

## SOME BASIC CONCEPTS OF CHEMISTRY

Chemistry is basically an experimental science. In it we study physical and chemical properties of substance and measure it upto possibility. The results of measurement can be reported in two steps : (a) Arithmetic number, (b) Unit of measurement.

Every experimental measurement varies slightly from one another and involves some error or uncertainty depending upon the skill of person making the measurements and measuring instrument. The closeness of the set of values obtained from identical measurements called **precision** and a related term, refers to the closeness of a single measurement to its true value called **accuracy**.

### Significant figures.

In the measured value of a physical quantity, the digits about the correctness of which we are surplus the last digit which is doubtful, are called the significant figures. Number of significant figures in a physical quantity depends upon the least count of the instrument used for its measurement.

(1) **Common rules for counting significant figures** : Following are some of the common rules for counting significant figures in a given expression:

**Rule 1.** All non zero digits are significant.

Example :  $x = 1234$  has four significant figures. Again  $x = 189$  has only three significant figures.

**Rule 2.** All zeros occurring between two non zero digits are significant.

Example :  $x = 1007$  has four significant figures. Again  $x = 1.0809$  has five significant figures.

**Rule 3.** In a number less than one, all zeros to the right of decimal point and to the left of a non zero digit are not significant.

Example :  $x = 0.0084$  has only two significant digits. Again,  $x = 1.0084$  has five significant figures. This is on account of rule 2.

**Rule 4.** All zeros on the right of the last non zero digit in the decimal part are significant.

Example :  $x = 0.00800$  has three significant figures 8, 0, 0. The zeros before 8 are not significant again 1.00 has three significant figures.

**Rule 5.** All zeros on the right of the non zero digit are not significant.

Example :  $x = 1000$  has only one significant figure. Again  $x = 378000$  has three significant figures.

**Rule 6.** All zeros on the right of the last non zero digit become significant, when they come from a measurement.

Example : Suppose distance between two stations is measured to be 3050 m. It has four significant figures. The same distance can be expressed as 3.050 km or  $3.050 \times 10^5$  cm. In all these expressions, number of significant figures continues to be four. Thus we conclude that change in the units of measurement of a quantity does not change the number of significant figures. By changing the position of the decimal point, the number of significant digits in the results does not change. Larger the number of significant figures obtained in a measurement, greater is the accuracy of the measurement. The reverse is also true.

(2) **Rounding off** : While rounding off measurements, we use the following rules by convention:

**Rule 1.** If the digit to be dropped is less than 5, then the preceding digit is left unchanged.

Example :  $x = 7.82$  is rounded off to 7.8, again  $x = 3.94$  is rounded off to 3.9.

**Rule 2.** If the digit to be dropped is more than 5, then the preceding digit is raised by one.

Example :  $x = 6.87$  is rounded off to 6.9, again  $x = 12.78$  is rounded off to 12.8.

**Rule 3.** If the digit to be dropped is 5 followed by digits other than zero, then the preceding digit is raised by one.

Example :  $x = 16.351$  is rounded off to 16.4, again  $x = 6.758$  is rounded off to 6.8.

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**Rule 4.** If digit to be dropped is 5 or 5 followed by zeros, then preceding digit is left unchanged, if it is even.

Example :  $x = 3.250$  becomes 3.2 on rounding off, again  $x = 12.650$  becomes 12.6 on rounding off.

**Rule 5.** If digit to be dropped is 5 or 5 followed by zeros, then the preceding digit is raised by one, if it is odd.

Example :  $x = 3.750$  is rounded off to 3.8. again  $x = 16.150$  is rounded off to 16.2.

### (3) Significant figure in calculation

(i) **Addition and subtraction** : In addition and subtraction the following points should be remembered :

(a) Every quantity should be changed into same unit.

(b) If a quantity is expressed in the power of 10, then all the quantities should be changed into power of 10.

(c) The result obtained after addition or subtraction, the number of figure should be equal to that of least, after decimal point.

(ii) **Multiplication and division**

(a) The number of significant figures will be same if any number is multiplied by a constant.

(b) The product or division of two significant figures, will contain the significant figures equal to that of least.

### Units for measurement.

The chosen standard of measurement of a quantity which has essentially the same nature as that of the quantity is called the **unit** of the quantity. Following are the important types of system for unit,

(1) C.G.S. System : Length (*centimetre*), Mass (*gram*), Time (*second*)

(2) M.K.S. System : Length (*metre*), Mass (*kilogram*), Time (*second*)

(3) F.P.S. System : Length (*foot*), Mass (*pound*), Time (*second*)

(4) S.I. System : The 11th general conference of weights and measures (October 1960) adopted *International system of units*, popularly known as the *SI units*. The SI has *seven basic units* from which all other units are derived called *derived units*. The standard prefixes which helps to reduce the basic units are now widely used.

**Dimensional analysis** : The seven basic quantities lead to a number of derived quantities such as pressure, volume, force, density, speed etc. The units for such quantities can be obtained by defining the derived quantity in terms of the base quantities using the base units. For example, speed (velocity) is expressed in distance/time. So the unit is  $m/s$  or  $ms^{-1}$ . The unit of force (mass  $\times$  acceleration) is  $kg\ ms^{-2}$  and the unit for acceleration is  $ms^{-2}$ .

### Seven basic SI units

Length	Mass	Time	Temperature	Electric Current	Luminous Intensity	Amount of substance
metre (m)	Kilogram (kg)	Second (s)	Kelvin (K)	Ampere (A)	Candela (Cd)	Mole (mol)

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### Derived Units

Physical quantity	Unit	Symbol
Area	square metre	$m^2$
Volume	cubic metre	$m^3$
Velocity	metre per second	$ms^{-1}$
Acceleration	metre per second square	$ms^{-2}$
Density	kilogram per cubic metre	$kg\ m^{-3}$
Molar mass	kilogram per mole	$kg\ mol^{-1}$
Molar volume	cubic metre per mole	$m^3\ mol^{-1}$
Molar concentration	mole per cubic metre	$mol\ m^{-3}$
Force	newton (N)	$kg\ m\ s^{-2}$
Pressure	pascal (Pa)	$N\ m^{-2}$
Energy work	joule (J)	$kg\ m^2\ s^{-2}, Nm$

### Standard prefixes use to reduce the basic units

Multiple	Prefix	Symbol	Submultiple	Prefix	Symbol
$10^{24}$	yotta	Y	$10^{-1}$	deci	d
$10^{21}$	zetta	Z	$10^{-2}$	centi	c
$10^{18}$	exa	E	$10^{-3}$	milli	m
$10^{15}$	peta	P	$10^{-6}$	micro	$\mu$
$10^{12}$	tera	T	$10^{-9}$	nano	n
$10^9$	giga	G	$10^{-12}$	pico	p
$10^6$	mega	M	$10^{-15}$	femto	f
$10^3$	kilo	k	$10^{-18}$	atto	a
$10^2$	hecto	h	$10^{-21}$	zeto	z
$10^1$	deca	da	$10^{-24}$	yocto	y

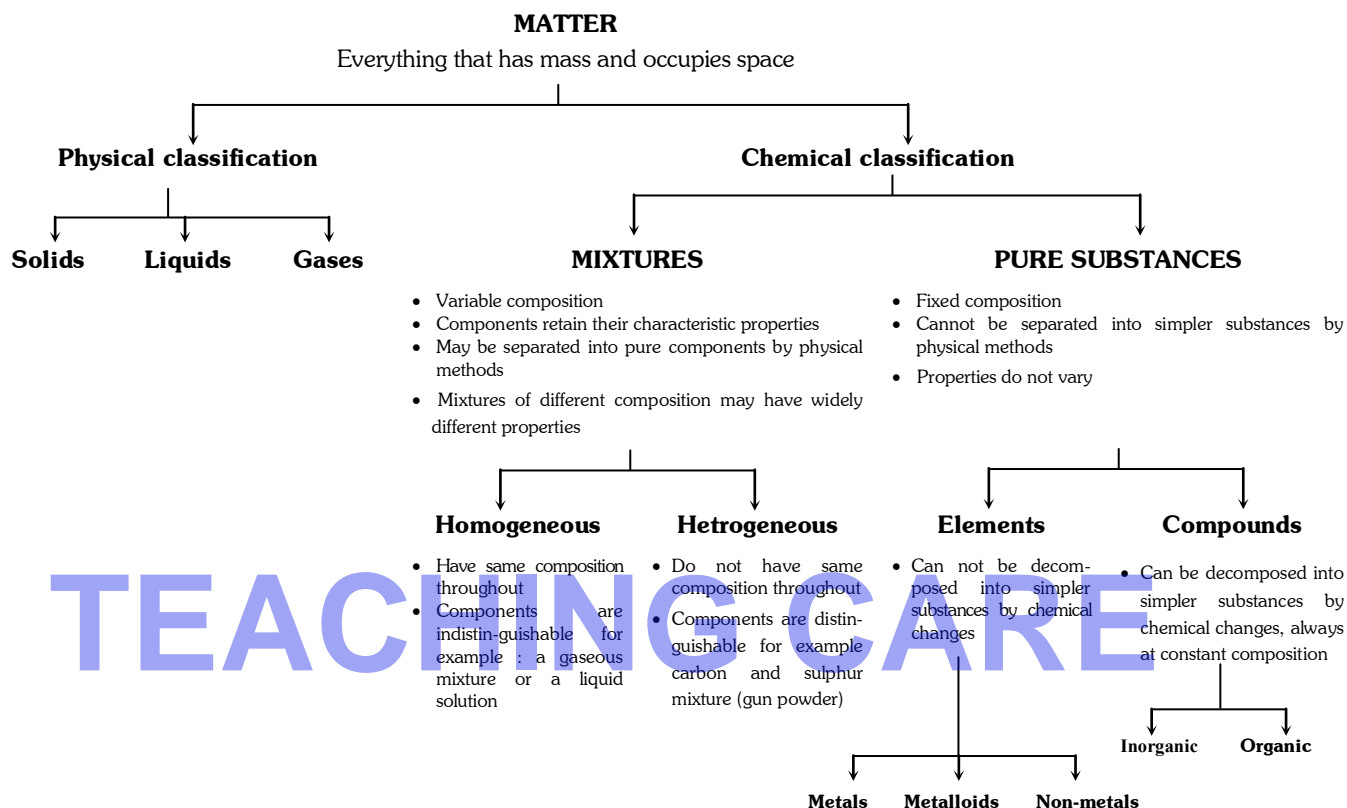
### Conversion factors

$1\ m = 39.37\ inch$	$1\ cal = 4.184\ J$	$1\ e.s.u. = 3.3356 \times 10^{-10}\ C$	$1\ mole\ of\ a\ gas = 22.4\ L\ at\ STP$
$1\ inch = 2.54\ cm$	$1\ eV = 1.602 \times 10^{-19}\ J$	$1\ dyne = 10^{-5}\ N$	$1\ mole\ a\ substance = N_0\ molecules$
$1\ litre = 1000\ mL$	$1\ eV/atom = 96.5\ kJ\ mol^{-1}$	$1\ atm = 101325\ Pa$	$1\ g\ atom = N_0\ atoms$
$1\ gallon\ (US) = 3.79\ L$	$1\ amu = 931.5016\ MeV$	$1\ bar = 1 \times 10^5\ N\ m^{-2}$	$t\ (^{\circ}F) = 9/5\ t\ (^{\circ}C) + 32$
$1\ lb = 453.59237\ g$	$1\ kilo\ watt\ hour = 3600\ kJ$	$1\ litre\ atm = 101.3\ J$	$1\ g\ cm^{-3} = 1000\ kg\ m^{-3}$
$1\ newton = 1\ kg\ m\ s^{-2}$	$1\ horse\ power = 746\ watt$	$1\ year = 3.1536 \times 10^7\ s$	$1\ \text{\AA} = 10^{-10}\ m$
$1\ J = 1\ Nm = 1\ kg\ m^2\ s^{-2}$	$1\ joule = 10^7\ erg$	$1\ debye\ (D) = 1 \times 10^{-18}\ esu\ cm$	$1\ nm = 10^{-9}\ m$

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### Matter and its classification.

Matter is the physical material of the universe which occupies space and has mass e.g., water, sugar, metals, plants etc. Matter can be classified as,



### Separation of mixtures or purification of an impure substance.

Each component of a mixture retains its own properties and thus a mixture can be separated into its components by taking advantages of the difference in their physical and chemical properties. The different methods which are employed to separate the constituents from a mixture to purify an impure sample of a substance are,

(1) **Sedimentation** : It is also called gravity separation. It is used for a mixture in which one component is a liquid and the other is insoluble solid heavier than the liquid. Example : Sand dispersed in water.

(2) **Filtration** : It is used for a mixture containing two components one of which is soluble in a particular solvent and the other is not. Example : (i) A mixture of salt and paper using water as solvent (ii) A mixture of sand and sulphur using carbon disulphide as solvent. (iii) A mixture of glass powder and sugar, using water as a solvent in which sugar dissolves but glass does not. (iv) A mixture of sand and sulphur, using carbon disulphide as the solvent in which sulphur dissolves but sand does not.

(3) **Sublimation** : It is used for a mixture containing a solid component, which sublimes on heating from non-volatile solids. Example : A mixture of sand + naphthalene or powdered glass +  $NH_4Cl$  / camphor / iodine etc. can be separated by the method of sublimation because substances like naphthalene,  $NH_4Cl$ , iodine, camphor etc. form sublimes whereas sand, glass etc. do not.

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(4) **Evaporation** : It is used for a mixture in which one component is a non-volatile soluble salt and other is a liquid. Example : Sodium chloride dissolved in sea-water.

(5) **Crystallization** : It is a most common method for a mixture containing solid components and based upon the differences in the solubilities of the components of the mixture into a solvent. For separation, a suitable solvent is first selected. It is of two types :

(i) **Simple crystallization** : It is applicable when there is a large difference in the solubilities of different components of a mixture.

(ii) **Fractional crystallization** : It is applicable when there is a small difference in the solubilities of different components of a mixture in the same solvent. Example :  $K_2Cr_2O_7$  and  $KCl$ . Here  $K_2Cr_2O_7$  is less soluble in water and hence crystallizes first. A series of repeated crystallization separates the two components in pure form.

(6) **Distillation** : It is the most important method for purifying the liquids. It involves the conversion of a liquid to its vapours on heating (vaporisation) and then cooling the vapours back into the liquid (condensation). It can be used to separate, (i) A solution of a solid in a liquid. e.g., aqueous copper sulphate solution. (ii) A solution of two liquids whose boiling points are different. Several methods of distillation are employed.

(i) **Simple distillation** : It is used only for such liquids which boil without decomposition at atmospheric pressure and contain non-volatile impurities. Example : (a) *Pure water from saline water.* (b) Benzene from toluene.

(ii) **Fractional distillation** : It is used for the separation and purification of a mixture of two or more miscible liquids having different boiling points. The liquid having low boiling point vaporises first, gets condensed and is collected in the receiver. The temperature is then raised to the boiling point of second so that the second liquid vaporises and is collected in other receiver. *If two liquids present in a mixture have their boiling points closer to each other, a fractionating column is used.* Example : (a) *Crude petroleum is separated into many useful products such as lubricating oil, diesel oil, kerosene and petrol by fractional distillation.* (b) A mixture of acetone and methyl alcohol.

(iii) **Vacuum distillation or distillation under reduced pressure** : It is used for such liquids which decompose on heating to their boiling points. At reduced pressure, the boiling point of liquid is also reduced. Example : *Glycerol is distilled under pressure as it decomposes on heating to its boiling point.*

(iv) **Steam distillation** : It is used for liquids which are partially miscible with water, volatile in steam. e.g., aniline, oils etc. are purified by steam distillation. The principle involved is of reduced pressure distillation. If  $P_w$  and  $P_l$  are vapour pressures of water and liquid at distillation temperature, then  $P_w + P_l = P = 76 \text{ cm}$  i.e., atmospheric pressure. Thus, a liquid boils at relatively low temperature than its boiling point in presence of steam.

Example : *Some solids like naphthalene, o-nitrophenol which are steam volatile can be purified. Nitrobenzene, chlorobenzene, essential oils are also extracted or separated by this process*

(7) **Solvent extraction** : It is used for the separation of a compound from its solution by shaking with a suitable solvent. The extraction follows Nernst distribution law. The solvent used must be insoluble with other phase in which compound is present as well as the compound should be more soluble in solvent. The extraction becomes more efficient if the given extracting liquid is used for more number of extractions with smaller amounts than using it once with all extracting liquid. Example : (i) Aqueous solution of benzoic acid by using benzene. (ii) Aqueous solution of Iodine by using chloroform or carbon tetrachloride. (iii) Flavour of tea from the tea leaves by boiling with water.

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(8) **Magnetic separation** : It is used for a mixture in which one component is magnetic while the other is non-magnetic. Example : iron ore from the non-magnetic impurities.

(9) **Chromatography** : It is based on the differences in the rates at which different components of a mixture are absorbed on a suitable solvent. It is used in separation, isolation, purification and identification of a substance. *It was proposed by a Russian botanist Tswett.*

(10) **Atmolysis** : It is used for separation of the mixture of gases or vapours. It is based on the difference in their rates of diffusion through a porous substance. Example :  $H_2$ ,  $SO_2$ ,  $CH_4$  and  $O_2$ ,  $U^{235}$  &  $U^{238}$  in the form of their hexa-fluorides.

(11) **Electrophoresis** : It is based upon the differences in the electrical mobility of the different components of the mixture.

(12) **Ultracentrifugation** : It is based upon the difference in sedimentation velocity of the components in a centrifugal field.

### Laws of chemical combination.

Various chemical reactions take place according to the certain laws, known as the *Laws of chemical combination*. These are as follows,

(1) **Law of conservation of mass** : *It was proposed by Lavoisier and verified by Landolt. According to this law, Matter is neither created nor destroyed in the course of chemical reaction though it may change from one form to other.* The total mass of materials after a chemical reaction is same as the total mass before reaction.

Example : A reaction between  $AgNO_3$  solution and  $KI$  solution.



According to the modified statement of the law, *The total sum of mass and energy of the system remains constant.*

(2) **Law of constant or definite proportion** : *It was proposed by Proust. According to this law, A pure chemical compound always contains the same elements combined together in the fixed ratio of their weights whatever its methods of preparation may be.*

Example :  $CO_2$  can be formed by either of the following processes:



$CO_2$  is collected separately as a product of each reaction and the analysis of  $CO_2$  of each collection reveals that it has the combination ratio of carbon and oxygen as 12 : 32 by weight.

(3) **Law of multiple proportion** : *It was proposed by Dalton and verified by Berzelius. According to this law, When two elements A and B combine to form more than one chemical compounds then different weights of A, which combine with a fixed weight of B, are in proportion of simple whole numbers.*

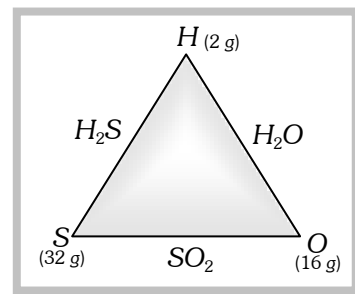
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Example : Nitrogen forms as many as five stable oxides. The analysis of these oxides ( $N_2O$ ,  $NO$ ,  $N_2O_3$ ,  $N_2O_4$  and  $N_2O_5$ ) reveals that for 28 gm. nitrogen, the weight of oxygen that combines is in the ratio 16 : 32 : 48 : 64 : 80 i.e., 1 : 2 : 3 : 4 : 5 in  $N_2O$ ,  $NO$ ,  $N_2O_3$ ,  $N_2O_4$  and  $N_2O_5$  respectively.

(4) **Law of equivalent proportion or law of reciprocal proportion** : It was proposed by Richter. According to this law, The weights of the two or more elements which separately react with same weight of a third element are also the weights of these elements which react with each other or in simple multiple of them.

Example : Formation of  $H_2S$ ,  $H_2O$  and  $SO_2$  can be done as follows,

(i) Hydrogen combines with sulphur forming hydrogen sulphide; 2gm. of hydrogen reacts with 32gm of sulphur. (ii) Hydrogen combines oxygen forming water; 2 gm. of hydrogen reacts with 16 gm. of oxygen. (iii) Sulphur combines with oxygen forming sulphur dioxide; 32 gm. of sulphur reacts with 32 gm. of oxygen i.e., in the ratio 32 : 32. This ratio is double of the ratio weights of these elements which combine with 2 gm. of hydrogen. i.e.,  $32/16 : 32/32 = 2 : 1$



Law of Reciprocal proportion can be used to obtain equivalent weights of elements. As elements always combine in terms of their equivalent weights.

(5) **Gay-Lussac's Law**: It was proposed by Gay-Lussac and is applicable only for gases. According to this law, When gases combine, they do so in volumes, which bear a simple ratio to each other and also to the product formed provided all gases are measured under similar conditions. The Gay-Lussac's law, was based on experimental observation.

Example : (i) Reaction between hydrogen and oxygen.  $H_{2(g)} + \frac{1}{2} O_{2(g)} \longrightarrow H_2O_{(v)}$

One volume of  $H_2$  reacts with half volume of  $O_2$  to give one volume of  $H_2O$ .

(ii) Reaction between nitrogen and hydrogen.  $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$

One volume of  $N_2$  reacts with three volumes of  $H_2$  to give two volumes of  $NH_3$ .

### Important hypothesis.

(1) **Atomic hypothesis** : Keeping in view various law of chemical combinations, a theoretical proof for the validity of different laws was given by John Dalton in the form of hypothesis called **Dalton's atomic hypothesis**. Postulates of Dalton's hypothesis is as follows,

(i) Each element is composed of extremely small particles called **atoms** which can take part in chemical combination.

(ii) All atoms of a given element are identical i.e., atoms of a particular element are all alike but differ from atoms of other element.

(iii) Atoms of different elements possess different properties (including different masses).

(iv) Atoms are indestructible i.e., atoms are neither created nor destroyed in chemical reactions.

(v) Atoms of elements take part to form molecules i.e., compounds are formed when atoms of more than one element combine.

(vi) In a given compound, the relative number and kinds of atoms are constant.

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(2) **Modern atomic hypothesis** : The main modifications made in Dalton's hypothesis as a result of new discoveries about atoms are,

(i) Atom is no longer considered to be indivisible.

(ii) Atoms of the same element may have different atomic weights. e.g., isotopes of oxygen  $O^{16}$ ,  $O^{17}$ , and  $O^{18}$ .

(iii) Atoms of different element may have same atomic weights. e.g., isobars  $Ca^{40}$  and  $Ar^{40}$ .

(iv) *Atom is no longer indestructible. In many nuclear reactions, a certain mass of the nucleus is converted into energy in the form of  $\alpha$ ,  $\beta$  and  $\gamma$  rays.*

(v) Atoms may not always combine in simple whole number ratios. e.g., in sucrose ( $C_{12}H_{22}O_{11}$ ), the elements carbon, hydrogen and oxygen are present in the ratio of 12 : 22 : 11 and the ratio is not a simple whole number ratio.

(3) **Berzelius hypothesis** : "Equal volumes of all gases contain equal number of atoms under same conditions of temperature and pressure". When applied to law of combining volumes, this hypothesis predicts that atoms are divisible and hence it is contrary to Dalton's hypothesis.

(4) **Avogadro's hypothesis** : "Equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules." Avogadro hypothesis has been found to explain as follows :

(i) Provides a method to determine the atomic weight of gaseous elements.

(ii) *Provides a relationship between vapour density (V.D.) and molecular masses of substances.*

$$\text{Vapour density} = \frac{\text{Volume of definite amount of a gas}}{\text{Volume of same amount of hydrogen}}$$

$$\text{or Vapour density} = \frac{\text{Mass of 'n' molecule of a gas}}{\text{Mass of 'n' molecule of hydrogen}}$$

$$\text{Vapour density} = \frac{\text{Mass of 1 molecule of a gas}}{\text{Mass of 1 molecule of hydrogen}}$$

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

(iii) *It helps in the determination of mass of fixed volume of a particular gas.*

$$\text{Vapour density} = \frac{\text{Mass of 1 molecule of gas}}{\text{Mass of 1 molecule of } H_2} = \frac{\text{Mass of 1 ml of gas}}{\text{Mass of 1 ml of } H_2} = \frac{\text{Mass of 1 ml of gas}}{0.0000897}$$

$$(\because 1 \text{ ml } H_2 = 0.0000897 \text{ gm.}) \text{ at NTP}$$

$$\therefore \text{Mass of 1 ml gas} = \text{V.D.} \times 0.0000897 \text{ gm.}$$

(iv) *It also helps in the determination of molar volume at N.T.P.*

$$\therefore \text{V.D.} \times 0.0000897 \text{ gm. gas has volume} = 1 \text{ ml}$$

$$\therefore 2 \times \text{V.D. (i.e., molecular mass) gm. has volume} = \frac{1 \times 2 \times \text{V.D.}}{\text{V.D.} \times 0.0000897} \text{ ml} = 22400 \text{ ml}$$

$$\therefore \text{Molar mass of a gas or its 1 mole occupies } 22.4 \text{ L volume at S.T.P.}$$



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**Note :** 22.4 litres of any gas at S.T.P. weigh equal to the molecular weight of the gas expressed in grams. This is called **Gram-molecular volume** (G.M.V.) law.

(v) It helps in determination of molecular formulae of gases and is very useful in gas analysis. By knowing the molecular volumes of reactants and products of reaction, molecular composition can be determined easily.

### Atomic, Molecular and Equivalent masses.

(1) **Atomic Mass** : It is the average relative mass of atom of element as compared with an atom of carbon-12 isotope taken as 12.

$$\text{Atomic mass} = \frac{\text{Average mass of an atom}}{1/12 \times \text{Mass of an atom of } C^{12}}$$

**Average atomic mass** : If an element exists in two isotopes having atomic masses 'a' and 'b' in the ratio  $m : n$ , then average atomic mass =  $\frac{(m \times a) + (n \times b)}{m + n}$ . Since the atomic mass is a ratio, it has no units and is expressed in *amu*,

1 *amu* =  $1.66 \times 10^{-24}$  g. One atomic mass unit (*amu*) is equal to  $\frac{1}{12}$ th of the mass of an atom of carbon-12 isotope.

**Gram atomic mass (GAM)** : Atomic mass of an element expressed in grams is called *Gram atomic mass* or *gram atom* or *mole atom*.

(i) Number of gram atoms or mole atoms =  $\frac{\text{Mass of an element}}{\text{GAM}}$

(ii) Mass of an element in *gm.* = Number of *gm.* atom  $\times$  GAM

(iii) Number of atoms in 1 GAM =  $6.02 \times 10^{23}$

$$\therefore \text{Number of atoms in a given substance} = \text{No. of GAM} \times 6.02 \times 10^{23} = \frac{\text{Mass}}{\text{GAM}} \times 6.02 \times 10^{23}$$

(iv) Number of atoms in 1 *gm* of element =  $\frac{6.02 \times 10^{23}}{\text{Atomic mass}}$

(v) Mass of one atom of the element (in *gm.*) =  $\frac{\text{GAM}}{6.02 \times 10^{23}}$

### Methods of determination of atomic mass

(i) **Dulong and Pettit's method** : According to Dulong and Pettit's law

$$\text{Atomic mass} \times \text{Specific heat} = 6.4 \text{ (approx.)}$$

$$\text{Atomic mass (approx.)} = \frac{6.4}{\text{Specific heat (in cal.)}}$$

This law is applicable to solid elements only except Be, B, C and Si.

$$\text{Atomic mass} = \text{Equivalent mass} \times \text{Valency}$$

$$\text{Valency} = \frac{\text{Approximate atomic mass}}{\text{Equivalent mass}}$$

(ii) **Vapour density method** : It is suitable for elements whose chlorides are volatile.

$$\text{Valency of the element} = \frac{\text{Molecular mass of chloride}}{\text{Equivalent mass of chloride}} = \frac{2 \times \text{Vapour density of chloride}}{\text{Equivalent mass of metal} + 35.5}$$

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$$\text{Atomic mass} = \text{Equivalent mass of metal} \times \text{Valency}$$

(iii) **Specific heat method** : It is suitable only for gases. The two types of specific heats of gases are  $C_p$  (at constant pressure) and  $C_v$  (at constant volume). Their ratio is known as  $\gamma$  whose value is constant (1.66 for monoatomic, 1.40 for diatomic and 1.33 for triatomic gases).

$$\text{Atomic mass of a gaseous element} = \frac{\text{Molecular mass}}{\text{Atomicity}}$$

(iv) **Volatile chloride method** : Different steps are given below

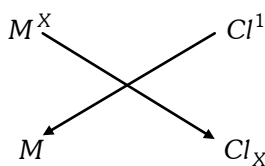
*Step I.* The element (M) whose atomic mass is to be determined is converted into its volatile chloride whose vapour density is determined by **Victor Meyer method**.

Thus, Molecular mass of the chloride =  $2 \times V. D.$

*Step II.* Equivalent mass of the element (M) of valency X is determined as usual.

Atomic mass of the element = Equivalent mass of the element  $\times X = Z \times X$

*Step III.* The formula of the volatile chloride is derived as below,



*Step IV.* Molecular mass of the chloride =  $(Z \times X) + (35.5 \times X) = X(Z + 35.5)$

$$\text{From (I) and (IV)} \quad X(Z + 35.5) = 2 \times \text{Vapour density} \quad \text{or} \quad X = \frac{2 \times V.D.}{Z + 35.5}$$

Thus, Atomic mass of the element =  $Z \times X$

(v) **Isomorphism method** : It is based on law of isomorphism which states that compounds having identical crystal structure have similar constitution and chemical formulae.

Example :  $K_2SO_4$ ,  $K_2CrO_4$  and  $K_2SeO_4$  (valency of S, Cr, Se = 6),

$ZnSO_4 \cdot 7H_2O$ ,  $MgSO_4 \cdot 7H_2O$ ,  $FeSO_4 \cdot 7H_2O$  (valency of Zn, Mg, Fe = 2).

### Applications of isomorphism

- The valencies of two elements forming isomorphism salts are essentially same. Thus if valency of one of the elements is known that of other will be the same.

- Masses of two elements, that combine with the same mass of other elements in their respective isomorphous compounds, are in the ratio of their atomic masses *i.e.*,

$$\frac{\text{Mass of element A that combines with a certain mass of other elements}}{\text{Mass of element B that combines with the same mass of other elements}} = \frac{\text{Atomic mass of A}}{\text{Atomic mass of B}}$$

- By knowing the percentage of two elements of their isomorphous compound and atomic mass of one element, atomic mass of other element can be calculated.

(2) **Molecular mass** : Molecular mass of a molecule, of an element or a compound may be defined as a number which indicates how many times heavier is a molecule of that element or compound as compared with

$\frac{1}{12}$  of the mass of an atom of carbon-12. Molecular mass is a ratio and hence has no units. It is expressed in *a.m.u.*

## SOME BASIC CONCEPTS OF CHEMISTRY

$$\text{Molecular mass} = \frac{\text{Mass of one molecule of the substance}}{1/12 \times \text{Mass of one atom of C - 12}}$$

Actual mass of one molecule = Molecular mass  $\times 1.66 \times 10^{-24}$  gm.

Molecular mass of a substance is the additive property and can be calculated by adding the atomic masses present in one molecule.

**Gram molecular mass (GMM) and Gram molar volume :** Molecular mass of an element or compound when expressed in gm. is called its gram molecular mass, gram molecule or mole molecule.

$$\text{Number of gm molecules or mole molecules} = \frac{\text{Mass of substances}}{\text{GMM}}$$

Mass of substances in gm = Number of gm. molecules  $\times$  GMM

Volume occupied by one mole of any gas at STP is called **Gram molar volume**. The value of gram molar volume is 22.4 litres. *Volume of 1 mole of any gas at STP = 22.4 litres*

### Expression for mass and density

Mass of 11.2L of any gas at STP = V.D. of that gas in gm.

$$\text{Density of a gas at NTP} = \frac{\text{Mol. mass in gm.}}{22400 \text{ ml}}$$

### Important generalisations

Number of atoms in a substance = Number of GMM  $\times 6.02 \times 10^{23} \times$  Atomicity

Number of electrons in given substance = Number of GMM  $\times 6.02 \times 10^{23} \times$  Number of electrons

**Methods of determination of molecular mass :** following methods are used to determine molecular mass

(i) **Diffusion method** (For gases) : The ratio of rates of diffusion of two gases is inversely proportional to the square root of their molecular masses.

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

(ii) **Vapour density method** (For gases only) : Mass of a fixed volume of the vapour is compared with the mass of the same volume of hydrogen under same conditions. The ratio of these masses is called **Vapour density** or **Relative density**.

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

**Note :** \* Vapour density of a substance increases due to molecular association (e.g.,  $\text{CH}_3\text{COOH}$ ,  $\text{HF}$ ) and decreases due to dissociation (e.g.,  $\text{NH}_4\text{Cl}$ ,  $\text{PCl}_5$ , etc.)

(iii) **Victor Meyer method** (For volatile liquids or solids) : It is based on Dalton's law of partial pressure and Avogadro's hypothesis (gram molar volume).

22400 ml of vapours of a substance = Molecular mass of that substance

(iv) **Colligative property method** (For non-volatile solids) : Discussed in colligative properties of solutions.

### Average atomic mass and molecular mass

## SOME BASIC CONCEPTS OF CHEMISTRY

$$\bar{A} \text{ (Average atomic mass)} = \frac{\sum A_i X_i}{\sum X_{\text{total}}}; \bar{M} \text{ (Average molecular mass)} = \frac{\sum M_i X_i}{\sum X_{\text{total}}}$$

Where  $A_1, A_2, A_3, \dots$  are atomic mass of species 1, 2, 3, ... etc. with % ratio as  $X_1, X_2, X_3, \dots$  etc. Similar terms are for molecular masses.

**(3) Equivalent mass :** The number of parts by mass of a substance that combines with or displaces 1.008 parts by mass of hydrogen or 8.0 parts of oxygen or 35.5 parts of chlorine is called its *equivalent mass* (EM). On the other hand *quantity of a substance in grams numerically equal to its equivalent mass is called its gram equivalent mass (GEM) or gram equivalent.*

$$\text{Number of GEM} = \frac{\text{Mass of the substance in grams}}{\text{GEM of the substance}}$$

### Expressions for equivalent mass (EM)

(i) EM of an element =  $\frac{\text{Atomic mass}}{\text{Valency}}$

(ii) EM of an acid =  $\frac{\text{Molecular mass}}{\text{Basicity}}$

(Basicity of acid is the number of replaceable hydrogen atoms in one molecule of the acid).

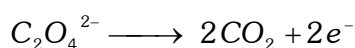
(iii) EM of a base =  $\frac{\text{Molecular mass}}{\text{Acidity}}$

(Acidity of a base is the number of replaceable -OH groups in one molecule of the base).

(iv) EM of a salt =  $\frac{\text{Formula mass}}{\text{Total positive or negative charge}}$

(v) EM of a reducing agent =  $\frac{\text{Formula mass}}{\text{Number of electrons lost per molecule or Total change in oxidation number}}$

Example : Eq. mass of oxalic acid ( $C_2H_2O_4$ )



(O.N. of C = +3)                      (O.N. of C = +4)

∴ Number of electrons lost = 2

or Change in O.N. per C atom (from +3 to +4) = 1

∴ Total change in O.N. =  $1 \times 2 = 2$

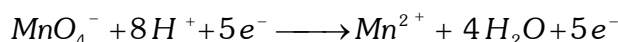
∴ Equivalent mass of  $C_2H_2O_4 = \frac{90}{2} = 45$

(vi) EM of an oxidising agent =  $\frac{\text{Formula mass}}{\text{Number of electrons gained per molecule or Total change in O.N.}}$

Equivalent mass of common oxidising agent changes with the medium of the reaction.

Example : (a) Equivalent mass of  $KMnO_4$  in acidic medium

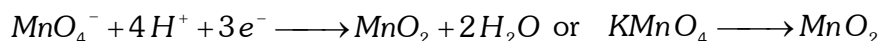
## SOME BASIC CONCEPTS OF CHEMISTRY



Change in O.N. of Mn from + 7 to + 2 = + 5

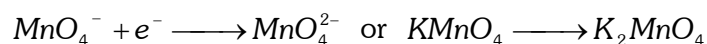
$$\therefore \text{Equivalent mass of } \text{KMnO}_4 = \frac{158}{5} = 31.60$$

(b) Equivalent mass of  $\text{KMnO}_4$  in neutral medium



$$\therefore \text{Equivalent mass of } \text{KMnO}_4 = \frac{158}{3} = 52.67$$

(c) Equivalent mass of  $\text{KMnO}_4$  in basic medium



$$\text{Equivalent mass of } \text{KMnO}_4 = \frac{158}{1} = 158$$

$$\text{(vii) EM of radicals} = \frac{\text{Formula mass of radical}}{\text{Number of units of charge}}$$

Example : Equivalent mass of  $\text{SO}_4^{2-} = \frac{96}{2} = 48$

### Methods of determination of equivalent mass

**(i) Hydrogen displacement method :** The mass of metal which displaces 11200 ml of hydrogen at NTP from an acid, alkali or alcohol is the equivalent mass of the metal.

$$\text{(a) Equivalent mass of metal} = \frac{\text{Mass of metal}}{\text{Mass of H}_2 \text{ displaced}} \times 1.008 = \frac{W}{M} \times 1.008\text{g}$$

$$\text{(b) Equivalent mass of metal} = \frac{\text{Mass of metal}}{\text{Vol. (ml) of H}_2 \text{ displaced at STP}} \times 11200 = \frac{W}{V} \times 11200$$

*This method is useful for metals which can displace hydrogen from acids or can combine with hydrogen (Mg, Zn, Na, Ca etc.)*

**(ii) Oxide formation method :** The mass of the element which combines with 8 grams of oxygen is the equivalent mass of the element.

$$\text{(a) Equivalent mass of metal} = \frac{\text{Mass of metal}}{\text{Mass of oxygen}} \times 8$$

$$\text{(b) Equivalent mass of metal} = \frac{\text{Mass of metal}}{\text{Vol. of O}_2 \text{ at S.T.P. in ml}} \times 5600$$

**(iii) Chloride formation method :** The mass of an element which reacts with 35.5 gm. of chlorine is the equivalent mass of that element.

$$\text{(a) Equivalent mass of metal} = \frac{\text{Mass of metal}}{\text{Mass of chlorine}} \times 35.5$$

$$\text{(b) Equivalent mass of metal} = \frac{\text{Mass of metal}}{\text{Vol. of Cl}_2 \text{ (in ml.) at STP}} \times 11200$$

## SOME BASIC CONCEPTS OF CHEMISTRY

(iv) **Neutralisation method** : (For acids and bases).

$$\boxed{\text{Equivalent mass of acid (or base)} = \frac{W}{V \times N}}, \text{ Where } W = \text{Mass of acid or base in gm.}, V = \text{Vol. of base or}$$

acid in litre required for neutralisation and  $N$  is Normality of base or acid

(v) **Metal displacement method** : It is based on the fact that one gm. equivalent of a more electropositive metal displaces one gm equivalent of a less electropositive metal from its salt solution.

$$\frac{\text{Mass of metal added}}{\text{Mass of metal displaced}} = \frac{\text{Eq. mass of metal added}}{\text{Eq. mass of metal displaced}}; \quad \frac{W_1}{W_2} = \frac{E_1}{E_2}$$

(vi) **Electrolytic method** : The quantity of substance that reacts at electrode when 1 faraday of electricity is passed is equal to its gram equivalent mass.

$$\text{Gram equivalent mass} = \text{Electrochemical equivalent} \times 96500$$

The ratio of masses of two metals deposited by the same quantity of electricity will be in the ratio of their equivalent masses.

$$\boxed{\frac{W_1}{W_2} = \frac{E_1}{E_2}}$$

(vii) **Double decomposition method** :  $AB + CD \longrightarrow AD \downarrow + CB$

$$\frac{\text{Mass of compound } AB}{\text{Mass of compound } AD} = \frac{\text{Eq. mass of } A + \text{Eq. mass of } B}{\text{Eq. mass of } A + \text{Eq. mass of } D}$$

$$\text{or } \frac{\text{Mass of salt taken } (W_1)}{\text{Mass of ppt. obtained } (W_2)} = \frac{\text{Eq. mass of salt } (E_1)}{\text{Eq. mass of salt in ppt. } (E_2)}$$

(viii) **Conversion method** : When one compound of a metal is converted to another compound of the same metal, then

$$\frac{\text{Mass of compound I } (W_1)}{\text{Mass of compound II } (W_2)} = \frac{E + \text{Eq. mass of radical I}}{E + \text{Eq. mass of radical II}} \quad (E = \text{Eq. mass of the metal})$$

(ix) **Volatile chloride method**

$$\text{Valency of metal} = \frac{2 \times \text{V.D. of Chloride}}{\text{Eq. mass of metal chloride}} = \frac{2 \times \text{V.D.}}{E + 35.5} \quad \therefore E = \frac{2 \times \text{V.D. of Chloride}}{\text{Valency}} - 35.5$$

(x) **Silver salt method** (For organic acids)

$$\text{Equivalent Mass of acid} = \frac{108 \times \text{Mass of silver salt}}{\text{Mass of Ag metal}} - 107$$

$$\text{Molecular mass of acid} = \text{Equivalent mass of acid} \times \text{Basicity}$$

**Example: 1** Oxygen contains 90%  $O^{16}$  and 10%  $O^{18}$ . Its atomic mass is  
 (a) 17.4 (b) 16.2 (c) 16.5 (d) 17

[KCET 1998]

**Solution:** (b) Average atomic mass of oxygen =  $\frac{90 \times 16 + 10 \times 18}{100} = 16.20$

**Example: 2** The total number of electrons present in 1.6 gm. of methane is

[IIT 1976; Roorkee 1985; CPMT 1987, 92]

## SOME BASIC CONCEPTS OF CHEMISTRY

- (a)  $6.02 \times 10^{23}$       (b)  $6.02 \times 10^{22}$       (c)  $6.02 \times 10^{21}$       (d)  $4.02 \times 10^{20}$

**Solution:** (a) Number of GMM of methane in 1.6 gm. of methane

$$= \frac{\text{Mass of methane}}{\text{Mol. mass of methane}} = \frac{1.6}{16} = 0.1$$

Number of electrons in 1 molecule of methane ( $CH_4$ ) =  $6 + 4 = 10$

Hence, Number of electrons in 1.6 gm. of methane

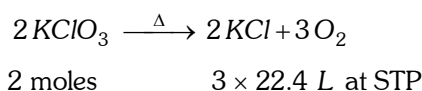
$$= \text{Number of GMM} \times 6.02 \times 10^{23} \times \text{No. of electrons}$$

$$= 0.1 \times 6.02 \times 10^{23} \times 10 = 6.02 \times 10^{23}$$

**Example: 3**  $KClO_3$  on heating decomposes to  $KCl$  and  $O_2$ . The volume of  $O_2$  at STP liberated by 0.1 mole  $KClO_3$  is [BIT 1991]

- (a) 4.36 L      (b) 3.36 L      (c) 2.36 L      (d) none of these

**Solution:** (b) On heating  $KClO_3$  dissociates as:



$\therefore$  2 moles of  $KClO_3$  on heating produces = 67.2 L of  $O_2$  at STP

$\therefore$  0.1 mole of  $KClO_3$  on heating produces =  $\frac{67.2}{2} \times 0.1$  L = 3.36 L of  $O_2$  at STP

**Example: 4** At S.T.P. the density of  $CCl_4$  vapour in g/L will be nearest to [CBSE PMT 1988]

- (a) 6.84      (b) 3.42      (c) 10.26      (d) 4.57

**Solution:** (a) 1 mole of  $CCl_4$  vapour =  $12 + 4 \times 35.5 = 154$  gm = 22.4 L at S.T.P.

$$\therefore \text{Density} = \frac{154}{22.4} \text{ gmL}^{-1} = 6.875 \text{ gmL}^{-1}$$

**Example: 5** 4.4 g of an unknown gas occupies 2.24 litres of volume at NTP. The gas may be [MP PMT 1995]

- (a) Carbon dioxide      (b) Carbon monoxide      (c) Oxygen      (d) Sulphur dioxide

**Solution:** (a) Mass of 2.24 litres gas = 4.4 gm.

$\therefore$  Mass of 22.4 litres gas = 44 gm.

Here, out of four given gases, the molecular mass of only carbon dioxide is 44 gm.

**Example: 6** One gm. of a metal carbonate gave 0.56 gm. of its oxide on heating. The equivalent mass of the metal is

- (a) 40      (b) 30      (c) 20      (d) 10

**Solution:** (c)  $\frac{\text{Mass of metal carbonate}}{\text{Mass of metal oxide}} = \frac{E + \text{Eq. mass of } CO_3^{2-}}{E + \text{Eq. mass of } O_2^{2-}}$  or  $\frac{1.0}{0.56} = \frac{E + 30}{E + 8}$

$$\text{or } (E + 8) \times 1.0 = (E + 30)(0.56) \quad (\text{Where } E \text{ is the equivalent mass of the metal})$$

$$\therefore E = 20$$

### The mole concept.

The mole (abbreviated as mol) is the SI base unit for a amount of a chemical species. It is always associated with a chemical formula and refers to Avogadro's number ( $6.022 \times 10^{23}$ ) of particles represented by the formula. It is designated as  $N_0$ . Thus, 12 molecules of  $HCl$  is a dozen, 144 molecules of  $HCl$  is a gross and  $6.022 \times 10^{23}$  molecules of  $HCl$  is a mole.

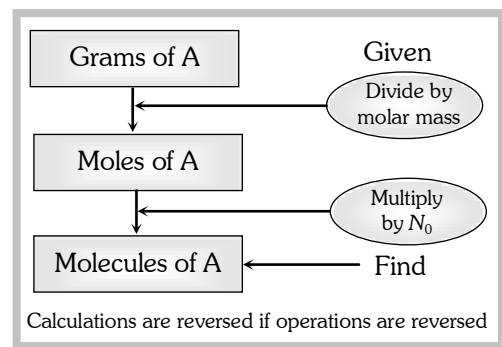
## SOME BASIC CONCEPTS OF CHEMISTRY

1 mole of a substance =  $6.022 \times 10^{23}$  species

The molar mass of a substance is the mass in grams of 1 mole of that substance.

$$\text{Mole of a substance} = \frac{\text{mass in grams}}{\text{molar mass}}$$

Under STP conditions when temperature is 273K and pressure is 1 atm, volume of one mole of an ideal gas is 22.4L



**Example: 7** The number of gram molecules of oxygen in  $6.02 \times 10^{24}$  CO molecules is

[IIT 1990]

- (a) 10 gm molecules      (b) 5 gm molecules      (c) 1 gm molecules      (d) 0.5 gm molecules

**Solution:** (b)  $6.02 \times 10^{23}$  molecules = 1 mole of CO

$$\begin{aligned} \therefore 6.02 \times 10^{24} \text{ CO molecules} &= 10 \text{ moles of CO} \\ &= 10 \text{ gms of Oxygen atom} = 5 \text{ gm molecules of } O_2 \end{aligned}$$

**Example: 8** 1 c.c of  $N_2O$  at NTP contains :

[CBSE PMT 1988]

- (a)  $\frac{1.8}{224} \times 10^{22}$  atoms      (b)  $\frac{6.02}{22400} \times 10^{23}$  molecules  
 (c)  $\frac{1.32}{224} \times 10^{23}$  electrons      (d) All the above

**Solution:** (d) 22400 c.c. =  $6.02 \times 10^{23}$  molecules

$$\begin{aligned} 1 \text{ c.c. of } N_2O &= \frac{6.02 \times 10^{23}}{22400} \text{ molecules} \\ &= \frac{3 \times 6.02 \times 10^{23}}{22400} \text{ atoms (Since } N_2O \text{ has three atoms)} \\ &= \frac{6.02 \times 10^{23}}{22400} \times 22 \text{ electron (Because number of electrons in } N_2O \text{ are 22)} \end{aligned}$$

**Example: 9** The mass of carbon present in 0.5 mole of  $K_4[Fe(CN)_6]$  is

- (a) 1.8 g      (b) 18 g      (c) 3.6 g      (d) 36 g

**Solution:** (d) 1 mole of  $K_4[Fe(CN)_6]$  = 6 gm atoms of carbon

$$\begin{aligned} 0.5 \text{ mole of } K_4[Fe(CN)_6] &= 3 \text{ gm atoms of carbon} \\ &= 3 \times 12 = 36 \text{ g} \end{aligned}$$

**Example: 10** The number of moles of oxygen in one litre of air containing 21% oxygen by volume under standard conditions is

[CBSE PMT 1995]

- (a) 0.186 mole      (b) 0.21 mole      (c) 0.0093 mole      (d) 2.10 mole

**Solution:** (c)  $\therefore$  100 ml of air at STP contains 21 ml of  $O_2$ .

$$\therefore 1000 \text{ ml of air at STP contains } 210 \text{ ml of } O_2.$$



## SOME BASIC CONCEPTS OF CHEMISTRY

$$\therefore \text{No. of moles of } O_2 = \frac{\text{Vol. of } O_2 \text{ in litres under STP conditions}}{22.4 \text{ litre}} = \frac{210/1000}{22.4} = \frac{21}{2240} = 0.0093 \text{ mole}$$

**Example: 11** The number of moles of  $BaCO_3$  which contains 1.5 moles of oxygen atoms is [EAMCET 1991]

- (a) 0.5                      (b) 1                      (c) 3                      (d)  $6.02 \times 10^{23}$

**Solution:** (a)  $\therefore$  1 mole of  $BaCO_3$  contains 3 moles of oxygen atoms.

$\therefore \frac{1}{2}$  mole (0.5) of  $BaCO_3$  contains 1.5 moles of oxygen atoms.

### Percentage composition & Molecular formula.

(1) **Percentage composition of a compound** : Percentage composition of the compound is the relative mass of each of the constituent element in 100 parts of it. If the molecular mass of a compound is  $M$  and  $B$  is the mass of an element in the molecule, then

$$\text{Percentage of element} = \frac{\text{Mass of element}}{\text{Molecular mass}} \times 100 = \frac{X}{M} \times 100$$

(2) **Empirical formula** : The chemical formula that gives the simplest whole number ratio between the atoms of various elements present in one molecule of the compound is known as empirical formula (simplest formula).

For example, the empirical formula of glucose is  $CH_2O$  which shows that  $C$ ,  $H$  and  $O$  are present in the ratio 1 : 2 : 1 in a molecule of glucose.

*Empirical formula mass of a compound is equal to the sum of atomic masses of all atoms present in the empirical formula of the compound.*

**Calculation of the empirical formula** : The empirical formula of a chemical compound can be deduced by knowledge of the, (i) Percentage composition of different elements. (ii) Atomic masses of the elements.

The following steps are involved in the calculation of the empirical formula,

**Step I.** Calculate the relative number of atoms or atomic ratio.

$$\text{Atomic ratio} = \frac{\text{Percentage of an element}}{\text{Atomic mass of the same element}}$$

**Step II.** Calculate the simplest atomic ratio.

**Step III.** Calculate the simple whole number ratio.

**Step IV.** Write the empirical formula.

(3) **Molecular formula** : The chemical formula that gives the actual number of atoms of various elements present in one molecule of the compound. For example, the molecular formula of glucose is  $C_6H_{12}O_6$ .

**Relation between empirical and molecular formula** : The molecular formula of a compound is a simple whole number multiple of its empirical formula.

Molecular formula =  $n \times$  Empirical formula;    Where  $n$  is any integer such as 1, 2, 3...etc.

The value of 'n' can be obtained from the following relation:  $n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$ .

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The molecular mass of a volatile compound can be determined by **Victor Meyer's method** or by employing the relation,  $\boxed{\text{Molecular mass} = 2 \times \text{Vapour density}}$ .

**Calculation of the molecular formula** : The molecular formula of a compound can be deduced from its :

(i) Empirical formula, (ii) Molecular mass

The determination of molecular formula involves the following steps:

*Setp I.* Calculation of empirical formula from the percentage composition.

*Setp II.* Calculation of empirical formula mass.

*Setp III.* Calculation of the value of 'n'.

*Setp IV.* Calculation of molecular formula by multiplying the empirical formula of the compound by 'n'.

**Example: 12** The oxide of a metal contains 40% by mass of oxygen. The percentage of chlorine in the chloride of the metal is [BIT Ranchi 1997]

- (a) 84.7                                      (b) 74.7                                      (c) 64.7                                      (d) 44.7

**Solution:** (b) % of oxygen =  $\frac{8}{m+8} \times 100 = 40$       or,  $40(m+8)=800$       or,  $m+8=20$       or,  $m=12$

$$\therefore \% \text{ of chlorine} = \frac{35.5}{m+35.5} \times 100 = \frac{35.5}{12+35.5} \times 100 = 74.7 \quad (\text{Where } m \text{ is the atomic mass of metal})$$

**Example: 13** The empirical formula of an organic compound containing carbon and hydrogen is  $CH_2$ . The mass of one litre of this organic gas is exactly equal to that of one litre of  $N_2$ . Therefore, the molecular formula of the organic gas is [EAMCET 1985]

- (a)  $C_2H_4$                                       (b)  $C_3H_6$                                       (c)  $C_6H_{12}$                                       (d)  $C_4H_8$

**Solution:** (a) Molar mass of 1L of gas = mass of 1L  $N_2$

$\therefore$  Molecular masses will be equal i.e., molecular mass of the gas = 28, hence formula is  $C_2H_4$

**Example: 14** A sample of pure compound is found to have  $Na = 0.0887$  mole,  $O = 0.132$  mole,  $C = 2.65 \times 10^{22}$  atoms. The empirical formula of the compound is [CPMT 1997]

- (a)  $Na_2CO_3$                                       (b)  $Na_3O_2C_5$   
(c)  $Na_{0.0887}O_{0.132}C_{2.65 \times 10^{22}}$                                       (d)  $NaCO$

**Solution:** (a)  $\therefore 6.02 \times 10^{23}$  atoms of C = 1 mole of C

$$\therefore 2.65 \times 10^{22} \text{ atoms of C} = \frac{1 \times 2.65 \times 10^{22}}{6.02 \times 10^{23}} \text{ mole} = \frac{2.65}{6.02 \times 10} = 0.044 \text{ mole}$$

Now,

Element	Relative number of moles	Simplest ratio of moles
Na	0.0887	$\frac{0.0887}{0.044} = 2$
O	0.132	$\frac{0.132}{0.044} = 3$
C	0.044	$\frac{0.044}{0.044} = 1$

## SOME BASIC CONCEPTS OF CHEMISTRY

Thus, the empirical formula of the compound is  $Na_2CO_3$ .

**Example: 15** An organic compound containing C, H and N gave the following on analysis: C = 40%, H = 13.3% and N = 46.67%. Its empirical formula would be [CBSE PMT 1999, 2002]

- (a) CHN                                      (b)  $C_2H_2N$                                       (c)  $CH_4N$                                       (d)  $C_2H_7N$

**Solution:** (c) Calculation of empirical formula

Element	Symbol	Percentage of element	At. mass of elements	Relative number of atoms = $\frac{\text{Percentage}}{\text{At. mass}}$	Simplest atomic ratio	Simplest whole number atomic ratio
Carbon	C	40	12	$\frac{40}{12} = 3.33$	$\frac{3.33}{3.33} = 1$	1
Hydrogen	H	13.3	1	$\frac{13.3}{1} = 13.3$	$\frac{13.3}{3.33} = 4$	4
Nitrogen	N	46.67	14	$\frac{46.67}{14} = 3.33$	$\frac{3.33}{3.3} = 1$	1

Thus, the empirical formula is  $CH_4N$ .

**Example: 16** An organic substance containing C, H and O gave the following percentage composition : C = 40.687%, H = 5.085% and O = 54.228%. The vapour density of the compound is 59. The molecular formula of the compound is

- (a)  $C_4H_6O_4$                                       (b)  $C_4H_6O_2$                                       (c)  $C_4H_4O_2$                                       (d) None of these

**Solution:** (a)

Element	Symbol	Percentage of element	Atomic mass of element	Relative number of atoms = $\frac{\text{Percentage}}{\text{At. mass}}$	Simplest atomic ratio	Simplest whole number atomic ratio
Carbon	C	40.687	12	$\frac{40.687}{12} = 3.390$	$\frac{3.390}{3.389} = 1$	2
Hydrogen	H	5.085	1	$\frac{5.085}{1} = 5.085$	$\frac{5.085}{3.389} = 1.5$	3
Oxygen	O	54.228	16	$\frac{54.228}{16} = 3.389$	$\frac{3.389}{3.389} = 1$	2

$\therefore$  Empirical formula is  $C_2H_3O_2$

$\therefore$  Empirical formula mass of  $C_2H_3O_2 = 59$

Also, Molecular mass =  $2 \times$  Vapour density =  $2 \times 59 = 118$

$\therefore n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{118}{59} = 2$

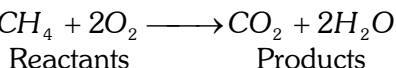
Now, Molecular formula =  $n \times$  (Empirical formula) =  $2 \times (C_2H_3O_2) = C_4H_6O_4$

## SOME BASIC CONCEPTS OF CHEMISTRY

∴ Molecular formula is  $C_4H_6O_4$ .

### Chemical equations and its balancing.

(1) **Chemical equations** : Chemical reactions are represented in a concise way by chemical equations. A chemical equation represents an actual chemical change taking place in terms of the symbols and the formulae of the reactants and products. e.g., Methane burns in oxygen to produce carbon dioxide and water. The chemical reaction can be represented as:



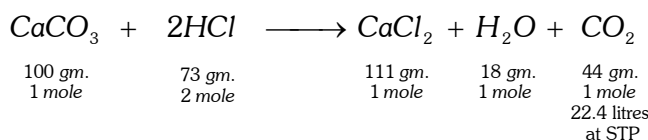
#### Essentials of a chemical equation

- It must represent an actual chemical reaction.
- It must be balanced i.e., the total number of atoms of various substances involved on both sides of the equation must be equal.
- It should be molecular. The elementary gases like hydrogen, oxygen etc. must be written in the molecular form as  $H_2$ ,  $O_2$  etc.

**Information conveyed by a chemical equation** : A Chemical equation conveys both qualitative and quantitative information.

- Qualitative information** : Qualitatively a chemical equation conveys the names of the reactants and products taking part in the reaction.
- Quantitative information** : Quantitatively a chemical equation conveys the following information :
  - It conveys the actual number of reactants and product species (atoms or molecules) taking part in the reaction.
  - It tells the relative masses of the reactants and products participating in the reaction.
  - It conveys the relative number of reactant and product moles.
  - It also conveys the volumes of the gaseous reactants and products if present.

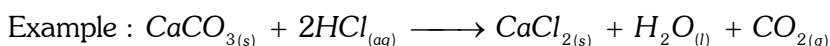
Example : Reaction between  $CaCO_3$  and aqueous  $HCl$ .



**Note** : \* All chemical equations are written under STP conditions provided no other conditions are mentioned.

**Limitations of a chemical equation and their removal** : Although a chemical equation conveys a number of informations, it suffers from certain limitations or drawbacks. The major limitations and the steps taken for their removal are given below:

- Physical states of the reactants and products** : The chemical equation fails to convey the physical states of the reactants and products. These are specified by the use of letters 's'(for solids), 'l'(for liquids), 'g'(for gases) and 'aq'(for aqueous solutions).





## SOME BASIC CONCEPTS OF CHEMISTRY

(iii) Oxidation-Number method, (iv) Ion-Electron method

The first two methods are discussed here, while the remaining two methods will be taken up for discussion in redox reactions.

**(i) Hit and Trial method :** This method involves the following steps:

(a) Write the symbols and formulae of the reactants and products in the form of skeleton equation.

(b) If an elementary gas like  $H_2$ ,  $O_2$  or  $N_2$  etc. appears on either side of the equation, write the same in the atomic form.

(c) Select the formula containing maximum number of atoms and start the process of balancing.

(d) In case the above method is not convenient, then start balancing the atoms which appear minimum number of times.

(e) Balance the atoms of elementary gases in the last.

(f) When the balancing is complete, convert the equation into molecular form.

Let us balance the skeleton equation  $Mg_3N_2 + H_2O \longrightarrow Mg(OH)_2 + NH_3$

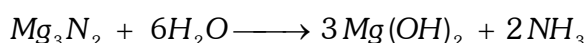
The balancing is done in the following steps

Step I. Balance the Mg atoms  $Mg_3N_2 + H_2O \longrightarrow 3Mg(OH)_2 + NH_3$

Step II. Balance the N atoms  $Mg_3N_2 + H_2O \longrightarrow 3Mg(OH)_2 + 2NH_3$

Step III. Balance the O atoms  $Mg_3N_2 + 6H_2O \longrightarrow 3Mg(OH)_2 + 2NH_3$

The hydrogen atoms are already balanced. Hence, final balanced equation is



**(ii) Partial equation method :** Chemical equations which involve a large number of reactants and products can not be balanced easily by the hit and trial method. In partial equation method, the overall reaction is assumed to take place into two or more simpler reactions known as *partial equations*. The balancing of the equation involves the following steps:

(a) Split the chemical equation into two or more simpler equations or partial equations.

(b) Balance each partial equation separately by hit and trial method.

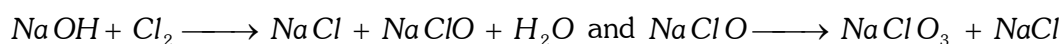
(c) Multiply the partial equations with suitable coefficient if necessary so as to cancel out the final substances which do not appear in the final equation.

(d) Finally, add up the partial equations to get the final equation.

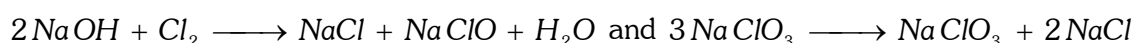
Let us balance the skeleton equation  $NaOH + Cl_2 \longrightarrow NaCl + NaClO_3 + H_2O$

This reaction is supposed to take place in the following steps:

The probable partial equations for the above reaction are:

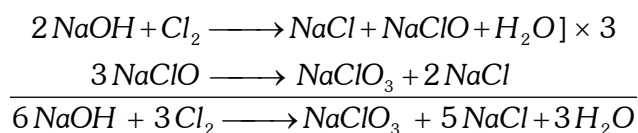


Balance the partial chemical equations separately by hit and trial method as



Multiply the first partial equation by 3 in order to cancel out  $NaClO$  which does not appear in the final equation. Finally add the two partial equations to get the final equation.

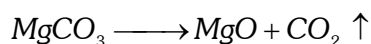
## SOME BASIC CONCEPTS OF CHEMISTRY



### Chemical Stoichiometry.

Calculation based on chemical equations is known as chemical stoichiometry. Stoichiometry can be broadly classified into two groups: (1) Gravimetric analysis (Stoichiometry-I), (2) Volumetric analysis (Stoichiometry-II)

**(1) Gravimetric analysis (Stoichiometry-I) :** With the help of chemical equation, we can calculate the weights of various substances reacting and weight of substances formed. For example,



This equation implies :

- (i) 1 mol of  $\text{MgCO}_3$  gives 1 mol of  $\text{MgO}$  and 1 mol of  $\text{CO}_2$ .
- (ii) 84 g of  $\text{MgCO}_3$  (Mol. wt. of  $\text{MgCO}_3$ ) gives 40 g of  $\text{MgO}$  and 44 g of  $\text{CO}_2$ .

Hence, chemical equation provide us information regarding :

- (i) Molar ratio of reactants and products.
- (ii) Mass ratio between reactants and products.
- (iii) Volume ratio between gaseous reactant and products.

The calculation based upon chemical equation (Stoichiometry-I) are based upon three types namely :

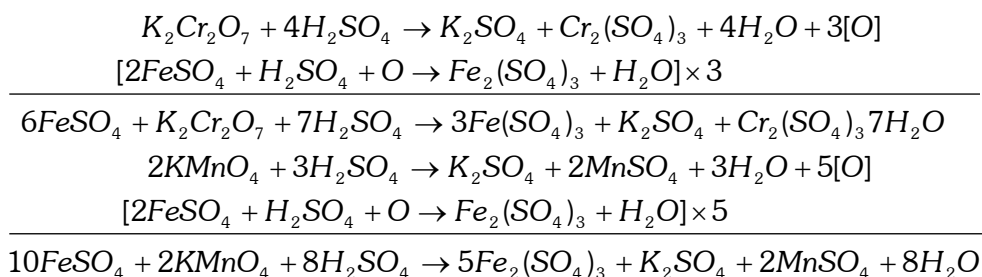
- (a) Mass-mass relationship (b) Mass-volume relationship (c) Volume-volume relationship

**(2) Volumetric analysis (Stoichiometry-II) :** It is a method which involves quantitative determination of the amount of any substance present in a solution through volume measurements. For the analysis a standard solution is required. (A solution which contains a known weight of the solute present in known volume of the solution is known as standard solution.)

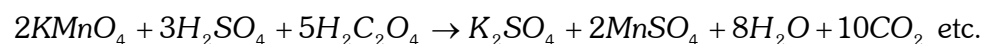
To determine the strength of unknown solution with the help of known (standard) solution is known as titration. Different types of titrations are possible which are summerised as follows :

(i) *Redox titrations* : To determine the strength of oxidising agents or reducing agents by titration with the help of standard solution of reducing agents or oxidising agents.

Examples:



Similarly with  $\text{H}_2\text{C}_2\text{O}_4$



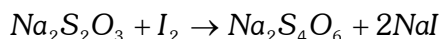
## SOME BASIC CONCEPTS OF CHEMISTRY

(ii) *Acid-base titrations* : To determine the strength of acid or base with the help of standard solution of base or acid.

Example:  $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$  and  $\text{NaOH} + \text{CH}_3\text{COOH} \rightarrow \text{CH}_3\text{COONa} + \text{H}_2\text{O}$  etc.

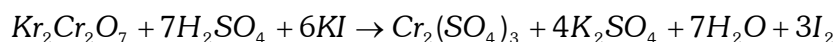
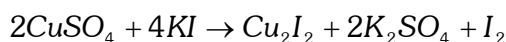
(iii) *Iodometric titrations* : To determine the reducing agents with the help of standard iodine solution is known as iodometry.

For example:  $\text{As}_2\text{O}_3 + 2\text{I}_2 + 2\text{H}_2\text{O} \rightarrow \text{As}_2\text{O}_5 + 4\text{HI}$   
Reducing agent

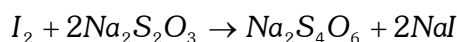


(iv) *Iodometric titrations* : To determine the oxidising agent indirectly by titration of liberated  $\text{I}_2$  with the help of standard hypo solution is known as iodometric titrations.

Examples: Oxidising agents such as  $\text{KMnO}_4$ ,  $\text{K}_2\text{Cr}_2\text{O}_7$ ,  $\text{CuSO}_4$ , ferric salts, etc. are reduced quantitatively when treated with large excess of  $\text{KI}$  in acidic or neutral medium and liberate equivalent amount of  $\text{I}_2$ .



This  $\text{I}_2$  is estimated with hypo



(v) *Precipitation titrations* : To determine the anions like  $\text{CN}^-$ ,  $\text{AsO}_3^{3-}$ ,  $\text{PO}_4^{3-}$ ,  $\text{X}^-$  etc, by precipitating with  $\text{AgNO}_3$  provides examples of precipitation titrations.



**End point and equivalence point** : The point at which titration is stopped is known as end point, while the point at which the acid and base (or oxidising and reducing agents) have been added in equivalent quantities is known as equivalence point. Since the purpose of the indicator is to stop the titration close to the point at which the reacting substances were added in equivalent quantities, it is important that the equivalent point and the end point be as close as possible.

**Normal solution** : A solution containing one gram equivalent weight of the solute dissolved per litre is called a normal solution; e.g. when 40 g of  $\text{NaOH}$  are present in one litre of  $\text{NaOH}$  solution, the solution is known as normal ( $N$ ) solution of  $\text{NaOH}$ . Similarly, a solution containing a fraction of gram equivalent weight of the solute dissolved per litre is known as subnormal solution. For example, a solution of  $\text{NaOH}$  containing 20 g ( $1/2$  of g eq. wt.) of  $\text{NaOH}$  dissolved per litre is a sub-normal solution. It is written as  $N/2$  or  $0.5 N$  solution.

### Formula used in solving numerical problems on volumetric analysis

(1) Strength of solution = Amount of substance in g  $\text{litre}^{-1}$

(2) Strength of solution = Amount of substance in g  $\text{moles litre}^{-1}$

(3) Strength of solution = Normality  $\times$  Eq. wt. of the solute

(4) Molarity =  $\frac{\text{Moles of solute}}{\text{Volume in litre}}$

(5) Number of moles =  $\frac{\text{Wt. in gm}}{\text{Mol. wt.}} = M \times V_{(ml)} = \frac{\text{Volume in litres}}{22.4}$  at NTP (only for gases)

(6) Number of millimoles =  $\frac{\text{Wt. in gm} \times 1000}{\text{mol. wt.}} = \text{Molarity} \times \text{Volume in ml.}$



## SOME BASIC CONCEPTS OF CHEMISTRY

$$(7) \text{ Number of equivalents} = \frac{\text{Wt. in gm}}{\text{Eq. wt.}} = x \times \text{No. of moles} \times \text{Normality} \times \text{Volume in litre}$$

$$(8) \text{ Number of milliequivalents (meq.)} = \frac{\text{Wt. in gm} \times 1000}{\text{Eq. wt.}} = \text{normality} \times \text{Volume in ml.}$$

$$(9) \text{ Normality} = x \times \text{No. of millimoles} = x \times \text{Molarity} = \frac{\text{Strength in gm litre}^{-1}}{\text{Eq. wt.}}$$

where  $x = \frac{\text{Mol. wt.}}{\text{Eq. wt.}}$ ,  $x^*$  = valency or change in oxi. Number.

$$(10) \text{ Normality formula, } N_1 V_1 = N_2 V_2$$

$$(11) \% \text{ by weight} = \frac{\text{Wt. of solvent}}{\text{Wt. of solution}} \times 100$$

$$(12) \% \text{ by volume} = \frac{\text{Wt. of solvent}}{\text{Vol. of solution}} \times 100$$

$$(13) \% \text{ by strength} = \frac{\text{Vol. of solvent}}{\text{Vol. of solution}} \times 100$$

$$(14) \text{ Specific gravity} = \frac{\text{Wt. of solution}}{\text{Vol. of solution}} = \text{Wt. of 1 ml. of solution}$$

$$(15) \text{ Formality} = \frac{\text{Wt. of ionic solute}}{\text{Formula Wt. of solute} \times V_{\text{ml}}}$$

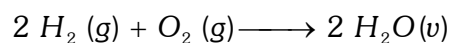
$$(16) \text{ Mol. Wt.} = V.D \times 2 \text{ (For gases only)}$$

### Limiting reagent or reactant.

In many situations, an excess of one or more substance is available for chemical reaction. Some of these excess substances will therefore be left over when the reaction is complete; the reaction stops immediately as soon as one of the reactant is totally consumed.

*The substance that is totally consumed in a reaction is called limiting reagent because it determines or limits the amount of product. The other reactant present in excess are called as excess reagents.*

Let us consider a chemical reaction which is initiated by passing a spark through a reaction vessel containing 10 mole of  $H_2$  and 7 mole of  $O_2$ .



Moles before reaction	10	7	0
Moles after reaction	0	2	10

The reaction stops only after consumption of 5 moles of  $O_2$  as no further amount of  $H_2$  is left to react with unreacted  $O_2$ . Thus  $H_2$  is a *limiting reagent* in this reaction

**Example: 17** When a solution containing 4.77 gm. of  $NaCl$  is added to a solution of 5.77 gm. of  $AgNO_3$ , the weight of precipitated  $AgCl$  is [IIT 1978]  
 (a) 11.70 gm.                      (b) 9.70 gm.                      (c) 4.86 gm.                      (d) 2.86 gm.

**Solution:** (c)

