c*) 10^{-10} cm (*d*) 10^{-12} cm.

8. Which of the following has the largest number of atoms ?
(*a*) 0·5 g atom of Cu (*b*) 0·635 g of Cu
(*c*) 0·25 moles of Cu atom
(*d*) 1 g of Cu. *(I.I.T. 1974)*

9. 27 g of Al (at mass = 27) will react with oxygen equal to
(*a*) 24 g (*b*) 8 g
(*c*) 40 g (*d*) 10 g. *(I.I.T. 1978)*

10. 2·76 g of silver carbonate (at mass of Ag = 108) on being heated strongly yields a residue weighing
(*a*) 2·16 g (*b*) 2·48 g
(*c*) 2·32 g (*d*) 2·64 g *(I.I.T. 1979)*

11. If 0·5 mol of $BaCl_2$ is mixed with 0·2 mol of Na_3PO_4, the maximum number of moles of $Ba_3(PO_4)_2$ that can be formed is
(*a*) 0·7 (*b*) 0·5
(*c*) 0·3 (*d*) 0·1. *(I.I.T. 1981)*

12. If 10^{21} molecules are removed from 200 mg of CO_2, then the number of moles of CO_2 left are
(*a*) $2{\cdot}88 \times 10^{-3}$ (*b*) $1{\cdot}66 \times 10^{-3}$
(*c*) $4{\cdot}54 \times 10^{-3}$ (*d*) $1{\cdot}66 \times 10^{-2}$. *(I.I.T. 1983)*

13. The number of gram molecules of oxygen in $6{\cdot}02 \times 10^{24}$ CO molecules is
(*a*) 10 g molecules (*b*) 5 g molecules
(*c*) 1 g molecule (*d*) 0·5 g molecule. *(I.I.T. 1990)*

14. One mole of CO_2 contains
(*a*) $6{\cdot}02 \times 10^{23}$ atoms of C
(*b*) $6{\cdot}02 \times 10^{23}$ atoms of O
(*c*) $18{\cdot}1 \times 10^{23}$ molecules of CO_2
(*d*) 3 g atoms of CO_2. *(M.L.N.R. 1990)*

15. Two elements X (atomic weight = 75) and Y (atomic weight = 16) combine to give a compound having 75·8 % of X. The formula of the compound is
(*a*) XY (*b*) X_2Y
(*c*) XY_2 (*d*) X_2Y_3. *(M.L.N.R. 1991)*

16. The largest number of molecules is in
(*a*) 54 g of nitrogen peroxide
(*b*) 28 g of carbon dioxide
(*c*) 36 g of water (*d*) 46 g of ethyl alcohol. *(M.L.N.R. 1991)*

ANSWERS

1. *b*	2. *d*	3. *b*	4. *d*	5. *a*	6. *b*	7. *a*	8. *a*	9. *a*	10. *c*
11. *d*	12. *a*	13. *b*	14. *a*	15. *a*	16. *c*				

17. The molecular weight of O_2 and SO_2 are 32 and 64 respectively. If one litre of O_2 at 15°C and 750 mm contains N molecules, the number of molecules in two litres of SO_2 under the same conditions of temperature and pressure will be

(*a*) N/2 (*b*) N
(*c*) 2N (*d*) 4N. *(M.L.N.R. 1991)*

18. Five grams of each of the following gases at 87°C and 750 mm pressure are taken. Which of them will have the least volume ?

(*a*) HF (*b*) HCl
(*c*) HBr (*d*) HI.
(M.L.N.R. 1991)

19. Which one of the following is the standard for atomic mass ?

(*a*) ${}^{1}_{1}H$ (*b*) ${}^{12}_{6}C$
(*c*) ${}^{14}_{6}C$ (*d*) ${}^{16}_{8}O$. *(I.I.T. 1993)*

20. Rearrange the following (I to IV) in the order of increasing masses and choose the correct answer from (*a*) (*b*) (*c*) and (*d*). Atomic masses : N = 14, O = 16, Cu = 63)

I. 1 molecule of oxygen 32
II. 1 atom of nitrogen 14
III. 1×10^{10} g molecular weight of oxygen
IV. 1×10^{-18} g atomic weight of copper

(*a*) II < I < III < IV (*b*) IV < III < II < I
(*c*) II < III < I < IV (*d*) III < IV < I < II.
(I.I.T. 1993)

21. The number of moles of H_2 in 0·224 litres of hydrogen gas at STP (273 K, 1 atm) (assuming ideal gas behaviour) is

(*a*) 1 (*b*) 0·1
(*c*) 0·01 (*d*) 0·001
(M.L.N.R. 1994)

22. Two containers P and Q of equal volume (1 litre each) contain 6g of O_2 and SO_2 respectively at 300 K and 1 atmosphere. Then

(*a*) No. of molecules in P is less than that in Q
(*b*) No. of the molecules in Q is less than that in P
(*c*) No. of molecules in P and Q are same
(*d*) Either (*a*) or (*b*) *(Haryana C.E.E.T. 1994)*

23. The maximum amount of $BaSO_4$ precipitated on mixing $BaCl_2$ (0·5 M) with H_2SO_4 (1M) will correspond to

(*a*) 0·5 M (*b*) 1·0 M
(*c*) 1·5 M (*d*) 2·0 M
(A.I.I.M.S. 1997)

24. Given the numbers : 161 cm, 0·161 cm, 0·0161 cm. The number of significant figures for the three numbers are

(*a*) 3, 4 and 5 respectively
(*b*) 3, 3 and 3 respectively
(*c*) 3, 3 and 4 respectively
(*d*) 3, 4 and 4 respectively. *(C.B.S.E. P.M.T. 1998)*

25. Haemoglobin contains 0·33% of iron by weight. The molecular weight of haemoglobin is approximately 67200. The number of iron atoms (at. wt. of Fe = 56) present in one molecule of haemoglobin is

(*a*) 6 (*b*) 1
(*c*) 4 (*d*) 2.
(C.B.S.E. P.M.T. 1998)

26. In the reaction $4\,NH_3\,(g) + 5\,O_2\,(g) \longrightarrow 4\,NO\,(g) + 6\,H_2O\,(l)$, when 1 mole of ammonia and 1 mole of O_2 are made to react to completion

(*a*) 1·0 mole of H_2O is produced
(*b*) 1·0 mole of NO will be produced
(*c*) all the oxygen will be consumed
(*d*) all the ammonia will be consumed.
(C.B.S.E. P.M.T. 1998)

27. The number of molecules in 16 g of methane is

(*a*) $3{\cdot}0 \times 10^{23}$ (*b*) $6{\cdot}02 \times 10^{23}$
(*c*) $\frac{16}{6{\cdot}02} \times 10^{23}$ (*d*) $\frac{16}{3{\cdot}0} \times 10^{23}$
(M.P. P.M.T. 1998)

28. 50 ml 10 N H_2SO_4, 25 ml 12 N HCl and 40 ml 5 N HNO_3 were mixed together and the volume of the mixture was made 1000 ml by adding water. The normality of the resultant solution will be :

(*a*) 1 N (*b*) 2 N
(*c*) 3 N (*d*) 4N.
(M.P. P.M.T. 1998)

29. The number of atoms in 4·25 g of NH_3 is approximately

(*a*) 1×10^{23} (*b*) 2×10^{23}
(*c*) 4×10^{23} (*d*) 6×10^{23}.
(C.B.S.E. P.M.T. 1999)

ANSWERS

17. *c*	18. *d*	19. *b*	20. *a*	21. *c*	22. *b*	23. *a*	24. *b*	25. *c*	26. *c*
27. *b*	28. *a*	29. *d*							

30. A molal solution is one that contains 1 mole of a solute in
(a) 1000 g of the solvent
(b) one litre of the solvent
(c) one litre of the solution
(d) 22·4 litres of the solution. (M.P. C.E.E. 1999)

31. A 100 ml solution of 0·1 N HCl was titrated with 0·2 N NaOH solution. The titration was discontinued after adding 30 ml of NaOH solution. The remaining titration was completed by adding 0·25 N KOH solution. The volume of KOH required for completing the titration is
(a) 70 ml (b) 32 ml
(c) 35 ml (d) 16 ml. (D.C.E. 1999)

32. One mole of calcium phosphide on reaction with excess of water gives
(a) one mole of phosphine
(b) two moles of phosphoric acid
(c) two moles of phosphine
(d) one mole of phosphorus pentoxide. (I.I.T. 1999)

33. Assuming fully decomposed, the volume of CO_2 released at STP on heating 9·85 *g* of $BaCO_3$ (Atomic mass, Ba = 137) will be
(a) 0·84 L (b) 2·24 L
(c) 4·06 L (d) 1·12 L. (C.B.S.E. P.M.T. 2000)

34. The specific heat of a metal is 0·16. Its *approximate* atomic weight would be
(a) 32 (b) 16
(c) 40 (d) 64. (A.I.I.M.S. 2000)

35. The weight of a molecule of the compound $C_{60}H_{122}$ is
(a) 1.4×10^{-21} g (b) 1.09×10^{-21} g
(c) 5.025×10^{23} g (d) 16.023×10^{23} g. (A.I.I.M.S. 2000)

36. The number of water molecules present in a drop of water (volume 0·0018 ml) at room temperature is
(a) 6.023×10^{19} (b) 1.084×10^{18}
(c) 4.84×10^{17} (d) 6.023×10^{23}. (D.C.E. 2000)

37. Which of the following contains maximum number of atoms ?
(a) 6.023×10^{21} molecules of CO_2
(b) 22·4 L of CO_2 at STP
(b) 0·44 *g* of CO_2
(d) None of these. (J.I.P.M.E.R. 2000)

38. Number of *g* of oxygen in 32·2 *g* $Na_2SO_4 . 10 H_2O$ is
(a) 20·8 (b) 22·4
(c) 2·24 (d) 2·08. (Haryana C.E.E.T. 2000)

39. The percentage of Se in peroxidase anhydrous enzyme is 0·5% by weight (atomic weight = 78·4). Then minimum molecular weight of peroxidase anhydrous enzyme is
(a) 1.568×10^4 (b) 1.568×10^3
(c) 15·68 (d) 3.136×10^4 (C.B.S.E. P.M.T. 2001)

40. The percentage of nitrogen in urea is about
(a) 46 (b) 85
(c) 18 (d) 28 (K.C.E.T. 2001)

41. How much of NaOH is required to neutralise 1500 cm^3 of 0·1 N HCl ? (Na = 23)
(a) 40 g (b) 4 g
(c) 6 g (d) 60 g (K.C.E.T. 2001)

42. 10 dm^3 of N_2 gas and 10 dm^3 of gas X at the same temperature contain the same number of molecules. The gas X is
(a) CO (b) CO_2
(c) H_2 (d) NO (K.C.E.T. 2001)

43. The volume of water to be added to 100 cm^3 of 0·5 N H_2SO_4 to get deci normal concentration is
(a) 100 cm^3 (b) 450 cm^3
(c) 500 cm^3 (d) 400 cm^3 (K.C.E.T. 2001)

44. The set of numerical coefficients that balances the equation
$K_2CrO_4 + HCl \longrightarrow K_2Cr_2O_7 + KCl + H_2O$ is
(a) 1, 1, 2, 2, 1 (b) 2, 2, 1, 1, 1
(c) 2, 1, 1, 2, 1 (d) 2, 2, 1, 2, 1
(e) 2, 2, 2, 1, 1 (Kerala E.E.E. 2001)

45. 250 ml of a sodium carbonate solution contains 2·65 grams of Na_2CO_3. If 10 ml of this solution is

ANSWERS

30. a	31. d	32. c	33. d	34. c	35. a	36. a	37. b	38. b	39. a
40. a	41. c	42. a	43. d	44. d					

diluted to one litre, what is the concentration of the resultant solution ? (mol. wt. of Na_2CO_3 = 106)

(a) 0·1 M (b) 0·001 M

(c) 0·01 M (d) 10^{-4} M

(E.A.M.C.E.T. 2001)

46. 7·5 grams of a gas occupy 5·6 lires of volume at STP. The gas is

(a) NO (b) N_2O

(c) CO (d) CO_2

(E.A.M.C.E.T. 2001)

47. An aqueous solution of 6·3 g of oxalic acid dihydrate is made up to 250 ml. The volume of 0·1 N NaOH required to completely neutralise 10 ml of this solution is

(a) 40 ml (b) 20 ml

(c) 10 ml (d) 4 ml. (I.I.T. 2001)

48. How many moles of electron weigh one kilogram ?

(a) $6{\cdot}023 \times 10^{23}$ (b) $\frac{1}{9{\cdot}108} \times 10^{31}$

(c) $\frac{6{\cdot}023}{9{\cdot}108} \times 10^{54}$ (d) $\frac{1}{9{\cdot}108 \times 6{\cdot}023} \times 10^{8}$

(I.I.T. 2002)

49. The prefix 10^{18} is

(a) giga (b) exa

(c) kilo (d) nano

(e) mega (Kerala M.E.E. 2002)

50. Number of atoms in 558·6 g Fe (molar mass Fe = 55·86 g mol^{-1}) is

(a) twice that in 60 g carbon

(b) $6{\cdot}023 \times 10^{22}$ (c) half that of 8 g He

(d) $558{\cdot}6 \times 6{\cdot}023 \times 10^{23}$ (A.I.E.E.E. 2002)

51. 3·92 g of ferrous ammonium sulphate are dissolved in 100 ml of water. 20 ml of this solution requires 18 ml of potassium permanganate during titration for complete oxidation. The weight of $KMnO_4$ present in one litre of the solution is

(a) 34·76 g (b) 12·38 g

(c) 1·238 g (d) 3·476 g

(Tamil Nadu C.E.T. 2002)

52. 3 g of an oxide of a metal is converted to chloride completely and it yielded 5 g of chloride. The equivalent weight of the metal is

(a) 33·25 (b) 3·325

(c) 12 (d) 20 (K.C.E.T. 2002)

53. Among $FeSO_4 . 7\,H_2O$ (A), $CuSO_4 . 5\,H_2O$ (B), $ZnSO_4 . 7\,H_2O$ (C), $MnSO_4 . 4\,H_2O$ (D), isomorphous salts are

(a) A and C (b) A and D

(c) C and B (d) A and B

(D.P.M.T. 2003)

54. What will be the volume of the mixture after the reaction ?

$(g)\ NH_3\ (1\ L) + (g)\ HCl\ (1{\cdot}5\ L) \longrightarrow NH_4Cl\ (s)$

(a) 1·5 L (b) 0·5 L

(c) 1 L (d) 0 L (D.P.M.T. 2003)

55. A compound has haemoglobin like structure. It has one Fe. It contain 4·6% of Fe. The approximate molecular mass is

(a) 100 g mol^{-1} (b) 1200 g mol^{-1}

(c) 1400 g mol^{-1} (d) 1600 g mol^{-1}

(D.P.M.T. 2003)

56. In Haber process, 30 litres of dihydrogen and 30 litres of dinitrogen were taken for reaction which yeilded only 50% of the expected product. What will be the compositon of the gaseous mixture under the aforesaid conditon in the end ?

(a) 20 litres NH_3, 25 litres N_2, 20 litres H_2

(b) 10 litres NH_3, 25 litres N_2, 15 litres H_2

(c) 20 litres NH_3, 10 litres N_2, 30 litres H_2

(d) 20 litres NH_3, 25 litres N_2, 15 litres H_2

(C.B.S.E. P.M.T. 2003)

57. What volume of hydrogen gas at 273 K and 1 atm pressure will be consumed in obtaining 21·6 g elemental boron (atomic mass = 10·8) from the reduction of boron trichloride by hydrogen ?

(a) 67·2 L (b) 44·8 L

(c) 22·4 L (d) 89·6 L

(A.I.E.E.E. 2003)

58. Which has maximum number of atoms ?

(a) 24 g of C (12) (b) 56 g of Fe (56)

(c) 27 g of Al (27) (d) 108 g of Ag (108)

(I.I.T. 2003)

59. Mixture X = 0·02 mol of $[Co(NH_3)_5SO_4]Br$ and 0·02 mol of $[Co(NH_3)_5Br]SO_4$ was prepared in 2 litre of solution

1 litre of mixture X + excess $AgNO_3 \longrightarrow$ Y

1 litre of mixture X + excess $BaCl_2 \longrightarrow$ Z

Number of moles of Y and Z are

(a) 0·01, 0·01 (b) 0·02, 0·01

ANSWERS

45. b	46. a	47. a	48. d	49. b	50. a	51. d	52. a	53. a	54. b
55. b	56. b	57. a	58. a						

(c) 0·01, 0·02 (d) 0·02, 0·02

(I.I.T. 2003)

60. The maximum number of molecules is present in

(a) 15 L of H_2 gas S.T.P. (b) 5 L of N_2 gas at S.T.P.

(c) 0·5 g of H_2 gas (d) 10 g of O_2 gas

(C.B.S.E. P.M.T. 2004)

61. 4 g of copper was dissolved in concentrated nitric acid. The copper nitrate on strong heating gave 5 g of its oxide. The equivalent weight of copper is

(a) 23 (b) 32 (c) 12 (d) 20

(Karnataka C.E.T. 2004)

62. Dulong and Petit's law is valid only for

(a) metals (b) non-metals

(c) gaseous elements (d) solid elements

(Karnataka C.E.T. 2004)

63. Number of atoms of oxygen present in 10·6 g Na_2CO_3 will be

(a) $6{\cdot}02 \times 10^{22}$ (b) $12{\cdot}04 \times 10^{22}$

(c) $1{\cdot}806 \times 10^{23}$ (d) $31{\cdot}80 \times 10^{28}$

(J & K.C.E.T. 2004)

64. One gram mole of a gas at NTP occupies 22·4 litres. This fact was derived from

(a) Law of gaseous volumes

(b) Avogadro's hypothesis

(c) Berzelius hypothesis

(d) Dalton's atomic theory (J & K.C.E.T. 2004)

65. Which of the following contains maximum number of molecules ?

(a) 100 cc of CO_2 at STP

(b) 150 cc of N_2 at STP

(c) 50 cc of SO_2 at STP

(d) 150 cc of O_2 at STP

(e) 200 cc of NH_3 at STP (Kerala C.E.E. 2004)

66. A sample of phosphorus trichloride (PCl_3) contains 1·4 moles of the substance. How many atoms are there is the sample ?

(a) 4 (b) 5·6

(c) $8{\cdot}431 \times 10^{23}$ (d) $3{\cdot}372 \times 10^{24}$

(e) $2{\cdot}409 \times 10^{24}$ (Kerala M.E.E. 2004)

67. A gas mixture contains 50% helium and 50% methane by volume. What is the percent by weight of methane in the mixture ?

(a) 19·97% (b) 20·05%

(c) 50% (d) 75%

(e) 80·03% (Kerala M.E.E. 2004)

68. 116 mg of a compound on vaporisation in a Victor Meyer's apparatus displaces 44·8 ml of air measured at STP. The molecular weight of the compound is

(a) 116 (b) 232

(c) 58 (d) 44·8

(e) 46·4 (Kerala M.E.E. 2004)

69. Number of water molecules in the drop of water, if 1 ml of water has 20 drops and A is Avogadro's number, is

(a) 0·5 A / 18 (b) 0·05 A

(c) 0·5 A (d) 0·05 A / 18

(U.P. C.P.M.T. 2004)

HINTS/EXPLANATIONS to Multiple Choice Questions

2. Carbon particles (solid phase) are dispersed into air (gaseous phase).

9. $\underset{4 \times 27\,g}{4\,Al} + \underset{3 \times 32\,g}{3\,O_2} \longrightarrow 2\,Al_2O_3$

10. $\underset{\substack{2 \times 108 + 12 + 3 \times 16 \\ = 276\,g}}{Ag_2CO_3} \longrightarrow \underset{\substack{2 \times 108 + 16 \\ = 232\,g}}{Ag_2O} + CO_2 \uparrow$

11. $3\,BaCl_2 + 2\,Na_3PO_4 \longrightarrow Ba_3(PO_4)_2 + 6\,NaCl$

3 moles of $BaCl_2$ react with 2 moles of Na_3PO_4. Na_3PO_4 will be the limiting reactant.

12. 200 mg of CO_2 = 0·200 g = $\frac{0{\cdot}2}{44}$ mol

$= \frac{0 \cdot 2}{44} \times 6 \cdot 023 \times 10^{23}$ molecules

$= 2 \cdot 738 \times 10^{21}$ molecules

After removing 10^{21} molecules, molecules left

$= (2 \cdot 738 - 1)\, 10^{21}$

$= 1 \cdot 738 \times 10^{21}$ molecules

$= \frac{1 \cdot 738 \times 10^{21}}{6 \cdot 023 \times 10^{23}}$ mol $= 2 \cdot 88 \times 10^{-3}$ mol.

13. 1 mol of CO ≡ 1 g atom of O

or $6{\cdot}02 \times 10^{23}$ molecules of CO ≡ $\frac{1}{2}$ g molecule of O_2

ANSWERS

59. a 60. a 61. b 62. d 63. c 64. b 65. e 66. d 67. e 68. c 69. d

$\therefore$ $6 \cdot 02 \times 10^{24}$ molecules of CO

$$\equiv \frac{1}{2} \times \frac{6 \cdot 02 \times 10^{24}}{6 \cdot 02 \times 10^{23}} = 5 \text{ g molecules}$$

15. $X = 75 \cdot 8\%$, $Y = 24 \cdot 2\%$. Find empirical formula.

16. $54 \text{ g } N_2O_5 = 54/108 \text{ mol} = 0 \cdot 50 \text{ mol}$

$28 \text{ g } CO_2 = 28/44 \text{ mol} = 0 \cdot 64 \text{ mol}$

$36 \text{ g } H_2O = 36/18 \text{ mol} = 2 \text{ mol}$

$46 \text{ g } C_2H_5OH = 46/46 \text{ mol} = 1 \text{ mol}$

17. Equal volumes of different gases under same conditions of temp. and pressure contain equal number of molecules.

18. 5 g HF = 5/20 mol, 5 g HCl = 5/36·5 mol, 5 g HBr = 5/81 mol, 5 g HI = 5/121 mol. Thus HI has least no. of moles and hence least volume.

22. $6 \text{ g } O_2 = 6/32$ mol, $6 \text{ g } SO_2 = 6/44$ mol. Thus vessel Q has less no. of moles.

25. Fe present in 67200 u $= \frac{0 \cdot 33}{100} \times 67200$

$= 222 \text{ u} = \frac{222}{56} = 4$ atoms.

26. O_2 will be limiting reactant. Hence all oxygen will be consumed.

28. $N_1V_1 + N_2V_2 + N_3V_3 = N_4V_4$

$10 \times 50 + 12 \times 25 + 5 \times 40 = N_4 \times 1000$

or $N_4 = 1$ N

29. $4 \cdot 25 \text{ g } NH_3 = 4 \cdot 25/17$ mol

$= 0 \cdot 25 \times 6 \cdot 022 \times 10^{23}$ molecules

$= (1 + 3) \times 0 \cdot 25 \times 6 \cdot 022 \times 10^{23}$ atoms

31. $N_1V_1 = N_2V_2 + N_3V_3$ *i.e.*

$$0 \cdot 1 \times 100 = 0 \cdot 2 \times 30 + 0 \cdot 25 \, V_3$$

32. $Ca_3P_2 + 6 \, H_2O \longrightarrow 3 \, Ca(OH)_2 + 2 \, PH_3$

34. Approx. atomic weight = 6·4/Specific heat (Dulong and Petit's law)

35. 1 mol of $C_{60}H_{122} = 60 \times 12 + 122 \times 1 \, g = 842$ g

$= 6 \cdot 023 \times 10^{23}$ molecules

36. $0 \cdot 0018 \text{ ml } H_2O = 0 \cdot 0018 \text{ g } H_2O$ $(\because d = 1 \, g \, ml^{-1})$

$1 \text{ mol } H_2O = 18 \text{ g} = 6 \cdot 023 \times 10^{23}$ molecules

38. 1 mol $Na_2SO_4.10 \, H_2O = 14 \, g$ atoms of (O)

i.e. 322 *g* $Na_2SO_4.10 \, H_2O$ contain (O)=224 *g*.

39. 0·5% by weight means 0·5 g Se is present in 100 g of peroxidase anhydrous enzyme. As at least one atom of Se must be present in the enzyme and atomic weight of Se = 78·4, therefore 1 g atom of Se *i.e.* 78·4 g will be present in enzyme

$$= \frac{100}{0 \cdot 5} \times 78 \cdot 4 \text{ g} = 1 \cdot 568 \times 10^4 \text{ g}$$

40. $NH_2CONH_2 \equiv 2 \text{ N}$

1 mol — 2 g atoms

60 g — 28 g

$\therefore$ %age of N $= \frac{28}{60} \times 100 = 46\%$

41. 1500 cm^3 of 0·1 N HCl $\equiv \frac{0 \cdot 1}{1000} \times 1500$

$= 0 \cdot 15$ g eq.

It will neutralise NaOH = 0·15 g eq.

$= 0 \cdot 15 \times 40 \text{ g} = 6$ g

42. No. of moles of N_2 and X should be equal. This can be so if X has same molecular weight as N_2.

43. $N_1V_1 = N_2V_2$ *i.e.* $0 \cdot 5 \times 100 = 0 \cdot 1 \times V_2$

or $V_2 = 500 \text{ cm}^3$

$\therefore$ Water to be added to 100 cm^3 solution

$= 500—100 = 400 \text{ cm}^3$.

44. $\underset{2}{2 \, K_2CrO_4} + \underset{2}{2 \, HCl} \longrightarrow \underset{1}{K_2Cr_2O_7} + \underset{2}{2 \, KCl} + \underset{1}{H_2O}$

45. Molarity of given Na_2CO_3 sol.

$$= \frac{2 \cdot 65}{106} \times \frac{1}{250} \times 1000 = 0 \cdot 1 \text{ M}$$

$M_1V_1 = M_2V_2$, $10 \times 0 \cdot 1 = 1000 \times M_2$

or $M_2 = 0 \cdot 001$ M

46. Mass of 22·4 L of gas at STP $= \frac{7 \cdot 5}{5 \cdot 6} \times 22 \cdot 4$

$= 30$ g

i.e. molecular mass of the gas = 30. Hence gas is NO.

47. Normality of oxalic acid sol.

$$= \frac{6 \cdot 3}{250} \times \frac{1000}{63} = 0 \cdot 4 \text{ N}$$

$10 \times 0 \cdot 4 = V \times 0 \cdot 1$ or $V = 40$ ml.

48. 1 mol of electrons weigh

$= (9 \cdot 108 \times 10^{-31}) \times (6 \cdot 023 \times 10^{23})$ kg

or $9 \cdot 108 \times 10^{-31} \times 6 \cdot 023 \times 10^{23}$ kg of electrons $\equiv 1$ mol of electrons

$\therefore$ 1 kg of electrons

$$= \frac{1}{9 \cdot 108 \times 6 \cdot 023} \times 10^8 \text{ mol}$$

50. $558 \cdot 5$ g Fe = 10 moles

$= 10 \times 6 \cdot 023 \times 10^{23}$ atoms

60 g C = 5 moles $= 5 \times 6 \cdot 023 \times 10^{23}$ atoms.

51. Eq. mass of ferrous ammonium sulphate = 392

$\because$ Normality of the salt solution

$$= \frac{3 \cdot 92}{392} \times \frac{1}{100} \times 1000 = 0 \cdot 1 \text{ N}$$

20 ml of 0·1 N salt sol ≡ 18 ml of $KMnO_4$ sol.

$\therefore$ Normality of $KMnO_4$ sol. $= \frac{20 \times 0 \cdot 1}{18} = \frac{1}{9}$ N

Eq. mass of $KMnO_4$ = 31·6

$\therefore$ Strength of $KMnO_4$ sol. $= \frac{1}{9} \times 31 \cdot 6$

$= 3 \cdot 5 \text{ g L}^{-1}$ approx.

52. $\frac{\text{mass of metal oxide}}{\text{mass of metal chloride}}$

$$= \frac{\text{Eq. mass of metal + Eq. mass of oxide}}{\text{Eq. mass of metal + Eq. mass of chloride}}$$

$= \frac{3}{5} = \frac{E + 8}{E + 35 \cdot 5}$. This gives E = 33·25.

54. 1 L NH_3 (*g*) reacts with 1 L HCl (*g*) to form NH_4Cl (*s*) which has negligible volume. Hence final mixture will contain only 0·5 L HCl.

55. 1 *g* atom of Fe (56 *g* Fe) is present in 1 mole of the compound. As 4·6 *g* Fe are present in 100 *g* of the compound, 56 *g* of Fe will be present in $\frac{100}{4 \cdot 6} \times 56\,g$ = 1217 *g* of the compound. Hence approximate molecular mass = 1200.

56. $N_2 (g) + 3 H_2 (g) \rightleftharpoons 2 NH_3 (g)$

1 L of N_2 reacts with 3 L of H_2 to form 2 L of NH_3. Thus N_2 is the limitinng reactant.

10 L N_2 will react with 30 L H_2 to form 20 L NH_3. As actual yield is 50% of the expected value, NH_3 formed = 10 L, N_2 reacted = 5 L, H_2 reacted = 15 L

$\therefore$ Mixture will contain 10 L NH_3, 25 L N_2, 15 L H_2.

57. $\underset{\substack{3 \times 22 \cdot 4 \text{ L} \\ = 67 \cdot 2 \text{ L}}}{2 BCl_3 + 3 H_2} \longrightarrow \underset{\substack{2 \times 10 \cdot 8 \\ = 21 \cdot 6\,g}}{2 B} + 6 HCl$

58. 24 g C $= \frac{24}{12}$ mol = 2 mol = 2 × N_0 atoms

56 *g* Fe $= \frac{56}{56}$ mol = 1 mol = N_0 atoms

27 g Al $= \frac{27}{27}$ mol = 1 mol = N_0 atoms

108 g Ag $= \frac{108}{108}$ mol = 1 mol = N_0 atoms

59. Mixture X will contain 0·02 mol Br^- ions and 0·02 mol SO_4^{2-} ions in 2 L solution. Hence 1 L of mixture X will contain 0·01 mol Br^- and 0·01 mol SO_4^{2-} ions. With excess of $AgNO_3$, 0·01 mol of AgBr *i.e.* Y is formed. With excess of $BaCl_2$, 0·01 mol of $BaSO_4$ *i.e.* Z is formed.

60. At STP, 22·4 L of any gas = 6·02 × 10^{23} molecules

$\therefore$ 15 L $H_2 = \frac{6 \cdot 02 \times 10^{23}}{22 \cdot 4} \times 15 = 4 \cdot 03 \times 10^{23}$,

5 L $N_2 = \frac{6 \cdot 02 \times 10^{23}}{22 \cdot 4} \times 5 = 1 \cdot 344 \times 10^{23}$

2 *g* H_2 = 6·02 × 10^{23} molecules

$\therefore$ 0·5 *g* $H_2 = \frac{6 \cdot 02 \times 10^{23}}{2} \times 0 \cdot 5$ molecules

= 1·505 × 10^{23} molecules

32 *g* O_2 = 6·02 × 10^{23} molecules

$\therefore$ 10 *g* $O_2 = \frac{6 \cdot 02 \times 10^{23}}{32} \times 10$ molecules

= 1·88 × 10^{23} molecules

61. Cu = 4 *g*, CuO = 5 *g* $\therefore$ Oxygen = 1 *g*. Thus 1 *g* oxygen combines with Cu = 4 *g*.

$\therefore$ 8 *g* oxygen will combine with Cu = 4 × 8 = 32 *g*

$\therefore$ Eq. wt. of Cu = 32.

63. Molar mass of Na_2CO_3 = 2 × 23 + 12 + 3 × 16 = 106 *g* mol^{-1}

$\therefore$ 10·6 *g* Na_2CO_3 = 0·1 mol

= 0·1 × 6·02 × 10^{23} molecules

= 3 × 0·1 × 6·02 × 10^{23} atoms of O

= 1·806 × 10^{23}.

65. 22400 cc of any gas at STP contains same number of molecules *i.e.* 6·02 × 10^{23}. Hence larger the volume at STP, greater is the number of molecules.

66. 1 mole of PCl_3 contains 4 × 6·02 × 10^{23} atoms $\therefore$ 1·4 moles will contain = 3·372 × 10^{24} atoms.

67. Equal volumes contain equal number of moles. Hence molar ratio of He : CH_4 = 1 : 1 $\therefore$ Ratio by weight = 4 : 16 = 1 : 4

$\therefore$ CH_4 present by weight $= \frac{4}{5} \times 100 = 80\%$.

68. 22400 ml of air at STP will be displaced by $= \frac{0 \cdot 116}{44 \cdot 8} \times 22400\,g = 58\,g$.

69. 1 drop of water $= \frac{1}{20}$ ml $= \frac{1}{20}\,g$

$(d_{H_2O} = 1\,g\,ml^{-1})$

18 *g* of water = A molecules

$\therefore$ $\frac{1}{20}\,g$ water $= \frac{A}{18} \times \frac{1}{20} = \frac{0 \cdot 05\,A}{18}$ molecules

ADDITIONAL QUESTIONS

For All Competitive Examinations

Assertion-Reason Type Questions

The following questions consist of two statements each printed as Assertion and Reason. While answering these questions, you are required to choose any one of the following five responses.

(a) If both Assertion and Reason are true and the Reason is a correct explanation of the Assertion.

(b) If both Assertion and Reason and true but Reason is not a correct explanation of the Assertion.

(c) If Assertion is true but the Reason is false.

(d) If the Assertion is false but the Reason is true.

(e) If both Assertion and Reason are false.

	Assertion	Reason
1.	Cinnabar is a chemical compound whereas brass is a mixture.	Cinnabar always contains 6·25 times as much mercury as sulphur by weight. Brass can be made with widely different ratios of copper and zinc.
2.	A single C^{12} atom has a mass exactly 12 amu and a mole of these atoms has a mass of exactly 12 grams.	A mole of atoms of any element has a mass in grams equal to its atomic weight.
3.	Pure water obtained from different sources such as river, well, spring, sea etc. always contains hydrogen and oxygen in the ratio of 1 : 8 by mass.	Mass of reactants and products during chemical or physical change is always the same.
4.	In a gaseous reaction, the ratio of volumes of reactants and gaseous products is in agreement with their molar ratio.	Volume of gas is inversely proportional to its moles at particular temperature and pressure.
5.	The standard unit for expressing the mass of atoms is amu.	amu is also called as avogram.
6.	Both 106 *g* of sodium carbonate and 12 *g* of carbon have same number of carbon atoms.	Both contain 1 *g* atom of carbon which contains $6{\cdot}023 \times 10^{23}$ carbon atoms.
7.	Average atomic mass of boron is 10·3.	Boron has two isotopes B^{10} and B^{11} whose percentage abundances are 19·6% and 80·4% respectively.
8.	Atomic mass of sodium is 23.	An atom of sodium is 23 times heavier than $\frac{1}{12}$th of the mass of carbon atom (C^{12}).

True/False Statements

Which of the following statements are not true ?

1. The zeros on the right of a decimal point are not significant.
2. Petrol is a homogeneous mixture of a number of hydrocarbons.
3. Mass of reactants is always equal to the mass of the products.
4. The volumes of oxygen which combine with a fixed volume of nitrogen in N_2O , NO and NO_2 bear a simple ratio to one another.
5. Equal volumes of different gases under similar conditions of temperature and pressure contain equal number of atoms.
6. Atom is not indestructible.
7. Empirical formula represents the actual number of atoms present in a molecule of the substance.
8. A balanced equation contains equal number of atoms of each element on both sides of the equation.

Fill In The Blanks

1. The total number of digits in a number including the last whose value is uncertain is called........
2. The S.I. unit of pressure is.........
3. 1 picometre = ...metre.
4. A pure substance that contains only one kind of atoms is called......
5. The symbol 'u' used for expressing atomic and molecular masses represents...........scale based on ....isotope.
6. The law which states that a chemical compound always contains the same elements combined in a fixed ratio by mass is called.........
7. The atomic mass of an element is the average relative mass of its atoms as compared with an atom of.........taken as.........
8. According to S.I. a mole is that amount of the substance which contains as many elementary entities as there are atoms in exactly.........kg of...... isotope.
9. The reactant which reacts completely and decides the amount of the product is called...........
10. The weight of 1×10^{22} molecules of $CuSO_4 \cdot 5H_2O$ is........... **(*I.I.T. 1991*)**

Matching Type Questions

Match enteries of column A with appropriate enteries of column B.

A		B	
1.	Petrol	1.	Compound
2.	Brass	2.	Element
3.	Sugar	3.	Mixture
4.	Gold	4.	Mixture

ANSWERS

ASSERTION-REASON TYPE QUESTIONS

1. *a* **2.** *a* **3.** *b* **4.** *c* **5.** *b* **6.** *a* **7.** *d* **8.** *a*

TRUE/FALSE STATEMENTS

1, 4, 5, 7.

FILL IN THE BLANKS QUESTIONS

1. significant figures **2.** N m^{-2} or kg m^{-1} s^{-1} **3.** 10^{-12} **4.** element **5.** unified, C-12 **6.** law of constant composition or law of definite proportions **7.** carbon, 12 **8.** 0·012, C—12 **9.** limiting reactant

10. $\frac{249 \cdot 5}{6 \cdot 02 \times 10^{23}} \times 10^{22} = 4 \cdot 14\ g.$

MATCHING TYPE QUESTIONS

1—3, 2—4, 3—1, 4—2.

HINTS/EXPLANATIONS to Assertion-Reason Type Questions

3. Both Assertion and Reason are correct statements but Reason is not the correct explanation. The correct explanation is based on law of constant composition.
4. Volume of a gas is directly proportional to its moles at a particular temperature and pressure.
5. Both Assertion and Reason are correct statements but Reason is not the correct explanation.
6. $106\ g\ Na_2CO_3 = 1$ mol

$\equiv 1$ g atom of C

$12\ g\ C = 1$ g atom of C.

7. Reason is correct but average atomic mass of B

$$= \frac{10 \times 19 \cdot 6 + 11 \times 80 \cdot 4}{100}$$

$$= 10 \cdot 8$$