MOLE CONCEPT

1. MOLE

A mole is the amount of substance that contains as many species [Atoms, molecules, ions or other particles] as there are atoms in exactly 12 gm of C-12.

$$\boxed{1\,mole = 6.022\times 10^{23}\,\text{species}}$$

2.1 Atomic mass

Atomic mass of an element can be defined as the number which indicates how many times the mass of one atom of the element is heavier in comparison to $\frac{1}{12}$ th part of the mass of one atom of Carbon-12.

$$\text{Atomic mass} = \frac{[\text{Mass of an atom of the element}]}{\frac{1}{12}\times[\text{Mass of an atom of carbon - 12}]} = \frac{\text{Mass of an atom in amu}}{1\text{ amu}}$$

2.2 Atomic mass unit (amu) or Unified mass (u)

The quantity $[\frac{1}{12}\times$ mass of an atom of C–12] is known as atomic mass unit.

The actual mass of one atom of C-12 = 1.9924×10^{-26} kg

$$\therefore\ 1\text{ amu} = \frac{1.9924\times 10^{-26}}{12}\text{ kg}$$

$$= 1.66\times 10^{-27}\text{ kg} = 1.66\times 10^{-24}\text{ gm} = \frac{1}{N_A}\text{ gm}$$

2.3 Gram atomic mass

The gram atomic mass can be defined as the mass of 1 mole atoms of an element.

e.g., Mass of one oxygen atom = 16 amu = $\frac{16}{N_A}$ gm.

Mass of N_A oxygen atom = $\frac{16}{N_A}.N_A$ = 16 gram

Illustration

(a) What is the mass of one atom of Cl? (b) What is the atomic mass of Cl?

(c) What is the gram atomic mass of Cl?

Sol. (a) Mass of one atom of Cl = 35.5 amu.

(b) Atomic mass of Cl = $\dfrac{\text{Mass of an atom in amu}}{1\text{amu}} = \dfrac{35.5\text{amu}}{1\text{amu}} = 35.5$

(c) Gram atomic mass of Cl = [Mass of 1 Cl atom × N_A]

$$= 35.5 \text{ amu} \times N_A = \frac{35.5}{N_A} \times N_A \text{ gram} = 35.5 \text{ gram}$$

Exercise

(a) What is the mass of one atom of S?

(b) What is the atomic mass of S ?

(c) What is the gram atomic mass of S?

Ans. (a) 32 amu, (b) 32, (c) 32 gram

3.1 Molecular mass

Molecular mass is the number which indicates how many times one molecule of a substance is heavier in comparison to $\frac{1}{12}$ th of the mass of one atom of C-12.

$$\text{Molecular mass} = \frac{\text{Mass of one molecule of the substance (in amu)}}{\frac{1}{12} \times [\text{Mass of an atom of C-12}]}$$

$$= \frac{\text{Mass of one molecule of the substance (in amu)}}{1\text{ amu}}$$

3.2 Gram Molecular mass

Gram molecular mass can be defined as the mass of 1 mole of molecules.

e.g., Mass of one molecule of $O_2 = 32 \text{ amu} = \frac{32}{N_A} \text{ gram}$.

Mass of N_A molecules of $O_2 = \frac{32}{N_A} \times N_A \text{gm} = 32 \text{ gm}$

Illustration

(a) What is the mass of one molecule of HNO_3?

(b) What is the molecular mass of HNO_3?

(c) What is the gram molecular mass of HNO_3?

Sol. (a) Mass of one molecule of $HNO_3 = (1 + 14 + 3 \times 16)$ amu $= 63$ amu.

(b) Molecular mass of $HNO_3 = \dfrac{63\ amu}{1\ amu} = 63$

(c) Gram molecular mass of HNO_3 = Mass of 1-molecule of $HNO_3 \times N_A$

$$= 63 \text{ amu} \times N_A = \frac{63}{N_A} \text{gm} \times N_A = 63 \text{ gram}$$

Exercise

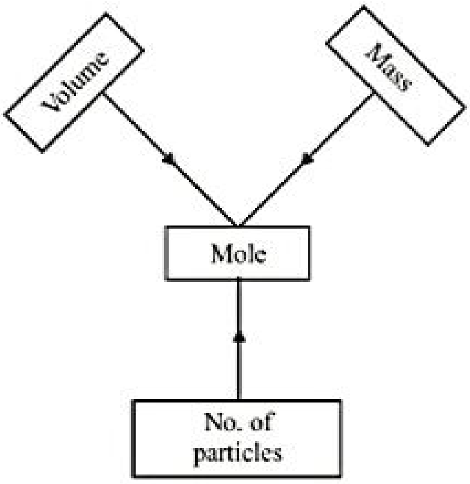
(a) What is the mass of one molecule of H_2SO_4.

(b) What is the molecular mass of H_2SO_4.

(c) What is the gram molecular mass of H_2SO_4.

Ans. (a) 98 amu (b) 98 (c) 98 gram

4. METHODS TO CALCULATE MOLES



4.1 From number of particles :

$$\boxed{\text{No. of mole} = \frac{\text{Given no. of Paritcles [atoms/molecules/ ions]}}{N_A}}$$

Illustration

A piece of Cu contains 6.022×10^{24} atoms. How many mole of Cu atoms does it contain?

Sol. No. of mole $= \dfrac{6.022 \times 10^{24}}{N_A} = \dfrac{6.022 \times 10^{24}}{6.022 \times 10^{23}} = 10$ mole

Exercise

5 mole of CO_2 are present in a gaseous sample. How many molecules of CO_2 are present in the sample?

Ans. $5\ N_A$

4.2 From given Mass :

(a) For atoms : No. of mole $= \dfrac{\text{Given mass of the substance(gm)}}{\text{Gram atomic mass}}$ = No. of g-atoms

(b) For molecules : No. of mole $= \dfrac{\text{Given mass of the substance(gm)}}{\text{Gram molecular mass}}$ = No. of g-molecules

Illustration

What will be the mass of 5 mole of SO_2?

Ans. Molecular mass of SO_2 = 64 gm

$$5 = \frac{\text{mass (gm)}}{64}$$

$\therefore$ mass = 320 gm

Exercise

(a) How many mole of O atoms are present in 88 gm CO_2?

(b) What will be the mass of 10 mole of H_3PO_4?

Ans. (a) 4 mole (b) 980 gm

4.3 From the given volume of a gas :

$$\boxed{n = \frac{\textit{volume of gas at 1bar pressure and 273 K (in litre)}}{22.7}}$$

S.T.P.: 1 bar pressure and 273 K.

$$\boxed{n = \frac{\text{volume of gas at 1atm and 273K(in litre)}}{22.4}}$$

Note: According to old IUPAC agreement, STP condition was 1 atm pressure and 273 K temperature but according to new agreement it is 1 bar pressure and 273K temperature. Although many books are still using the condition of 1 atm and 273K for STP.

If volume is given under any other condition of temperature and pressure, then use the ideal gas equation to find the no. of moles.

$$\boxed{\text{No. of mole(n)} = \frac{PV}{RT}}$$

Units of Pressure :

1 atm = 76 cm Hg = 760 torr = 1.01325 bar = 1.01325×10^5 pa.

Units of temperature :

$$\boxed{K = {}^{\circ}C + 273}$$

Value of R :

R = 0.0821 litre-atm/mole.K

= 8.314 J/mole.K = 1.987 $\approx$ 2 cal/mole.K

Units of volume :

$$1\ dm^3 = 10^3 cm^3 = 1\ litre = 10^{-3} m^3 = 10^3 ml$$

$$1 m^3 = 10^3\ litre$$

Illustration

A sample of He gas occupies 5.6 litre volume at 1 atm and 273 K. How many mole of He are present in the sample?

Sol. No. of mole $= \frac{5.6}{22.4} = 0.25$

Exercise

How much volume will be occupied by 2 mole CO_2 gas at STP?

Ans. 45.4 L

Note : We can use the following relationship as per requirement of question.

$$\text{No. of mole} = \frac{\text{No. of particle}}{N_A} = \frac{\text{mass (gm)}}{[\text{gm at. or mol. mass}]}$$

$$= \frac{V(\ell)\text{ occupied by a Gas at STP}}{22.7} = \frac{V(\ell)\text{ occupied by a Gas at 1atm and 273K}}{22.4}$$

Illustration

How many molecules of O_2 are present in 5.6 litres of O_2 at 1 atm and 273 K?

Sol. $$\frac{\text{No. of molecules}}{N_A} = \frac{V(\ell) \text{ at 1 atm and 273 K}}{22.4}$$

$$\frac{\text{No. of molecules}}{N_A} = \frac{5.6}{22.4} = \frac{1}{4} \Rightarrow \text{No. of molecule} = \frac{N_A}{4} = 1.505 \times 10^{23}$$

Exercise

How many molecules of water are present in 9 gram of water?

Ans. 3.011×10^{23}

5. LAWS OF CHEMICAL COMBINATION

5.1 Law of conservation of mass (Lavoisier – 1774) :

In any physical or chemical change, mass can neither be created nor be destroyed.

It means :

Total mass of the reactants = total mass of the products.

This relationship holds good when reactants are completely converted into products.

In case the reacting material are not completely consumed the relationship will be –

Total mass of the reactants = Total mass of the products + mass of unreacted reactants

Limitation : In nuclear reactions, some mass of reactant is converted into energy, so mass of reactant is always less than that of product.

Illustrations

1.7 gram of silver nitrate dissolved in 100 gram of water is taken. 0.585 gram of sodium chloride dissolved in 100 gram of water is added to it and chemical reaction occurs. 1.435 gm of AgCl and 0.85 gm $NaNO_3$ are formed. Show that these results illustrate the law of conservation of mass.

Sol. Total mass before chemical change = mass of $AgNO_3$ + Mass of NaCl + Mass of water

= 1.70 + 0.585 + 200 = 202.285 gram

Total mass after the chemical reaction = mass of AgCl + Mass of $NaNO_3$ + Mass of water

= 1.435 + 0.85 + 200 = 202.285 gram

Thus in the given reaction

Total mass of reactants = Total mass of the products.

Exercise

If 6.3 gram of $NaHCO_3$ are added to 15 gram CH_3COOH solution. The residue is found to weigh 18 gram. What is the mass of CO_2 released in this reaction?

Ans. 3.3 gram

5.2 Law of constant composition : [Proust 1799]

A chemical compound always contains the same element combined together in fixed proportion by mass.

Limitations : In case of isotopes, ratio is not fixed

e.g.	$^{12}CO_2$	$^{14}CO_2$
	12: 32	14 : 32
	3 : 8	7 : 16

Illustration

1.08 gram of Cu wire was allowed to react with nitric acid. The resulting solution was dried and ignited when 1.35 gram of copper oxide was obtained. In another experiment 2.3 gram of copper oxide was heated in presence of Hydrogen yielding 1.84 gram of copper. Show that the above data are in accordance with law of constant composition?

Sol. **Case-I**

$$Cu \xrightarrow{HNO_3} Cu(NO_3)_2 \xrightarrow{\Delta} CuO$$

1.08 gram 1.35 gram

1.35 gram CuO contains $\longrightarrow$ 1.08 gram Cu.

$$100 \text{ gram CuO contains} \longrightarrow \frac{1.08}{1.35} \times 100 = 80 \text{ gram Cu}$$

% Cu in CuO = 80%

% O in CuO = 20%

Case-II

$$CuO + H_2 \longrightarrow Cu + H_2O$$

2.3 gram 1.84 gram

2.30 gram CuO contains $\longrightarrow$ 1.84 gram Cu.

$$100 \text{ gram CuO contains} \longrightarrow \frac{1.84 \times 100}{2.30} = 80 \text{ gram Cu}$$

% Cu in CuO = 80%

% O in CuO = 20%

Both sample have the same composition & hence the data are in accordance with law of constant composition.

Exercise

7.95 gram of cupric oxide was reduced by heating in a current of hydrogen and the weight of copper that remained was 6.35 gram. In another experiment, 19.05 gram of Cu was dissolved in the nitric acid and the resulting copper nitrate is converted into cupric oxide by ignition. The weight of cupric oxide formed was 23.85 gram. Show that these results illustrate the law of constant composition.

5.3 Law of multiple proportion : [Dalton 1806]

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

Illustration

Two compounds each containing only tin and oxygen had the following composition.

	Mass % of Tin	Mass % of oxygen
Compound A	78.77	21.23
Compound B	88.12	11.88

Show that these data illustrate the law of multiple proportion?

Sol.

In compound A

21.23 parts of oxygen combine with 78.77 parts of tin.

1 part of oxygen combine with $\frac{78.77}{21.23} = 3.7$ parts of Sn.

In compound B

11.88 parts of oxygen combine with 88.12 parts of tin.

1 part of oxygen combined with $\frac{88.12}{11.88} = 7.4$ parts of tin.

Thus the mass of Tin in compound A and B which combine with a fixed mass of oxygen are in the ratio 3.7 : 7.4 or 1 : 2. This is a simple ratio. Hence the data illustrate the law of multiple proportion.

Exercise

Carbon and oxygen are known to form two compounds. The carbon content in one of these is 42.9% while in the other it is 27.3%. Show that these data are in agreement with the law of multiple proportion.

5.4 Law of reciprocal proportion : [Richter 1794]

When two different elements combine with the same mass of a third element, the ratio in which they do so will be same or simple multiple if both directly combined with each other.

Illustration

The % composition of NH_3, H_2O and N_2O_3 is as given below :

$NH_3 \longrightarrow$ 82.35% N and 17.65% H.

$H_2O \longrightarrow$ 88.9% O and 11.1% H

$N_2O_3 \longrightarrow$ 63.15% O and 36.85% N

On the basis of above data prove the law of reciprocal proportion?

Sol. (i) For NH_3 1-part of hydrogen reacts with $= \dfrac{82.35}{17.65} = 4.67$ part N.

(ii) For H_2O 1-part of hydrogen reacts with $= \dfrac{88.90}{11.10} = 8.01$ part O.

Thus the ratio N : O = 4.67 : 8.01 = 1: 1.71

(iii) For N_2O_3 : N and O reacts with each other N : O = 36.85 : 63.15 = 1 : 1.71

Because the two ratios are same, thus law of reciprocal proportion is proved.

5.5 Gay-Lussac's law of gaseous volumes [Gay-Lussac–1808] :

When gases combined or produced in a chemical reaction, they do so in a simple ratio by volume provided all the gases are at same temperature and pressure.

6. SIGNIFICANCE OF CHEMICAL EQUATIONS

A chemical equation describes the chemical process both qualitatively and quantitatively. The stoichiometric coefficients in the chemical equation give the quantitative information of the chemical process. These coefficients represent the relative number of molecules or moles of the reactants and products, e.g.,

	$2\,NH_3\,(g)$	$\rightarrow$	$N_2\,(g)$	$+$	$3\,H_2\,(g)$
	2 molecules		1 molecule		3 molecules
or	2 N molecules		N molecules		3 N molecules
or	2 moles		1 mole		3 moles

Again, Avogadro's principle states that under the same conditions of temperature and pressure, equal volumes of gases contain the same number of molecules. Thus, for homogeneous gaseous reactions, the stoichiometric coefficients of the chemical equation also signify the relative volumes of each reactant and product under the same conditions of temperature and pressure, e.g.,

	$H_2(g)$	$+$	$I_2\,(g)$	$\rightarrow$	$2\,HI\,(g)$
	1 molecule		1 molecule		2 molecule
or	1 mole		1 mole		2 mole
or	1 volume		1 volume		2 volume (T & P constant)
or	1 pressure		1 pressure		2 pressure (T & V constant)

6.1 LIMITING REAGENT

The reactant which gives least amount of product on being completely consumed is known as limiting reagent. It may also be defined as the reactant that is completely consumed when a reaction goes to completion. It comes into the picture when reaction involves two ore more reactants. For solving such reactions, first step is to calculate Limiting Reagent.

Calculation of Limiting Reagent:

Method-I : By calculating the required amount by the equation and comparing it with given amount. [Useful when only two reactant are there]

Method-II : By calculating amount of any one product obtained taking each reactant one by one irrespective of other reactants. The one giving least product is *limiting reagent*.

Method-III : Divide given moles of each reactant by their stoichiometric coefficient, the one with least ratio is *limiting reagent*. [Useful when number of reactants are more than two.]

Illustration

If 20gm of $CaCO_3$ is treated with 20gm of HCl, how many grams of CO_2 can be generated according to following reaction?

$CaCO_3(g) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(\ell) + CO_2(g)$

Sol. Mole of $CaCO_3 = \frac{20}{100} = 0.2$

Mole of HCl $= \frac{20}{36.5} = 0.548$

$$\left[\frac{Mole}{Stoichiometric\ co\text{-}efficient}\right] \text{ for } CaCO_3 = \frac{0.2}{1} = 0.2$$

$$\left[\frac{Mole}{Stoichiometric\ co\text{-}efficient}\right] \text{ for HCl} = \frac{0.548}{2} = 0.274$$

So $CaCO_3$ is limiting reagent

According to reaction :

100 gm of $CaCO_3$ gives 44gm of CO_2

20 gm $CaCO_3$ will give $\frac{44}{100} \times 20 = 8.8$ gm CO_2

Exercise

Calculate the mass of carbon tetrachloride can be produced by the reaction of 144gm of carbon with 71 gm of Chlorine.

Ans. 77 gm

6.2 PROBLEMS RELATED WITH MIXTURE

Illustration

4 gram of a mixture of $CaCO_3$ and Sand (SiO_2) is treated with an excess of HCl and 0.88 gm of CO_2 is produced. What is the percentage of $CaCO_3$ in the original mixture?

Ans. $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

$SiO_2 + HCl \rightarrow$ No reaction

$CaCO_3 = x$ gm

100 gm $CaCO_3$ gives $\longrightarrow$ 44 gm CO_2

x gm $CaCO_3$ gives $\longrightarrow$ 0.88 gm CO_2

$$\Rightarrow \quad \frac{100}{x} = \frac{44}{0.88} \quad \Rightarrow \quad x = 2 \text{ gram}$$

$$\% \ CaCO_3 = \frac{2}{4} \times 100 = 50\%$$

Exercise

44 gram sample of a natural gas, consisting of methane [CH_4] and ethylene [C_2H_4] was burned in excess of oxygen yielding 132 gm CO_2 and some H_2O as products. What is the mole % of ethylene in the sample?

Ans. 50%

6.3 PERCENTAGE YIELD

In gener59al, when a reaction is carried out in the laboratory we do not obtain the theoretical amount of product. The amount of product that is actually obtained is called the actual yield. Knowing the actual yield and theoretical yield, the % yield can be calculated by the following formula–

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoritical yield}} \times 100\,\%$$

Illustration

For the reaction

$$CaO + 2HCl \rightarrow CaCl_2 + H_2O$$

1.12 gram of CaO is reacted with excess of hydrochloric acid and 1.85 gm $CaCl_2$ is formed. What is the % yield of the reaction?

Sol. $CaO + 2HCl \rightarrow CaCl_2 + H_2O$

56 gm CaO will produce 111 gm $CaCl_2$

1.12 gram of CaO will produce $\rightarrow \frac{111}{56} \times 1.12 = 2.22$ gm

Thus Theoretical yield = 2.22 gm

Actual yield = 1.85 gm

$$\% \text{ yield} = \frac{1.85}{2.22} \times 100 = 83.33\ \%$$

Exercise

Calculate the mass of KCl that is produced from 1 mole of $KClO_3$ if % yield of reaction is 80?

Ans. 59.6 gm

6.4 PERCENTAGE PURITY

Depending upon the mass of the product, the equivalent amount of reactant present can be determined with the help of given chemical equation. Knowing the actual amount of the reactant taken and the amount calculated with the help of a chemical equation, the purity can be determined, as

$$\text{Percentage purity} = \left[\frac{\text{Amount of reactant calculated from the chemical equation}}{\text{Actual amount of reactant taken}}\right] \times 100\%$$

Illustration

Calculate the amount of (CaO) in kg that can be produced by heating 200 kg lime stone that is 90% pure $CaCO_3$.

Sol. Mass of Pure $CaCO_3 = \frac{200 \times 90}{100} = 180$ kg

$$\begin{array}{llll} CaCO_3 & \longrightarrow & CaO & + \ CO_2 \\ 100 \text{ kg} & & 56 \text{ kg} & \\ 180 & & x & \end{array}$$

$$\frac{100}{180} = \frac{56}{x} \Rightarrow x = 100.8 \text{ kg}$$

Exercise

Calculate the mass of coal sample in kg containing 0.05% mass of iron pyrite [FeS_2] that can produced 44.8 litre of SO_2 at 1atm and 273 with 40% reaction yield?

$$FeS_2 + \frac{5}{2}O_2 \longrightarrow FeO + 2SO_2$$

Ans. 600 kg

6.5 PROBLEMS RELATED WITH SEQUENTIAL REACTION

Illustration

$KClO_4$ can be prepared by Cl_2 and KOH by a series of reactions as given below

$Cl_2 + 2KOH \longrightarrow KCl + KClO + H_2O$

$3KClO \longrightarrow 2KCl + KClO_3$

$4KClO_3 \longrightarrow 3KClO_4 + KCl$

(i) Calculate mass of Cl_2 in gram required to produce 1385 gm $KClO_4$?

(ii) Calculate the total mass of KCl produced in I[st], 2[nd] and 3[rd] reaction?

Sol. (i) Let mole of Cl_2 required = x

$$\underset{x}{Cl_2} + 2KOH \longrightarrow \underset{x}{KCl} + \underset{x}{KClO} + H_2O$$

$$\underset{x}{3KClO} \longrightarrow \underset{\frac{2x}{3}}{2KCl} + \underset{\frac{x}{3}}{KClO_3}$$

$$\underset{\frac{x}{3}}{4KClO_3} \longrightarrow \underset{\frac{x}{4}}{3KClO_4} + \underset{\frac{x}{12}}{KCl}$$

Mole of $KClO_4$ formed $= \dfrac{1385}{138.5} = 10$

$$\frac{x}{4} = 10, \qquad x = 40$$

$\therefore$ Mass of Cl_2 required for the reaction $= 40 \times 71 = 2840$ gm

(ii) KCl produced from I[st], II[nd] and III[rd] reaction $= \left(x + \dfrac{2x}{3} + \dfrac{x}{12}\right)$ mole

$= \dfrac{21x}{12}$ mole $= \dfrac{21}{12} \times 40 = 70$ mole

$\therefore$ Mass of KCl produced $= 70 \times 74.5 = 5215$ gram

Exercise

Sulphur trioxide may be prepared by the following two reactions

$$S_8 + 8O_2 \longrightarrow 8SO_2$$
$$2SO_2 + O_2 \longrightarrow 2SO_3$$

How many gram of SO_3 will be produced from 1 mole of S_8.

Ans. 640 gram

6.6 PROBLEM RELATED WITH PARALLEL REACTIONS

Illustration

If 6 moles of $KClO_3$ are decomposed according to following reactions calculate the moles of $KClO_4$ produced if mole of O_2 produced are 3?

$$KClO_3 \longrightarrow KCl + \frac{3}{2}O_2 \qquad \text{... (i)}$$

$$4KClO_3 \longrightarrow 3KClO_4 + KCl \qquad \text{... (ii)}$$

Sol. Let x-mole $KClO_3$ reacts in reaction (i) and y mole $KClO_3$ reacts in reaction (ii)

$$KClO_3 \longrightarrow KCl + \frac{3}{2}O_2 \qquad \text{... (i)}$$

x mole $\qquad \frac{3x}{2}$ mole

$$4KClO_3 \longrightarrow 3KClO_4 + KCl \qquad \text{... (ii)}$$

y mole

From question $\qquad x + y = 6$

and $\frac{3x}{2} = 3$

$\therefore \qquad x = 2$ mole $\qquad$ and $\qquad y = 4$ mole

It means 4 mole $KClO_3$ reacts in reaction (ii)

From reaction (ii)

4 mole $KClO_3$ gives 3 mole $KClO_4$

Ans. 3 mole

Exercise

$$A_2 + 2B_2 \longrightarrow A_2B_4 \quad \text{and} \quad \frac{3}{2}A_2 + 2B_2 \longrightarrow A_3B_4$$

Two substances A_2 and B_2 are allowed to react completely to form A_2B_4 and A_3B_4 mixture, leaving none of the reactions. Using this information. Calculate the composition of final mixture when 5/4 mole of A_2 and 2 mole of B_2 is taken?

Ans. $A_3B_4 = 0.5$ mole, $A_2B_4 = 0.5$ mole

6.7 PRINCIPLE OF ATOM CONSERVATION

The principle of conservation of mass, expressed in the concepts of atomic theory means the conservation of atoms. And if atoms are conserved, moles of atoms shall also be conserved. This is known as the principle of atom conservation. This principle is in fact the basis of the mole concept.

In order to solve problems of nearly all stoichiometric calculations, let us first see how this principle works. Choose an example,

$$KClO_3\,(s) \rightarrow KCl\,(s) + O_2\,(g)$$

Apply the principle of atom conservation (POAC) for K atoms.

Moles of K atoms in reactant = mole of K atoms in products

Moles of K atoms in $KClO_3$ = moles of K atoms in KCl.

Now, since 1 molecule of $KClO_3$ contains 1 atom of K

or 1 mole of $KClO_3$ contains 1 mole of K, similarly, 1 mole of KCl contains 1 mole of K

Thus, mole of K atoms in $KClO_3 = 1 \times$ moles of $KClO_3$

and Mole of K atoms in KCl = $1 \times$ moles of KCl

$\therefore$ mole of $KClO_3$ = mole of KCl

or $$\frac{\text{wt. of } KClO_3 \text{ in g}}{\text{mol. wt. of } KClO_3} = \frac{\text{wt. of } KCl \text{ in g}}{\text{mol. wt. of } KCl}$$

The above equation gives the weight relationship between $KClO_3$ and KCl which is important in stoichiometric calculations.

Again, applying the principle of atom conservation for O atoms,

Moles of O in $KClO_3$ = moles of O in O_2

But since 1 mole of $KClO_3$ contains 3 moles of O and 1 mole of O_2 contains 2 moles of O,

Thus, mole of O in $KClO_3 = 3 \times$ moles of $KClO_3$

Moles of O in $O_2 = 2 \times$ moles of O_2

$\therefore$ $3 \times$ moles of $KClO_3 = 2 \times$ moles of O_2

or $$3 \times \frac{\text{wt. of } KClO_3}{\text{mol. wt. of } KClO_3} = 2 \times \frac{\text{wt. of } O_2}{\text{mol. wt. of } O_2}$$

Mole of O_2 may also be expressed in volume.

Illustration

All carbon atoms present in $KH_3(C_2O_4)_2.2H_2O$ weighing 254 gm is converted to CO_2. How many gram of CO_2 were obtained?

Sol. Apply P.O.A.C. on carbon atom

$4 \times$ mole of $KH_3(C_2O_4)_2.2H_2O = 1 \times$ mole of CO_2

$$4 \times \frac{254}{254} = 1 \times \frac{w_{CO_2}}{44}$$

$\therefore$ Mass of $CO_2 = 4 \times 44 = 176$ gram

Exercise

A sample of KNO_3 weighing W_1 gram undergo a series of reaction in such a way that all nitrogen atom are converted to $K_2Zn_3[Fe(CN)_6]_2$. How many gram of the product were obtained ?
[Given M.wt of $KNO_3 = M_1$, M.wt of $K_2Zn_3[Fe(CN)_6)]_2 = M_2$]

Ans. $\dfrac{W_1M_2}{12M_1}$

7.1 AVERAGE ATOMIC MASS

$$\text{Average atomic mass} = \frac{\text{total mass}}{\text{total mole of atoms}}$$

Let a sample contains n_1 mole of atoms with atomic mass M_1 and n_2 mole of atoms with atomic mass M_2. then

$$\boxed{M_{av} = \frac{n_1M_1 + n_2M_2}{n_1 + n_2}}$$

Illustration

Find the average atomic mass of a mixture containing 25% by mole Cl^{37} and 75% by mole Cl^{35} ?

Sol. $n_1 = 25$ $\quad M_1 = 37$

$n_2 = 75$ $\quad M_2 = 35$

$$M_{av} = \frac{25\times37 + 75\times35}{25+75} = 35.5$$

Exercise

Ag^{107} and Ag^{109}

average atomic mass = 108.5 find % by mole of each isotope

Ans. $Ag^{107} = 25\%$, $Ag^{109} = 75\%$

7.2 AVERAGE MOLECULAR MASS

$$\text{Average molecular mass} = \frac{\text{total mass}}{\text{total mole of molecules}}$$

Let a sample contains n_1 mole of molecules with molecular mass M_1 and n_2 mole of molecules with molecular mass M_2 , then

$$M_{av} = \frac{n_1M_1 + n_2M_2}{n_1 + n_2}$$

Illustration

Air is a mixture of O_2 and N_2 in which O_2 is present 20% by mole and N_2 is present 80% by mole. Find out the average molecular mass of air ?

Sol. $n_1 = 20$ $\quad M_1 = 32$

$n_2 = 80$ $\quad M_2 = 28$

$$M_{av} = \frac{n_1M_1 + n_2M_2}{(n_1 + n_2)} = \frac{20 \times 32 + 80 \times 28}{(20 + 80)} = 28.8$$

Exercise

Two gases A and B [M.wt of A = 20 and M. wt of B = 30] are mixed in the mole ratio a : b and the average molecular mass of the gas mixture is 24. What will be the average molar mass of the gas mixture, if gases A and B are mixed in the ratio b : a ?

Ans. 26

8. EMPIRICAL & MOLECULAR FORMULA

Empirical formula of a compound represents the ratio of different atoms present in a molecule.
Molecular formula of a compound represents the exact no. of atoms present in a molecule.
For Hydrogen peroxide, Empirical formula = HO, Molecular formula = H_2O_2

Steps for writing the empirical formula

The percentage of the elements in the compound is determined by suitable methods and from the data collected, the empirical formula is determined by the following steps:

(i) Divide the percentage of each element by its atomic mass. This will give the relative number of moles of atoms of various elements present in the compound.

(ii) Divide the quotients obtained in the above step by the smallest of them so as to get a simple ratio of moles of various elements.

(iii) Multiply the figures so obtained, by a suitable integer, if necessary, in order to obtain a whole number ratio.

(iv) Finally write down the symbols of the various elements side by side and put the above number as the subscripts to the lower right hand corner of each symbol. This will represent the empirical formula of the compound.

Steps for writing the molecular formula

(i) Calculate the empirical formula as described above.

(ii) Find out the empirical formula mass by adding the atomic masses of all the atoms present in the empirical formula of the compound.

(iii) Divide the molecular mass (determined experimentally by some suitable method) by the empirical formula mass and find out the value of n.

Illustration

Calculate the empirical formula for a compound that contains 26.6% potassium, 35.4% chromium and 38.1% oxygen by mass ?
Given : [Atomic wt : - K = 39; Cr = 52; O = 16]

Sol. Element	Mass Percentage	Atomic mass	Relative no. of atoms	Simple Ratio
K	26.6	39	$\frac{26.6}{39} = 0.68$	$\frac{0.68}{0.68} = 1$
Cr	35.4	52	$\frac{35.4}{52} = 0.68$	$\frac{0.68}{0.68} = 1$
O	38.1	16	$\frac{38.1}{16} = 2.38$	$\frac{2.38}{0.68} = 3.5$

K : Cr : O 1 : 1 : 3.5
2 : 2 : 7 (whole no. ratio)

Empirical formula :- $K_2Cr_2O_7$

Exercise

A carbon compound containing only carbon and oxygen has an approximate molecular mass of 290. On analysis it is found to contain 50% by mass of each element what is the molecular formula of the compound?

Ans. $C_{12}O_9$

9. VAPOUR DENSITY

Some times in numericals molecular mass of volatile substance is not given, instead vapour density is given. Vapour density can be defined as

$$V.D. = \frac{\text{Density of gas at a given T and P}}{\text{Density of } H_2 \text{ at same T and P}}$$

or, $$V.D = \frac{M_{gas}}{2}$$

$$M_{gas} = 2 \times V.D.$$

Illustration

A compound of nitrogen and oxygen was found to contain 7 : 16 by mass N and O respectively. Calculate molecular formula of the compound if V.D. is 46 ?

Sol. Let mass of N = 7 K gram

Mass of O = 16 K gram

Element	Mass	Atomic mass	Relative no. of atoms	Simple ratio
N	7 K	14	$\frac{7K}{14} = 0.5K$	$\frac{0.5K}{0.5K} = 1$
O	16 K	16	$\frac{16K}{16} = K$	$\frac{K}{0.5K} = 2$

N : O = 1 : 2

Empirical formula = NO_2

$(\text{Empirical formula})_n = \text{molecular formula}$

$$n = \frac{\text{M.mass}}{\text{empirical formula mass}} = \frac{2 \times \text{V.D}}{46} = \frac{2 \times 46}{46} = 2$$

Molecular formula $= (NO_2)_2 = N_2O_4$

10. EXPERIMENTAL METHODS FOR DETERMINATION OF ATOMIC MASS & MOLECULAR MASS

10.1 For determination of atomic mass :-

DULONG'S & PETIT'S LAW

In case of heavy solid elements, it is observed that product of atomic mass and specific heat capacity is almost constant.

$$\text{Atomic mass} \times \text{Specific heat capacity (Cal/gm}\ ^{o}\text{C)} \approx 6.4$$

It should be remembered that this law is an empirical observation and this gives an approximate value of atomic mass.

Illustration :

The approximate specific heat capacity of a metal is 0.836 J/gm °C. Find out the approximate atomic mass of the metal?

Ans. Atomic mass $\times$ specific heat ≈ 6.4

$$\text{Atomic mass} \approx \frac{6.4}{\text{specific heat}} = \frac{6.4}{0.836/4.2} = 32$$

10.2 For determination molecular mass :

Some of the most commonly used methods for determination of molecular mass are:

(i) Victor Meyer's Method

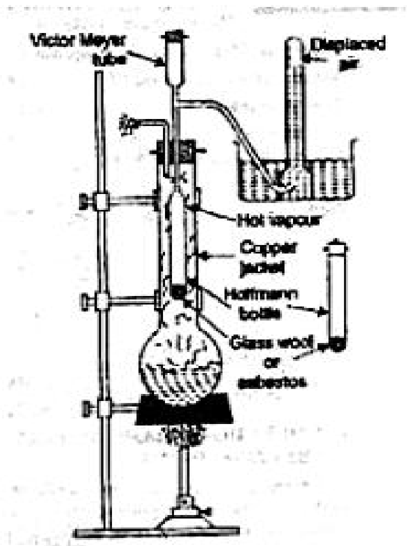
(ii) Silver Salt Method

(iii) Platinic chloride method

10.2.1 VICTOR MEYER'S METHOD

(Applicable for volatile substance)

A known mass of the volatile substance is taken in the Hoffmann's bottle and is vapourised by throwing the Hoffmann's bottle into the Vector Meyer's tube. The vapours displace an equal volume of the moist air and the moist air displaced by the vapours is measured at the room temperature and atmospheric pressure. Following diagram gives the experimental set-up for the Victor - Meyer's process.



Calculations involved

Let the mass of the substance taken be	= W g
Volume of moist vapours collected	= V cm^3
Room temperature	= T K
Barometric pressure	= P mm
Aqueous tension at T K	= p mm
Pressure of dry vapour	= (P - p) mm.

Calculation of molecular mass(M).

$$\frac{(P-p)}{760} \times \frac{V}{1000} = \frac{w}{M} \times RT$$

$$M = \frac{w \times RT \times 760 \times 1000}{(P-p) \times V}$$

Applying PV = nRT for the dry vapour and using n = w/M

Vapour pressure of liquid

The pressure exerted by the vapours in equilibrium with its liquid state is called vapour pressure of liquid. In case of liquid water it is also known as aqueous tension.

It depends only on temperature of liquid and is independent of

(i) Amount of liquid

(ii) Surface area of liquid

(iii) Volume of container

Illustration

0.15 g of a substance displaced 58.9 cm^3 of air at 300 K and 746 mm pressure Calculate the molecular mass. (Aq. Tension at 300 K = 26.7 mm).

Sol. Mass of the substance = 0.15 g
Volume of air displaced (V) = 58.9 cm^3
Temperature (T) = 300 K
Pressure (P) = 746 - 26.7 = 719.3 mm

$$\text{Molecular mass} = \frac{719.3}{760} \times \frac{58.9}{1000} = \frac{0.15}{M} \times 0.821 \times 300$$

$\therefore$ Molecular mass = 66.27 g/mol

10.2.2 SILVER SALT METHOD

(Applicable to organic acids only).

A known mass of the acid is dissolved in water followed by the subsequent addition of silver nitrate solution till the precipitation of silver salt is complete. The precipitate is separated, dried, weighed and ignited till decomposition is complete. The residue of pure silver left behind is weighed.

$$\underset{\text{organic acid}}{HX} \xrightarrow{AgNO_3} \underset{\substack{\text{Silver salt} \\ (W\ g)}}{AgX} \xrightarrow{\text{Ignite}} \underset{\substack{\text{Silver} \\ (x\ g)}}{Ag}$$

Calculations involved

Let the mass of the silver salt formed = W g
The mass of Ag formed = x g
For polybasic acid of the type H_nX (n is basicity)

$$\underset{\text{Organic acid}}{H_nX} \xrightarrow{AgNO_3} \underset{\substack{\text{Silver salt} \\ (W\ g)}}{Ag_nX} \xrightarrow{\text{Ignite}} \underset{\substack{\text{silver} \\ (x\ g)}}{nAg}$$

Mass of silver that gives x g of Ag = Wg

Mass of silver that gives n g- atom (108 g) of $Ag = \frac{108nW}{x} g$

$$\text{Molar Mass of salt} = \frac{108 \times nW}{x} g$$

Molar mass of acid = (molar mass of salt) – n (atomic mass of silver) + n (atomic mass of H)

$$\frac{108 \times nW}{x} - n \times 108 + n \times 1 = n\left(\frac{108W}{x} - 107\right) g\ mol^{-1}$$

Illustration

0.41 g of the silver salt of a dibasic organic acid left a residue to 0.216 g of silver on ignition. Calculate the molecular mass of the acid

Sol. Molecular mass of the silver salt taken (W) = 0.41 g

Mass of Ag formed (x) = 0.216 g

$$H_2X \rightarrow \underset{w=0.41\,g}{Ag_2X} \rightarrow \underset{x=0.216\,g}{2Ag}$$

$$\text{Now molar mass of acid} = n\left(\frac{108W}{x} - 107\right) g\,mol^{-1}$$

$$= 2\left(\frac{108 \times 0.41}{0.216} - 107\right) g\,mol^{-1} = 196 g\,mol^{-1}$$

10.2.3 PLATINIC CHLORIDE METHOD

(Applicable for finding the molecular masses of organic bases.)

A known mass of organic base is allowed to react with chloroplatinic acid (H_2PtCl_6) in presence of conc. HCl to form insoluble platinic chloride. The precipitate of platinic chloride is separated, dried, weighed and is subsequently ignited till decomposition is complete. The residue left is platinum which is again weighed. The molecular mass is then calculated by knowing the mass of the platinic chloride salt and that of platinum left.

If B represents the molecule of monoacidic organic base, then, the formula of platinic chloride salt is $B_2H_2PtCl_6$

$$\underset{\text{Organic base}}{B} \xrightarrow[\text{conc HCl}]{H_2PtCl_6} \underset{\text{Platinic chloride salt}(Wg)}{B_2H_2PtCl_6} \xrightarrow{\text{Ignite}} \underset{(xg)}{Pt}$$

Calculations involved

Let the mass of platinic chloride salt = Wg

The mass of platinum residue left = x g

It may be noted that salt formed with diacidic base would be $B_2(H_2PtCl_6)_2$: with triacidic base would be $B_2(H_2PtCl_6)_3$ and with polyacidic base would be $B_2(H_2PtCl_6)_n$

$$\text{Mass of salt which gives 195g (1 g- atom) of Pt} = \frac{W \times 195 \times n}{x}$$

$$\text{Molar mass of salt} = \frac{W \times 195 \times n}{x} g\,mol^{-1}$$

Now from the formula $B_2(H_2PtCl_6)_n$

Molar mass of salt = [2 × Molar mass of base] + n × [Molar mass of H_2PtCl_6]

Molar mass of base = $\frac{1}{2}(Molar\ mass\ of\ salt - Molar\ mass\ of\ H_2PtCl_6)$

$$= \frac{1}{2}\left(\frac{W \times 195 \times n}{x} - n \times 410\right) = \frac{n}{2}\left[\frac{W \times 195}{x} - 410\right] \text{g mol}^{-1}$$

Illustration

0.98 g of the chloroplatinate of some diacidic base when ignited left 0.39 g of platinum as residue. What is the molecular mass of the base? (At. Mass of Pt = 195)

Sol. Mass of the chloroplatinate salt (W) = 0.98 g

Mass of platinum (x) = 0.39 g

Acidity of the base (n) = 2

Now molar mass of the base $= \frac{n}{2}\left(\frac{W \times 195}{x} - 410\right) \text{gmol}^{-1}$

$$= \frac{2}{2}\left(\frac{0.98 \times 195}{0.39} - 410\right) \text{gmol}^{-1} = 80 \text{ gmol}^{-1}$$

11. EUDIOMETRY - GAS ANALYSIS

The study of gaseous reactions is done in a eudiometer tube with the help of Gay-Lussac's law and Avogadro's law. Eudiometer tube is a closed graduated tube open at one end. The other end is a closed one which is provided with platinum terminals for passing electricity for electric spark, through the known volume of mixture of gases and known volume of oxygen gas. Volume of CO_2 formed is determined by absorbing in KOH solution, O_2 is determined by dissolving unreacted O_2 in alkaline pyrogallol and water vapours formed are determined by noting contraction in volume caused due to cooling. Eudiometry helps :

(i) To study composition of gaseous mixture
(ii) To study volume – volume relationship
(iii) To determine molecular formula of gaseous hydrocarbons and
(iv) To determine molecular formula of gases

11.1 GAY LUSSAC LAW

According to Gay - Lussac's law, the volumes of gaseous reactants reacted and the volumes of gaseous products formed, all are measured at the same temperature and pressure, bear a simple ratio.

$$\underset{\substack{1\text{mol}\\1\text{vol}}}{N_2(g)} + \underset{\substack{3\text{mol}\\3\text{vol}}}{3H_2(g)} \longrightarrow \underset{\substack{2\text{mol}\\2\text{vol}}}{2NH_3(g)}$$

11.2 AVOGADRO'S LAW

In 1812, Amadeo Avogadro stated that samples of different gases which contain the same number of molecules (any complexity, size, shape) occupy the same volume at the same temperature and pressure.

For ideal gas at constant Temperature & Pressure, pressure is directely proportional to no. of moles

11.3 SOME ABSORBENTS OF GASES

The absorbent which is used for specific gas is listed below

Absorbent	Gas or gases absorbed
Turpentine oil	O_3
Alkaline pyrogallol	O_2
Ferrous sulphate solution	NO
Heated magnesium	N_2
Heated palladium	H_2
Ammonical coprous chloride	O_2, CO, C_2H_2 or $CH \equiv CH$
Copper sulphate solution	H_2S, PH_3, AsH_3
Conc. H_2SO_4	H_2O i.e., moisture, NH_3
NaOH or KOH solution	CO_2, NO_2, SO_2, X_2, all acidic oxides

11.4 VOLUME EXPANSION AND CONTRACTION IN THE EUDIOMETER TUBE

$$aA(g) + bB(g) \longrightarrow cC(g) + dD(g)$$

Δn_g = No. of gaseous products – No. of gaseous reactants = $(c + d) - (a + b)$

(a) If $\Delta n_g > 0$, then expansion will occur

(b) If $\Delta n_g = 0$, No contraction / expansion (volume remains constant)

(c) If $\Delta n_g < 0$, then contraction will occur

11.5 ASSUMPTIONS

1. All gases are assumed to be ideal.
2. Nitrogen gas formed during reaction will not react with any other gas.
3. The volume of solids and liquids are negligible in comparision to the volume of gas.

11.6 GENERAL REACTIONS FOR COMBUSTION OF ORGANIC COMPOUNDS

(i) When an organic compound is hydrocarbon :

$$C_xH_y + \left(x + \frac{y}{4}\right) O_2 \rightarrow xCO_2 + \frac{y}{2} H_2O$$

(ii) When an organic compound contain carbon, hydrogen and oxygen :

$$C_xH_yO_z + \left(x + \frac{y}{4} - \frac{z}{2}\right) O_2 \rightarrow xCO_2 + \frac{y}{2} H_2O$$

(iii) When an organic compound contain carbon, hydrogen and nitrogen :

$$C_xH_yN_z + \left(x + \frac{y}{4}\right) O_2 \rightarrow xCO_2 + \frac{y}{2} H_2O + \frac{z}{2} N_2$$

11.7 ANALYSIS OF GASEOUS MIXTURE

Illustration

10 ml of a mixture of CO, CH_4 and N_2, exploded with excess of oxygen, gave a contraction of 6.5 ml. There was a further contraction of 7 ml when the residual gas was treated with KOH. What is the composition of the original mixture ?

Sol. Let the volume of CO = a ml
CH_4 = b ml
$\therefore$ $N_2 = (10 - a - b)$ ml

The explosion reactions are

$$\begin{array}{llll} CO & + 1/2\, O_2 & \rightarrow CO_2 & \\ a & a/2 & a & \\ CH_4 & + 2O_2 & \rightarrow CO_2 & + 2H_2O\,(l) \\ b & 2b & b & \end{array}$$

contraction in volume = a + a/2 + b + 2b – a–b = 6.5
or, a + 4b = 13 ....(1)

CO_2 is absorbed by KOH solution
$\therefore$ a + b = 7 ....(2)
From equation (1) and (2) a = 5, b = 2

$$\therefore \quad \begin{cases} \text{vol. of } CO = 5\text{mL} \\ \text{vol. of } CH_4 = 2\text{mL} \\ \text{vol. of } N_2 = 10 - 5 - 2 = 3\text{mL} \end{cases}$$

Exercise

A mixture of Cl_2 gas (2.8 L) and H_2 gas (3L) was exploded in a eudiometer tube. Calculate the composition by volume of the resulting mixture if all measurements are done under similar conditions of temperature and pressure.

Ans. Volume of H_2 gas left behind = 0.2 L , Volume of HCl gas formed = 5.6 L

11.8 VOLUME - VOLUME ANALYSIS

Illustration

1 litre of a mixture of CO and CO_2 is taken. This mixture is passed through a tube containing red hot charcoal. The volume now becomes 1.6 litres. The volumes are measured under the same condition. Find the composition of the mixture by volume.

Sol. Let the volume of CO_2 in the mixture be x litres

On passing through red hot charcoal, $\underset{\text{x litre}}{CO_2} + C \longrightarrow \underset{\text{2x litre}}{2CO}$

Now, total volume of CO = $(1 - x) + 2x = 1.6$ (given)

$\therefore$ $x = 0.6$ litre

$\therefore$ volume of CO_2 in the mixture = 0.6 litre

volume of CO in the mixture = $1 - 0.6 = 0.4$ litre

Exercise

60 ml of a mixture of nitrous oxide and nitric oxide was exploded with excess of hydrogen. If 38 ml of N_2 was formed, calculate the volume of each gas in the mixture.

Ans. NO = 44 ml

N_2O = 16 ml **Ans.**

11.9 DETERMINATION OF MOLECULAR FORMULA OF GASEOUS HYDROCARBONS

Illustration

A gaseous hydrocarbon (V.D. = 15) having volume 10 ml was exploded with excess of O_2. On cooling, a contraction in volume by 25 ml was observed. Calculate the molecular formula of the hydrocarbon.

Sol. Let the molecular formula of hydrocarbon = C_xH_y

Reaction $\quad C_xH_y + \left(x + \frac{y}{4}\right)O_2 \longrightarrow xCO_2 + \frac{y}{2}H_2O$

$\quad\quad 10\text{ ml} \quad 10\left(x + \frac{y}{4}\right)\text{ ml} \quad\quad 10\,x\text{ ml}$

From reaction contraction in volume is given as : $10 + 10\left(x + \frac{y}{4}\right) - 10x = 25$ (given)

$\therefore$ $y = 6$

Now, Formula of hydrocarbon = C_xH_6

Mol. wt. of $C_xH_6 = (12 \times x) + (6 \times 1) = 2 \times 15$

$\therefore x = 2$

Hence, Molecular formula of hydrocarbon = C_2H_6

Exercise

500 ml of a hydrocarbon gas burnt in excess of oxygen yields 2500 ml of CO_2 and 3.0 litres of water vapours, all the volume being measured at the same temperature and pressure. What is the formula of hydrocarbon gas?

Ans. C_5H_{12}

11.10 DETERMINATION OF MOLECULAR FORMULA OF GASES

Illustration

60 ml of a mixture of equal volumes of Cl_2 and an oxide of chlorine were heated and then cooled back to the original temperature. The resulting gas mixture was found to have a volume of 75 ml. On treatment with caustic soda, the volume contracted to 15 ml. Assuming that all measurements where made at the same temperature & pressure, deduce the simplest formula of the oxide of chlorine. (The oxide of chlorine on heating decomposes quantitatively to give O_2 & Cl_2)

Sol. Let Cl_xO_y be the oxide.

$$Cl_xO_y \rightarrow \frac{x}{2} Cl_2 + \frac{y}{2} O_2$$

1 Vol.	$\frac{x}{2}$ Vol.	$\frac{y}{2}$ Vol.
30 ml	15 x ml	15y ml

Volume of Cl_xO_y = Vol. of Cl_2 initially = 30 ml
After cooling volume = 75 ml
This corresponds to volume of Cl_2 initially plus volume of Cl_2 produced & O_2 produced.

$$V(Cl_2) + V(Cl_2 \text{ produced}) + V(O_2) = 75 \text{ ml}$$

NaOH absorbs Cl_2 apart from CO_2. So after NaOH treatment, the residual volume corresponds to the volume of O_2 = 15 ml
and V (Cl_2 produced) = 75 – V (Cl_2) – V(O_2) = 75 – 30 – 15 = 30 ml
$15x = 30 \Rightarrow x = 2$
and $15y = 15 \Rightarrow y = 1$
Hence, formula of oxide of chloride is Cl_2O.

Exercise

50 ml of pure and dry oxygen was subjected to silent electric discharge and on cooling to the original temperature, the volume of ozonised oxygen was found to be 47 ml. The gas was then brought in contact with turpentine oil, the remaining gas occupied a volume of 41 ml. Find the molecular formula of ozone.

Ans. O_3

12. CONCENTRATION TERMS

A solution is a homogeneous mixture of two or more substances, the composition of which may vary within limits. "A solution is a special kind of mixture in which substances are intermixed so intimately that they can not be observed as separate components". The substance which is to be dissolved is called **solute** while the medium in which the solute is dissolved to get a homogeneous mixture is called the **solvent.** A solution is termed as binary and ternary if it consists of two and three components respectively.

12.1 METHODS OF EXPRESSING CONCENTRATION OF SOLUTION

Concentration of solution is the amount of solute dissolved in a known amount of the solvent or solution. The concentration of solution can be expressed in various ways as discussed below.

12.1.1 Percentage : It refers to the amount of the solute per 100 parts of the solution. It can also be called as parts per hundred (pph). It can be expressed by any of following four methods :

(i) Weight by weight percentage (%w/w) $= \dfrac{\text{Wt.}\,of\ solute\ (g)}{Wt.\,of\ solution\ (g)} \times 100$

e.g., 10% Na_2CO_3 solution w/w means 10g of Na_2CO_3 is dissolved in 100g of the solution. (It means 10g Na_2CO_3 is dissolved in 90 of solvent)

(ii) Weight by volume percent (% w/v) $= \dfrac{\text{Wt.}\,of\ solute(g)}{Volume\ of\ solution\ (cm^3)} \times 100$

e.g., 10% Na_2CO_3 (w/v) means 10g Na_2CO_3 is dissolved in 100 cm^3 of solution.

(iii) Volume by volume percent (% v/v) $= \dfrac{\text{Volume}\,of\ solute\ (cm^3)}{Volume\ of\ solution(cm^3)} \times 100$

e.g., 10% ethanol (v/v) means 10cm^3 of ethanol dissolved in 100 cm^3 of solution.

(iv) Volume by weight percent (%v/w) $= \dfrac{\text{Vol. of solute}}{\text{Wt. of solution}} \times 100$

e.g., 10% ethanol (v/w) means 10cm^3 of ethanol dissolved in 100g of solution.

Illustration

Concentrated nitric acid used as laboratory reagent is usually 69% by mass of nitric acid. Calculate the volume of the solution which contains 23 g nitric acid. The density of concentrated acid is 1.41 g cm^{-3}.

Sol. Given HNO_3 is 69% by mass;
density of HNO_3 = 1.41 g cm^{-3}.
Thus (i) 69 g HNO_3 is present in conc. HNO_3 = 100 g

$$23 \text{ g } HNO_3 \text{ is present in conc. } HNO_3 = \frac{100}{69} \times 23 = 33.33 \text{ g}$$

$$\text{(ii) Volume of solution required} = \frac{\text{Mass}}{\text{Density}} = \frac{33.33}{1.41} = \mathbf{23.64\ mL}$$

Exercise

Calculate % w/w of NaOH in a solution containing 40% w/v NaOH . Density of solution is d (g/ml).

Ans. $\% \text{ by mass} = \frac{40}{d}\%$

12.1.2 Parts per million (ppm) and parts per billion (ppb) : When a solute is present in very small quantity, it is convenient to express the concentration in parts per million and parts per billion. It is the number of parts of solute per million (10^6) or per billion (10^9) parts of solution. It is independent of the temperature.

$$\text{ppm} = \frac{\text{Mass of solute component}}{\text{Total mass of solution}} \times 10^6$$

$$\text{ppb} = \frac{\text{Mass of solute component}}{\text{Total mass of solution}} \times 10^9$$

Illustration

Calculate the parts per million of SO_2 gas in 250 ml water (density 1 g cm^{-3}) containing 5×10^{-4} g of SO_2 gas.

Sol. Mass of SO_2 gas = 5×10^{-4} g;
Mass of H_2O = Volume × Density = 250 cm^3 × 1 g cm^{-3} = 250 g

$$\therefore \text{ Parts per million of } SO_2 \text{ gas} = \frac{5 \times 10^{-4}}{250\text{g}} \times 10^6 = \mathbf{2}$$

Exercise

Units of parts per million (ppm) or per billion (ppb) are often used to describe the concentrations of solutes in very dilute solutions. The units are defined as the number of grams of solute per million or per billion grams of solvent. Bay of Bengal has 1.9 ppm of lithium ions. What is the molality of Li^+ in this water ?

Ans. 2.7×10^{-4}

12.1.3 Normality (N) : It is defined as the number of gram equivalents of a solute present per litre of the solution. Unit of normality is gram equivalents litre^{-1}. Normality changes with temperature since it involves volume. When a solution is diluted x times, its normality also decreases by x times. Concentration in terms of normality is generally expressed as,

N = Normal solution; 5N = Penta normal, 10N = Deca normal;

N/2 = semi normal; N/10 = Deci normal; N/5 = Penti normal

N/100 or 0.01 N = centinormal, N/1000 or 0.001 = millinormal.

Mathematically normality can be calculated by following formulas,

(i) Normality

$$N = \frac{\text{Number of gram equivalents of solute}}{\text{Volume of soltuion}(\ell)} = \frac{\text{Weight of solute in gram}}{\text{equivalent weight of solute} \times \text{Volume of solution}(\ell)}$$

$$= \frac{\text{wt. of solute}}{\text{eq. wt. of solute}} \times \frac{1000}{\text{wt. of solution (ml)}}$$

$$\text{Equivalent weight of solute} = \frac{\text{Gram molar mass}}{\text{Acidity/Basicity/Cation valency}}$$

Acidity = No. of replaceable OH^- in a molecule of a base

For example : Acidity of NaOH = 1

Basicity = No. of replaceable H^+ in a molecule of an acid

For example : Basicity of $H_2SO_4 = 2$

(ii) If volume V_1 and normality N_1 is so changed that new normality is N_2 and volume is V_2 then,

$\boxed{N_1V_1 = N_2V_2}$ (No. of equivalents remains same in mixing and dilution)

(iii) When two solutions of the same solute and solvent are mixed then normality of mixture (N) is

$$\boxed{N = \frac{N_1V_1 + N_2V_2}{V_1 + V_2}}$$

Illustration

Calculate normality of 7g /L H_2SO_4.

Sol. $N_{H_2SO_4} = \frac{g/L}{Eq.wt.} = \frac{7}{49} = \frac{1}{7} = 0.143$ g eq l^{-1}

Exercise

How many grams of oxalic acid are to be dissolved in 250 ml water to prepare 0.1 N solution? (Eq. wt. of ox. acid = 63)

Ans. 1.575g

12.1.4 Molarity (M) : Molarity of a solution is the number of moles of the solute per litre of solution (or number of millimoles per ml of solution). Unit of molarity is mol/litre or mol/dm^3. For example, a molar (1M) solution of sugar means a solution containing 1 mole of sugar per litre of the solution. Solutions in terms of molarity is generally expressed as,

1M = One molar solution, 2M = Molarity is two, $\frac{M}{2}$ or 0.5M = Semimolar solution,

$\frac{M}{10}$ or 0.1M = Decimolar solution, $\frac{M}{100}$ or 0.01 M = Centimolar solution

$\frac{M}{1000}$ or 0.001 M = Millimolar solution

Mathematically, molarity can be calculated by following formulas :

(i) $M = \frac{\text{No. of moles of solute(n)}}{\text{Vol. of solution in litres}} = \frac{\text{wt. of solute (gm)}}{\text{gm mol.wt. of solute}} \times \frac{1000}{\text{wt. of solution (ml)}}$

(ii) If molarity and volume of the solution are changed from M_1, V_1 to M_2, V_2. Then,

$$M_1V_1 = M_2V_2$$

(iii) In balanced chemical equation, if n_1 moles of reactant-1 react with n_2 moles of reactant-2 . Then,

$$n_1A + n_2B \longrightarrow \text{Product}$$

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$$

(iv) If two solutions of the same solute are mixed then molarity (M) of resulting solution

$$M = \frac{M_1V_1 + M_2V_2}{(V_1 + V_2)}$$

Illustration

A bottle of commercial sulphuric acid (density 1.787 g ml^{-1}) is labelled as 86% by weight. What is the molarity of acid?

Sol. (i) Molarity of $H_2SO_4 = \frac{\text{Wt. of } H_2SO_4 \text{ in 1L solution}}{\text{mol. wt. of } H_2SO_4}$

But wt. of given H_2SO_4 per litre $= \frac{86}{100} \times 1.787 \times 1000 = 1536.82$ g.

Hence molarity of $H_2SO_4 = \frac{1536.82}{98} = 15.68 \text{ mol L}^{-1}$

Exercise

A sample contains I_2 and benzene. The mole fraction of $I_2 = 0.2$. Calculate molarity of solution if
(i) density of solution is d gm/ml (ii) density of I_2 & benzene are d_{I_2} & $d_{benzene}$

Ans. (i) $M = 1.77\,d$ (ii) $M = \frac{0.2}{\left(\frac{50.8}{d_{I_2}} + \frac{62.4}{d_{benzene}}\right)} \times 1000$

12.1.5 Formality (F) : Formality of solution may be defined as the number of gram formula units of the ionic solute dissolved per litre of the solution. It is represented by **F**. Commonly, the term formality is used to express the concentration of the ionic solids which do not exist as molecules but exist as network of ions. A solution containing one gram formula mass of solute per litre of the solution has formality equal to one and is called **Formal solution**. It may be mentioned here that the formality of a solution changes with change in temperature.

$$\text{Formality(F)} = \frac{\text{Number of gram formula units of solute}}{\text{Volume of solution in litres}}$$

$$= \frac{\text{Mass of ionic solute(g)}}{\text{gram formula unit mass of solute} \times \text{Volume of solution (l)}}$$

Illustration

What will be the formality of KNO_3 solution having strength equal to 2.02 g per litre ?

Sol. Strength of $KNO_3 = 2.02 \text{ gL}^{-1}$
and g formula weight of $KNO_3 = 101$ g

$\therefore$ Formality of $KNO_3 = \frac{\text{strength in } gl^{-1}}{\text{g.formula wt.of } KNO_3} = \frac{2.02}{101} = \mathbf{0.02\ F}$

Exercise

Calculate the formality of NaCl solution, 5.85 g of which have been dissolved to form 250 ml of the given solution

Ans. 0.4 F

12.1.6 Molality (m) : It is the number of moles of the solute per 1000g of the solvent. Unit of molality is mol/kg. For example, a 0.2 molal (0.2m) solution of glucose means a solution obtained by dissolving 0.2 mole of glucose in 1000 gm of water. Molality (m) does not depend on temperature since it involves measurement of weight of liquids.

Mathematically molality can be calculated by following formulas,

(i) $$m = \frac{\text{Number of moles of solute}}{\text{Weight of solvent in kg}} = \frac{\text{Number of moles of solute}}{\text{Weight of solvent in gm}} \times 1000$$

(ii) $$= \frac{\text{Wt. of solute}}{\text{Mol. wt. of solute}} \times \frac{1000}{\text{Weight of solvent in gm}}$$

Illustration

The density of a 3 M $Na_2S_2O_3$ solution is 1.25 g ml^{-1}. Calculate percentage by mass of $Na_2S_2O_3$ and molalities of Na^+ and $S_2O_3^{2-}$ ions.

Sol. (i) Total mass of $Na_2S_2O_3$ solution = Vol of solution × Density of solution

= 1000 ml × 1.25 g ml^{-1} = 1250 g

Wt. of 3M $Na_2S_2O_3$ in 1000 ml solution = 3 × Mol. wt of $Na_2S_2O_3$ (158) = 474 g

[$\because$ Mol. wt. of $Na_2S_2O_3$ = 158 g mol^{-1}]

$$\therefore \% \text{ of } Na_2S_2O_3 = \frac{\text{Wt. of } Na_2S_2O_3}{\text{Total wt. of solution}} \times 100 = \frac{474}{1250} \times 100 = 37.92\%$$

(ii) To find molality of Na^+ ions

$\because$ 3 M $Na_2S_2O_3$ contain 6 moles of Na^+ ions in one litre

$$\therefore \text{Molality of } Na^+ \text{ ions} = 6 \times \frac{1000}{\text{wt. of solvent } (1250 - 474 = 776\text{g})} = 6 \times \frac{1000}{776} = 7.73 \text{ m}$$

(iii) To find molality of $S_2O_3^{2-}$

$\because$ 3M $Na_2S_2O_3$ contain 3 moles of $S_2O_3^{2-}$ ions in one litre

$$\text{Molality of } S_2O_3^{2-} = 3 \times \frac{1000}{\text{wt. of solvent} = 776} = 3.865 \text{ m}$$

Exercise

Find molality of 40% by mass (w/w) NaOH.

Ans. $\frac{50}{3}$ m

12.1.7 Mole fraction (χ) : Mole fraction may be defined as the ratio of number of moles of one component to the total number of moles of all the components (solvent and solute) present in the solution. It is denoted by the letter χ. It may be noted that the mole fraction is independent of the temperature. Mole fraction is dimensionless. If a solution contains the components A and B and suppose that W_A gram of A and W_B gram of B are present in it.

Number of moles of A is given by, $n_A = \frac{W_A}{M_A}$ and the number of moles of B is given by, $n_B = \frac{W_B}{M_B}$

where M_A and M_B are molecular mass of A and B respectively.

Total number of moles of A and B = $n_A + n_B$

Mole fraction of A, $\chi_A = \frac{n_A}{n_A + n_B}$

Mole fraction of B, $\chi_B = \frac{n_B}{n_A + n_B}$

The sum of mole fractions of all the components in the solution is always one.

$$\chi_A + \chi_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B} = 1$$

$$\chi_A + \chi_B = 1$$

Thus, if we know the mole fraction of one component of a binary solution, the mole fraction of the other can be calculated.

Illustration

Find out the masses of acid and water required to prepare 1 mole of CH_3COOH solution of 0.3 mole fraction of CH_3COOH.

Sol. $\chi_{CH_3COOH} = 0.3$

$\chi_{H_2O} = 1 - 0.3 = 0.7$

Wt. of CH_3COOH = χ_{CH_3COOH} × mol. wt.(CH_3COOH) = 0.3 × 60 = **18 g**

Wt. of water = χ_{H_2O} × mol. wt. (H_2O) = 0.7 × 18 = **12.6 g**

Exercise

Mole fraction of a solute in benzene is 0.2. Find out the molality of the solution.

Ans. 3.2 mol kg^{-1}

12.1.8 Mass Fraction : Mass fractino of a component in a solution is the mass of the component divided by the total mass of the solution. For a solution containing w_A gm of A and w_B gm of B.

Mass fraction of A = $\frac{w_A}{w_A + w_B}$ Mass fraction of B = $\frac{w_B}{w_A + w_B}$

Note : It may be noted that molality, mole fraction, mass fraction etc. are preferred to molarity, normality, formality etc. Because the former involve the weights of the solute and solvent where as later involve volumes of solutions. Temperature has no effect on weights but it has significant effect on volumes.

12.2 MOLARITY OF IONIC COMPOUNDS

In ionic compounds, calculate the dissociated mole of each ion per mole of molecule considering the degree of dissociation.

Illustration

Find the molarity of various ions in 0.1 M $BeCl_2$ solution, considering degree of dissociation of $BeCl_2$ as 80%.

Sol. $BeCl_2 \xrightarrow{80\%} Be^{+2} + 2Cl^-$

0.1	0	0
0.1–0.08	0.08	0.16

Be^{+2} = 0.08M, Cl^- = 0.16 M

Exercise

Calculate 'm' of all the ions present in the solution of $Al_2(SO_4)_3$ for 1M solution.
Given $d_{solution}$ = 2.342 gm/ml.

Ans. Al^{3+} = 1m, SO_4^{2-} = 1.5m

12.3 MIXING OR DILUTION OF SOLUTIONS

Illustration

20 ml N/2 HCl, 60 ml N/10 H_2SO_4 and 150 ml N/5 HNO_3 are mixed together. Calculate the normality of the $[H^+]$ in final solution.

Sol. $N_1V_1 + N_2V_2 + N_3V_3 = NV$

$$\left(\frac{1}{2}\times 20\right) + \left(\frac{1}{10}\times 60\right) + \left(\frac{1}{5}\times 150\right) = N\ (230)$$

$10 + 6 + 30 = N\ (230)$

$46 = N\ (230)$

$N = 0.2$

Exercise

200 ml of 1M HCl solution is mixed with 800 ml of 2M HCl solution.

(a) Calculate the final molarity of the solution.

(b) If density of final solution is 1.2 gm/ml. Calculate molality 'm'

Ans (a) 1.8M, (b) 1.5 m

12.4 MIXING OF ACID & BASE SOLUTIONS

In case of mixing of acid and base, calculate equivalent or milliequivalent of acid and base separately, substract the lower value from higher value, which gives the number of equivalent in the final solution.

Illustration

50 ml N/2 HCl, 50 ml N/5 H_2SO_5 and 200 ml N/10 NaOH are mixed. What will be normality of $[H^+]$ ions ?

Sol. $\text{meq of acids} = \left(50\times\frac{1}{2}\right) + \left(50\times\frac{1}{5}\right) = 25 + 10 = 35$

$\text{meq of base} = 200 \times \frac{1}{10} = 20$

After mixing meq of acid $= 35 - 20 = 15$ $\therefore$ Normality of H^+ ion, $N = \frac{15}{300} = 0.05$ **Ans.**

Exercise

Calculate $[H^+]$ in a solution if 0.2 M 100 ml H_2SO_4 is mixed with 0.1 M 300 ml NaOH solution.

Ans. 0.025M

12.5 PROBLEM INVOLVING PRECIPITATION

In problems involving precipitation, complete reaction is assumed and the limiting reagent completely reacts. Ions in the precipitated product is not considered in finding concentration of that ions.

Illustration

10 ml of HCl solution gave 0.1435 g of AgCl when treated with excess of $AgNO_3$. Find the molarity of the acid solution (Ag = 108)

Sol.

HCl +	$AgNO_3$	$\rightarrow$	AgCl +	HNO_3
(1+35.5)	(excess)		(108 + 35.5)	
= 36.5 g			= 143.5 g	

143.5 g of AgCl is obtained from HCl = 36.5 g

0.1435 g of AgCl is obtained from HCl = $\frac{36.5}{143.6} \times 0.1435\text{g} = 0.0365$

$\therefore$ Molarity of HCl solution $= \frac{0.0365}{36.5} \times \frac{1000}{10} = 0.1\text{ M}$

Exercise

If 200 ml of 0.1 NaCl in mixed with 100 ml of 0.2 M $AgNO_3$ solution. Calculate molarity of all the ions in the final solution.

Ans. $Na^+ = 0.067\text{ M},\ NO_3^- = 0.067\text{ M}$

12.6 SOME TYPICAL CONCENTRATION TERMS

12.6.1 Strength of H_2O_2 solution :

The strength of H_2O_2 is aqeous solution is expressed in the following two ways :

(i) **Percentage strength :** The mass of H_2O_2 present in 100 ml of the aqueous solution is termed as percentage strength. For example, a 25% solution (w/v) of H_2O_2 means that 25 grams of H_2O_2 are present in 100 ml of the solution.

(ii) **Volume strength :** Strength of the sample of H_2O_2 is generally expressed in terms of the volume of oxygen at 0°C and 1 atm that one volume of the sample of hydrogen peroxide gives on heating. The commercial samples are marked as '10 volume'. '15 volume' or '20 volume'. 10 volume means that one volume of the sample of hydrogen peroxide gives 10 volumes of oxygen at 0°C and 1 atm. 1 ml of a 10 volume solution of H_2O_2 will liberate 10 ml of oxygen at 0°C and 1 atm.

Illustration

Find the volume strength of 1.6 M H_2O_2 solution.

Sol. Strength of the solution = Molarity × mol. mass = 1.6 × 34 = 54.4 gL^{-1}

$2H_2O_2 \rightarrow$	$2H_2O + O_2$	
68 g		22.4 L
54.4		$\frac{22.4}{68} \times 54 = 19.92\text{L}$

$\therefore$ Volume strength = **'19.92 V'**

Exercise

For '44.8 V' H_2O_2 solution having d = 1.136 g/ml calculate
(i) Molarity of H_2O_2 solution. (ii) Mole fraction of H_2O_2 solution.

Ans. (i) 4 M ; (ii) 0.06

12.6.2 Percentage labelling of oleum

Oleum is fuming sulphuric acid which contains extra SO_3 dissolved in H_2SO_4. To convert this extra SO_3 into H_2SO_4, water has to be added $(SO_3 + H_2O \rightarrow H_2SO_4)$. The amount of sulphuric acid obtained when just sufficient water is added into 100 g of oleum so that all SO_3 present in it is converted into H_2SO_4 is called percentage labelling of oleum.

Illustration

An oleum is labelled as 109%. Calculate mass percent of free SO_3 and H_2SO_4

Sol. It means that water added = 9 g in 100 gm oleum

Let us calculate how much SO_3 can be converted into H_2SO_4 by 9 g of water

$$\underset{80\text{ g}}{SO_3} + \underset{18\text{ g}}{H_2O} \rightarrow \underset{98\text{ g}}{H_2SO_4}$$

i.e., 9 g H_2O can dissolve 40 g SO_3 to form 49 g H_2SO_4

$\therefore$ Mass of SO_3 in 100 g oleum = 40 g

and % by mass of SO_3 = 40 %

and % by mass of H_2SO_4 = 60 %

Exercise

A mixture is prepared by mixing 10 gm H_2SO_4 and 40 gm SO_3 calculate,
(a) mole fraction of H_2SO_4 (b) % labelling of oleum

Ans. (a) 0.169; (b) 118 %

12.7 RELATIONSHIP BETWEEN DIFFERENT CONCENTRATION TERMS

1. $N = M \times n \text{ factor}$
2. $M = \dfrac{md}{1 + mM_2/1000}$
3. $m = \dfrac{1000 \times x_2}{x_1 M_1}$
4. $M = \dfrac{1000dx_2}{x_1M_1 + x_1M_2}$
5. $d = M\left(\dfrac{1}{m} + \dfrac{M_2}{1000}\right)$
6. Volume strength of $H_2O_2 = 5.6 \times N = \dfrac{5.6 \times \text{Percentagestrength}}{\text{Eq.wt.of } H_2O_2 (17)} \times 10$

7. Volume strength of $H_2O_2 = 11.2 \times M = \dfrac{11.2 \times \text{Percentagestrength} \times 10}{\text{Mol.wt. of } H_2O_2\,(34)}$

8. In oleum labelled as $(100 + x)\%$

$$\%\text{ of free } SO_3 = \left(\frac{80 \times x}{18}\right) (\text{w/w})$$

where N = Normality

M = Molarity	m = molality
d = density of solution	M_2 = Molecular mass of solute
x_2 = Mole fraction of solute	x_1 = Mole fraction of solvent
M_1 = Molecular mass of solvent	d = Density of solution.

SOLVED EXAMPLES

Q.1 Naturally occurring chlorine is 75.53% Cl^{35} which has an atomic mass of 34.969 amu and 24.47% Cl^{37} which has a mass of 36.966 amu. Calculate the average atomic mass of chlorine-

(A) 35.5 amu (B) 36.5 amu (C) 71 amu (D) 72 amu

Ans. (A)

Sol. Average atomic mass $= \dfrac{\%\text{ of I isotope} \times \text{its atomic mass} + \%\text{ of II isotope} \times \text{its atomic mass}}{100}$

$$= \frac{75.53 \times 34.969 + 24.47 \times 36.96}{100}$$

= 35.5 amu.

Q.2 How many carbon atoms are present in 0.35 mol of $C_6H_{12}O_6$ -

(A) 6.023×10^{23} carbon atoms (B) 1.26×10^{23} carbon atoms

(C) 1.26×10^{24} carbon atoms (D) 6.023×10^{24} carbon atoms

Ans. (C)

Sol. $\therefore$ 1 mol of $C_6H_{12}O_6$ has = 6 N_A atoms of C

$\therefore$ 0.35 mol of $C_6H_{12}O_6$ has

= $6 \times 0.35\ N_A$ atoms of C

= $2.1\ N_A$ atoms

= $2.1 \times 6.022 \times 10^{23}$ **= 1.26×10^{24} carbon atoms**

Q.3 Calculate the mass in gm of $2N_A$ molecules of CO_2 -

(A) 22 gm (B) 44 gm (C) 88 gm (D) None of these.

Ans. (C)

Sol. $\because$ N_A molecules of CO_2 has molecular mass = 44 gm

$\therefore$ $2N_A$ molecules of CO_2 has molecular mass = 44×2 = **88 gm.**

Q.4 How many years it would take to spend Avogadro's number of rupees at the rate of 1 million rupees in one second -

(A) 19.098×10^{19} years (B) 19.098 years

(C) 19.098×10^{9} years (D) None of these

Ans. (C)

Sol. $\because$ 10^6 rupees are spent in 1sec.

$\therefore$ 6.022×10^{23} rupees are spent in $= \dfrac{1 \times 6.022 \times 10^{23}}{10^6}$ sec

$= \dfrac{1 \times 6.023 \times 10^{23}}{10^6 \times 60 \times 60 \times 24 \times 365}$ years **= 19.098×10^9 year**

Q.5 Calculate the number of Cl^- and Ca^{+2} ions in 222 g anhydrous $CaCl_2$.

(A) $2N_A$ ions of Ca^{+2} 4 N ions of Cl^- (B) $2N_A$ ions of Cl^- & 4N ions of Ca^{+2}

(C) $1N_A$ ions of Ca^{+2} & 1N ions of Cl^- (D) None of these.

Ans. (A)

Sol. $\because$ mol. wt. of $CaCl_2$ = 111 g

$\because$ 111 g $CaCl_2$ has = N_A ions of Ca^{+2}

$\therefore$ 222g of $CaCl_2$ has $\dfrac{N_A \times 222}{111}$ = $2N_A$ ions of Ca^{+2}

Also $\because$ 111 g $CaCl_2$ has = $2N_A$ ions of Cl^-

$\therefore$ 222 g $CaCl_2$ has = $\frac{2N_A \times 222}{111}$ ions of Cl^- **= $4N_A$ ions of Cl^- .**

Q.6 What is the molecular mass of a substance, each molecule of which contains 9 carbon atoms, 13 hydrogen atoms and 2.33×10^{-23} g of other component?

Ans. 135.04

Sol. The molecule has C, H and other component.

Mass of 9 C atoms = 12 × 9 = 108 amu

Mass of 13 H atoms = 13 × 1 = 13 amu

Mass of other component = $\frac{2.33\times10^{-23}}{1.66\times10^{-24}} = 14.04 amu$

$\therefore$ Total mass of one molecule = 108 + 13 + 14.04 = 135.04 amu

$\therefore$ Mol. mass of substance **= 135.04**

Q.7 The density of O_2 at 0°C and 1 atm is 1.429g / litre. The molar volume of gas is -

(A) 22.4 lit. (B) 11.2 lit (C) 33.6 lit (D) 5.6 lit.

Ans. (A)

Sol. $\because$ 1.429 gm of O_2 gas occupies volume = 1 litre.

$\therefore$ 32 gm of O_2 gas occupies = $\frac{32}{1.429}$ **= 22.4 litre/mol.**

Q.8 How many molecules are in 5.23 gm of glucose ($C_6H_{12}O_6$) -

(A) 1.65×10^{22} (B) 1.75×10^{22} (C) 1.75×10^{21} (D) None of these

Ans. (B)

Sol. $\because$ 180 gm glucose has = N_A molecules

$\therefore$ 5.23 gm glucose has = $\frac{5.23 \times 6.022 \times 10^{23}}{180}$ **= 1.75×10^{22} molecules**

Q.9 How many g of S are required to produce 10 moles and 10g of H_2SO_4 respectively?

Ans. 320 g, 3.265 g

Sol. $\because$ 1 mole of H_2SO_4 has = 32g S

$\therefore$ 10 mole of H_2SO_4 has = 32 × 10 = **320 g S**

Also, 98g of H_2SO_4 has = 32 g S

$\therefore$ 10 g of H_2SO_4 has = (32 × 10)/98 = **3.265 g S**

Q.10 P and Q are two elements which form P_2Q_3 and PQ_2 molecules. If 0.15 mole of P_2Q_3 and PQ_2 weighs 15.9 g and 9.3g, respectively, what are atomic mass of P and Q?

Ans. P = 26, Q = 18

Sol. Let at. mass of P and Q be a and b respectively,

$\therefore$ Mol. mass of P_2Q_3 = 2a + 3b

and Mol. mass of PQ_2 = a + 2b

$\therefore$ (2a + 3b) × 0.15 = 15.9

and (a + 2b) × 0.15 = 9.3

a = 26, b = 18

atomic mass of **P = 26**

atomic mass of **Q = 18**

Q.11 A hydrate of iron (III) thiocyanate $Fe(SCN)_3$, was found to contain 19% H_2O. What is the formula of the hydrate?

Ans. $Fe(SCN)_3.3H_2O$

Sol. Let the hydrate be $Fe(SCN)_3.mH_2O$

Molecular mass of hydrate = 56 + 3 × (32 + 12 + 14) + 18m = 230 + 18m

$\therefore \quad \%\text{ of } H_2O = \dfrac{18m \times 100}{230 + 18m} = 19 \quad \text{or} \quad m = 2.99 \approx 3$

$\therefore$ Formula is **$Fe(SCN)_3.3H_2O$**

Q.12 The vapour density of a mixture containing NO_2 and N_2O_4 is 38.3 at 27°C. Calculate the mole of NO_2 in 100 mole mixture.

Ans. 33.48

Sol. Mol. mass of mixture of NO_2 and N_2O_4 = 38.3 × 2 = 76.6

Let a mole of NO_2 be present in 100 mole mixture

= mass of NO_2 + mass of N_2O_4 = mass of mixture,

a × 46 + (100 – a) × 92 = 100 × 76.6

$\therefore$ a = **33.48 mole**

Q.13 What is the weight of 3.01×10^{23} molecules of ammonia -

(A) 17 gm (B) 8.5 gm (C) 34 gm (D) None of these

Ans. (B)

Sol. $\because$ 6.022×10^{23} molecules of NH_3 has weight = 17 gm

$\therefore$ 3.01×10^{23} molecules of NH_3 has weight $= \dfrac{17 \times 3.01 \times 10^{23}}{6.022 \times 10^{23}}$ = **8.50 gm**

Q.14 How many moles of potassium chlorate to be heated to produce 5.6 litre oxygen at 0°C and 1 atm ?

Ans. 1/6

Sol.

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

Mole ratio for reaction: 2 mole, 2 mole, 3 mole

$\because$ 3 × 22.4 litre O_2 is formed by 2 mole $KClO_3$

5.6 litre O_2 is formed by $\dfrac{2 \times 5.6}{3 \times 22.4}$ = **1/6 mole $KClO_3$**

Q.15 How many molecules are present in 1mℓ of water vapours at 0°C and 1 atm -

(A) 1.69×10^{19} (B) 2.69×10^{-19} (C) 1.69×10^{-19} (D) 2.69×10^{19}

Ans. (D)

Sol. $\because$ 22.4 litre water vapour at 0°C and 1 atm has = 6.022×10^{23} molecules

$\therefore$ 1×10^{-3} litre water vapours has $= \dfrac{6.022 \times 10^{23}}{22.4} \times 10^{-3}$ = **2.69×10^{19}**

Q.16 Calculate the weight of lime (CaO) obtained by heating 200 kg of 95% pure lime stone ($CaCO_3$).

(A) 104.4 kg (B) 105.4 kg

(C) 212.8 kg (D) 106.4 kg

Ans. (D)

Sol. 100 kg impure sample has pure $CaCO_3 = 95$ kg

$\therefore$ 200 kg impure sample has pure $CaCO_3 = \frac{95 \times 200}{100} = 190$ kg.

$CaCO_3 \rightarrow CaO + CO_2$

$\because$ 100 kg $CaCO_3$ gives $CaO = 56$ kg.

$\therefore$ 190 kg $CaCO_3$ gives $CaO = \frac{56 \times 190}{100}$ = **106.4 kg.**

Q.17 Zinc and hydrochloric acid react according to the reaction :

$$Zn(s) + 2HCl\,(aq.) \longrightarrow ZnCl_2(aq.) + H_2(g)$$

If 0.30 mole of Zn are added to hydrochloric acid containing 0.52 mole HCl, how many moles of H_2 are produced?

Ans. 0.26

Sol.

	$Zn(s)$ +	$2HCl(aq.) \longrightarrow$	$ZnCl_2(aq.)$ +	$H_2(g)$
Initial moles	0.30	0.52	0	0
Final moles	0.04	0	0.26	0.26

Moles of H_2 produced = **0.26**

Q.18 4 g of an impure sample of $CaCO_3$ on treatment with excess HCl produces 0.88 g CO_2. What is percent purity of $CaCO_3$ sample?

Ans. 50%

Sol. $CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$

44g $CO_2 \equiv 100$ g $CaCO_3$

$$0.88 \text{ g } CO_2 \equiv \frac{100 \times 0.88}{44} = 2.0 \text{ g } CaCO_3$$

$\therefore$ Percentage purity $= \frac{2}{4} \times 100 =$ **50%**

Q.19 A mixture of FeO and Fe_3O_4 when heated in air to constant weight gains 5% in its weight. Find out composition of mixture.

Ans. FeO = 20.28%,

Fe_3O_4 = 79.72%

Sol. Let weight of FeO and Fe_3O_4 be a and b g, respectively.

$$2FeO + \frac{1}{2}O_2 \longrightarrow Fe_2O_3$$

$$2Fe_3O_4 + \frac{1}{2}O_2 \longrightarrow 3Fe_2O_3$$

$\because$ 144 g FeO gives 160 g Fe_2O_3

$\therefore$ a g FeO gives 160 × a/144 g Fe_2O_3

Similarly, weight of Fe_2O_3 formed by $Fe_3O_4 = \frac{160 \times 3b}{464}$

Now if a + b = 100 ; then $\frac{160 \times a}{144} + \frac{160 \times 3b}{464} = 105$

Solving these two equations : a = 21.06 and b = 78.94

Percentage of **FeO ≈ 20.28%** and percentage of **Fe_3O_4 ≈ 79.72%**

Q.20 The reaction, $2C + O_2 \longrightarrow 2CO$ is carried out by taking 24 g of carbon and 96g O_2, find out:
(A) Which reactant is left in excess?
(B) How much of it is left
(C) How many mole of CO are formed?
(D) How many g of carbon should be taken so that nothing is left at the end of the reaction?

Sol.

	2C	+ $O_2 \longrightarrow$	2CO
Mole before reaction	$\frac{24}{12} = 2$	$\frac{96}{32} = 3$	0
Mole after reaction	0	2	2

(A) $\therefore$ O_2 is left in excess.
(B) 2 mole of O_2 or 64 g O_2 is left.
(C) 2 mole of CO or 56 g CO is formed.
(D) To use O_2 completely total 6 mole of carbon or 72 g carbon is needed.

Q.21 8g of sulphur are burnt to from SO_2, which is oxidised by Cl_2 water. The solution is treated with $BaCl_2$ solution. The amount of $BaSO_4$ precipitated is :
(A) 1.0 mole (B) 0.5 mole (C) 0.75 mole (D) 0.25 mole

Ans. (D)

Sol. $$S \xrightarrow{O_2} SO_2 \xrightarrow{Cl_2} SO_4^{2-} \xrightarrow{BaCl_2} BaSO_4$$

Mole of $BaSO_4$ formed = moles of sulphur $= \frac{8}{32} = \frac{1}{4}$

Q.22 8 litre of H_2 and 6 litre of Cl_2 are allowed to react to maximum possible extent. Find out the final volume of reaction mixture. Suppose P and T remains constant throughout the course of reaction-
(A) 7 litre (B) 14 litre (C) 2 litre (D) None of these.

Ans. (B)

Sol.

	H_2	+ Cl_2	$\rightarrow$ 2 HCl
Volume before reaction	8 lit	6 lit	0
Volume after reaction	2	0	12

$\therefore$ Volume after reaction = Volume of H_2 left + Volume of HCl formed
= 2 + 12 = **14 lit**

Q.23 When measured at the same temperature and pressure, hydrogen reacts with oxygen to form water in the volume ratio 2 : 1. Calculate the volume of O_2 gas measured at 137°C and 760 mm pressure that will combine with 100 ml of H_2 at 0°C and 200 mm pressure.

Ans. 19.76 ml

Sol. Reaction : $\underset{2Vol}{2H_2} + \underset{1Vol}{O_2} \rightarrow 2H_2O$

2 volume of H_2 required O_2 = 1 Vol

$\therefore$ 100 mL of H_2 required $O_2 = \frac{100}{2} = 50$ ml at 0°C and 200 mm

$\therefore P_1 = 200$ mm, $V_1 = 50$ mL, $T_1 = 0 + 273 = 273$ K, $P_2 = 760$ mm, $V_2 = ?$, $T_2 = 137 + 273 = 410$ K

But, $$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \text{ (gas equation)}$$

Hence, $V_2 = \frac{P_1V_1T_2}{P_2T_1}$

$\therefore$ $V_2 = \frac{200mm \times 50mL \times 410K}{760mm \times 273K}$ **= 19.76 ml Ans.**

Q.24 A gaseous mixture containing 49.5 ml of hydrogen chloride and H_2 gases was kept in contact with Na/Hg. The volume of mixture decreased to 42 ml. If 99.5 ml of the same mixture is mixed 49.5 ml of gaseous ammonia and then exposed to water, calculate the final volume. All measurements of volume being done under same conditions of temperature and pressure.

Ans. 60.5 ml

Sol. Volume of H_2 gas + HCl (g) = 49.5 mL. When kept in contact with Na/Hg only HCl will react with Na as follows :

$$\underset{\ }{2Na} + \underset{2Vol}{HCl} \longrightarrow 2NaCl + \underset{1Vol}{H_2}$$

Reduction in volume = 2 Vol – 1 Vol = 1 Vol for 2 Vol of HCl

Actual reduction in volume = 49.5 – 42 = 7.5 ml

$\therefore$ Volume of HCl present in 49.5 ml mixture = 7.5 × 2 = 15 ml

Volume of H_2 = 49.5 – 15 = 34.5 ml

$\therefore$ 99.5 mL of mixture would contain HCl(g) = $15 \times \frac{99.5}{49.5}$ = 30 ml

and H_2 = 99.5 – 30 = 69.5 ml

When mixed with NH_3, the reaction will be :

$$\underset{\substack{1Vol \\ 30\,ml}}{NH_3(g)} + \underset{\substack{1Vol \\ 30\,ml}}{HCl(g)} \longrightarrow NH_4Cl(s)$$

$\therefore$ Residual NH_3 = 49.5 – 30 = 19.5 ml

It will also dissolve in H_2O. So, only gas left behind is H_2 gas. Hence :

Residual H_2 = 99.5 – 30 = 69.5 mL or final volume **= 69.5 ml**

Q.25 A mixture of ethyne (C_2H_2) and H_2 was mixed with 65 ml of O_2 and exploded in a eudiometer tube. On cooling, it was found to have undergone a contraction of 35 ml. When treated with a KOH solution, a further contraction of 34 ml took place and 15 ml of O_2 alone was left behind. Find the percentage composition of ethyne and H_2 in the mixture.

Ans. C_2H_2 = 53.2 %, H_2 = 46.8%

Sol. Let mixture contains H_2 = x ml and C_2H_2 = y ml

Reaction (i) $\underset{2\,Vol}{2H_2} + \underset{1\,Vol}{O_2} \longrightarrow \underset{2\,Vol}{2H_2O}$

or x ml $\quad \frac{x}{2}$ ml

(ii) $\underset{2\,Vol}{2C_2H_2} + \underset{5\,Vol}{5O_2} \longrightarrow \underset{4\,Vol}{4CO_2} + \underset{2\,Vol}{2H_2O}$

or y ml $\quad \frac{5y}{2}$ ml $\quad$ 2y ml

From question

$2y = 34 \Rightarrow y = 17$

and , $\frac{x}{2} + \frac{5y}{2} = 65 - 17$

$\therefore$ $x = 15$

$\therefore$ $x = 2 \times 7.5 = 15.0$ ml

Thus, volume of H_2 = 15.0 mL; volume of C_2H_2 = 17 ml

$$\therefore \quad \%\text{ age of } H_2 = \frac{\text{Vol. of } H_2}{\text{Vol. of } H_2 + \text{Vol. of } C_2H_2} \times 100$$

$$= \frac{15}{15+17} \times 100 = 46.8$$

$$\therefore \quad \%\text{ age of } C_2H_2 = 100 - 46.88 = 53.2$$

Q.26 0.9 g of a solid organic compound (molecular weight 90) containing carbon, hydrocarbon and oxygen, was heated with oxygen corresponding to a volume of 224 ml at S.T.P. After combustion, the total volume of the gases was 560 ml at S.T.P. On treatment with potassium hydroxide, the volume decreased to 112 ml. Determine the molecular formula of the compound.

Ans. $C_2H_2O_4$

Sol. Let the formula of organic compound containing C, H and oxygen = $C_xH_yO_z$,
On combustion, CO_2 and H_2O (zero volume at 0°C) are formed
So, the residual gas will contain CO_2 and unused O_2,
Since CO_2 gas is absorbed by KOH, the volume of CO_2 formed = 560 – 12 = 448 ml

Volume of O_2 used = 224 – 112 = 112 ml

$$\text{Mole of } CO_2 \equiv \frac{1}{22400} \times 448$$

$$\text{No. of mole of compound} = \frac{0.9}{90} = 0.01 \text{ mol}$$

$\therefore$ 0.01 mole compound produced = 0.02 mole of CO_2

$$1 \text{ mole compound produced } CO_2 = 0.2 \times \frac{1}{0.01} = 2\text{mol } CO_2$$

$\therefore$ 2 mol $CO_2 \equiv$ 2 mol C atoms

$\therefore$ Formula of compound = $C_2H_yO_z$

Mol. wt. of $C_2H_yO_z = (2 \times 12) + 1 \times y + 16 \times z = 24 + y + 16z$

$$\text{Reaction } C_2H_yO_z + \frac{1}{2}O_2 \rightarrow 2CO_2 + \frac{y}{2} H_2O$$

Atoms of oxygen involved in the above reaction are :

$$z + 1 = 4 + \frac{y}{2};\ z = 4 + \frac{y}{2} - 1 = 3 + \frac{y}{2}$$

$$\therefore \quad z = 3 + \frac{y}{2}$$

But mol. wt. $24 + y + 16z = 90$

Hence $24 + y + 16\left(3+\frac{y}{2}\right) = 90$

$$y = 2$$

Substituting the value of y = 2

$$z = 3 + \frac{2}{2} = 4$$

$\therefore$ Molecular formula of compound = $C_2H_yO_z = C_2H_2O_4$ Ans.

Q.27 At high temperatures, the compound S_4N_4 decomposes completely into N_2 and sulphur vapours. If all measurements are made under the same conditions of temperature and pressure, it is found that for each volume of S_4N_4 decomposed 2.5 volumes of gaseous products are formed. What is the molecular formula of sulphur ?

Ans. S_8

Sol. Reaction

$$\begin{array}{lcl} N_4S_4 & \xrightarrow{\text{High temp.}} & 2N_2 + \text{sulphur vapours} \\ 1 \text{ Vol} & & 2.5 \text{ Vol } (2 + 0.5) \\ 1 \text{ molecule} & & 2.5 \text{ molecules} \end{array}$$

2.5 molecules of N_2 + sulphur vapours = 4 atoms of nitrogen + 0.5 molecule of sulphur vapour. Hence the formula S_4N_4 indicates that :

0.5 molecule of sulphur contains S-atoms = 4

1 molecule of sulphur contains S-atoms = $\frac{4}{0.5} = 8$

$\therefore$ Formula of sulphur = S_8

Q.28 10 ml of a gaseous organic compound containing C, H and O only was mixed with 100 ml of oxygen and exploded under conditions which allowed the water formed to condense. The volume of the gas after explosion was 90 ml. On treatment with potash solution, a further contraction in volume of 20 ml was observed. Given that the vapour density of the compound is 23, deduce the molecular formula. All volume measurements were carried out under the same conditions.

Ans. C_2H_6O

Sol. Let formula of organic compound = $C_xH_yO_z$

$$\begin{array}{lcccc} \text{Reaction } C_xH_yO_z & + \left(x+\frac{y}{4}-\frac{z}{2}\right) O_2 & \longrightarrow & xCO_2 & + \frac{y}{2}H_2O \\ 10 \text{ ml} & 10\left(x+\frac{y}{4}-\frac{z}{2}\right) & & 10\,x \text{ ml} & \frac{10\times y}{2} = 5y \text{ ml} \end{array}$$

(a) Since potash solution (KOH) absorbs CO_2, so volume of CO_2 gas = 20 ml.

$$10\,x = 20;\ x = 2$$

(b) Contraction in volume after explosion and cooling as obtained from reaction

$$10 + 10\left(x+\frac{y}{4}-\frac{z}{2}\right) - 10\,x = 10 + 100 - 90 = 20 \text{ ml (given)}$$

$$10 + 10\left(2+\frac{y}{4}-\frac{z}{2}\right) - (10 \times 2) = 20$$

$$5y - 10\,z = 20 \qquad \text{........(1)}$$

But mol. wt. of $C_xH_yO_z$ or $C_2H_yO_z = (2 \times 12) + (y \times 1) + (16 \times z)$
$= 24 + y + 16z = 2 \times 23$ (given)

or $y + 16z = 46 - 24 = 22$ ....(2)

From (1) and (2) $z = 1$
$y = 6$

Since $x = 2$, $y = 6$ and $z = 1$, the molecular formula of $C_xH_yO_z = C_2H_6O$

Q.29 A mixture of formic acid and oxalic acid is heated with concentrated H_2SO_4. The gas produced is collected and on its treatment with KOH solution the volume of the gas decreased by one-sixth. Calculate the molar ratio of the two acids in the original mixture.

Ans. 4 : 1

Sol. Let moles of HCOOH and $H_2C_2O_4$ are a and b respectively

The decomposition of the acid takes place as follows

HCOOH	$\rightarrow$	H_2O	+	CO		
a moles		a moles		a moles		
$H_2C_2O_4$	$\rightarrow$	H_2O +		CO	+	CO_2
b moles		b moles		b moles		b moles

H_2O is absorbed by H_2SO_4 and CO_2 is absorbed by KOH.

Thus as given, $\dfrac{\text{volume of } CO_2}{\text{total volume of } (CO + CO_2)} = \dfrac{b}{a+b+b} = \dfrac{1}{6}$

or $\dfrac{a}{b} = 4$

Molar ratio of HCOOH and $H_2C_2O = 4 : 1$

Q.30 9 volumes of a gaseous organic compound A and just sufficient amount of oxygen required for its complete combustion, yielded, 4 volume of CO_2, 6 volumes of water vapour and 2 volumes of N_2, all volumes measured at the same temperature and pressure. If the compound A contained only C, H and N, (i) how many volumes of oxygen are required for complete combustion, and (ii) what is the molecular formula of the compound A?

Ans. (i) 7 volume (ii) $C_2H_6N_2$

Sol. Suppose the compound A is $C_xH_yN_z$

(i)	$C_xH_yN_z$	+	O_2	$\rightarrow$	CO_2 +	H_2O (vapour) +	N_2
or	$(9 - v)$ moles		v moles		4 moles	6 moles	2 moles

Applying POAC for O atoms,

$2 \times$ moles of $O_2 = 2 \times$ moles of $CO_2 + 1 \times$ moles of H_2O

$2v = 2 \times 4 + 1 \times 6 = 14$; $v = 7$ volumes

(ii) The reaction is

	$C_xH_yN_z$	+	O_2	$\rightarrow$	CO_2 +	H_2O (vapour) +	N_2
	2 vols.		7 vols.		4 vols.	6 vols.	2 vols.
or	2 moles		6 moles		4 moles	6 moles	2 moles

Applying POAC for C, H and N, we get respectively

$x \times$ moles of $C_xH_yN_z = 1 \times$ moles of CO_2

$x \times 2 = 1 \times 4; x = 2$

$y \times$ moles of $C_xH_yN_z = 2 \times$ moles of H_2O (vapour)

$y \times 2 = 2 \times 6; y = 6$

$z \times$ moles of $C_xH_yN_z = 2 \times$ moles of N_2

$z \times 2 = 2 \times 2; z = 2$

Hence the compound is $C_2H_6N_2$

Q.31 15 ml of gaseous hydrocarbon required for complete combustion 357 ml of air (21% of O_2 by volume) and the gaseous products occupied 327 ml (all volumes being measured at S.T.P.). What is the formula of hydrocarbon ?

Ans. C_3H_8

Sol. Let C_xH_y be the hydrocarbon

$$C_xH_y \quad + \left(x+\frac{y}{4}\right) O_2 \rightarrow \quad xCO_2 + \frac{y}{2} H_2O$$

$$1\text{ Vol.} \qquad \left(x+\frac{y}{4}\right)\text{Vol.} \qquad x\text{ Vol.}$$

From equation, the contraction

$= 1 + (x + y/4) - (x + 0) = 1 + y/4$

for 15 ml gas, contraction =

$15 (1 + y/4) = (15 + 357) - (327) = 45$

$y = 8$

The gaseous products after contraction = 327

This includes vol. of CO_2 plus volume of N_2 in the air (O_2 is completely used up). So calculate the volume of N_2 in the air.

Vol. of $O_2 = 0.21 \times 357 = 75$ ml

Vol. of $N_2 = 357 - 75 = 282$ ml

Now, Vol. of N_2 + Vol. of $CO_2 = 327$ ml

Vol. of $CO_2 = 327 - 282 = 45$ ml

The volume of CO_2 produced $= 15\,x$

$15\,x = 45$

$x = 3$

Hence the hydrocarbon is C_3H_8

Q.32 Calculate the molarity and molality of a solution of H_2SO_4 (sp. gr. = 1.98) containing 27% H_2SO_4 by mass.

Ans. 3.3 M, 3.77 M

Sol. Vol of 100 g of 27% $H_2SO_4 = \frac{\text{wt.}}{\text{d}} = \frac{100}{1.198}$ ml

$$M_{H_2SO_4} = \frac{\text{wt./mol.wt.}}{\text{vol. of solution (litre)}} = \frac{27 \times 1.198 \times 1000}{98 \times 100} = \mathbf{3.3\ mol}\ L^{-1}$$

$$m_{H_2SO_4} = \frac{\text{wt./mol.wt.}}{\text{wt. of solvent(kg)}} = \frac{27 \times 1000}{(100-27) \times 98} = \mathbf{3.77}\text{ mol } Kg^{-1}$$

Q.33 How many milliliter of concentrated sulphuric acid of density 1.84 g ml^{-1} containing 98% H_2SO_4 by mass are required to make (a) 4 litre of 1 N solution and (b) 200 mL of 0.25 M solution.

Ans. (a) 108.7 (b) 2.72 ml

Sol. (a) Gram eq. wt. of $H_2SO_4 = \frac{98}{2} = 49$ g.

To find wt. of H_2SO_4 (W) required to prepare 4 L of 1N H_2SO_4, we have :

$$1 = \frac{\text{wt. of } H_2SO_4}{\text{g. eq. wt. of } H_2SO_4 \times 4(l)} = \frac{\text{wt. of } H_2SO_4}{49 \times 4(l)}$$

Or wt. of $H_2SO_4 = 196$ g.

Since acid is 98% thus amount of solution $= \frac{196 \times 100}{98} = 200$ g.

Thus, volume (mass/density) of 98% H_2SO_4 solution $= \frac{200}{1.84}$ **= 108.7 ml**

(b) $0.25 = \frac{\text{wt.of } H_2SO_4 \times 1000}{\text{g. mol. wt. of } H_2SO_4 \times \text{vol. of solution in ml}}$

$\therefore$ wt. of $H_2SO_4 = \frac{0.25 \times 98 \times 200}{1000} = 4.9$ g

$\therefore$ wt. of 98% H_2SO_4 solution $= \frac{4.9 \times 100}{98} = 5$ g

Volume of H_2SO_4 solution $= \frac{\text{Mass}}{\text{Density}} = \frac{5}{1.84}$ **= 2.72 ml**

Q.34 Derive the relation between molality (m) and mole fraction of solute, χ_2

Sol. Molality, m means, m mole of solute in 1000 g of solvent which is equal to $1000/M_1$ mol where M_1 = molar mass of the solvent.

$$\therefore \quad \text{Mole fraction, } \chi_2 = \frac{\text{moles of solute}}{\text{Moles of solute} + \text{Moles of solvent}} = \frac{m}{m + \frac{1000}{M_1}} = \frac{mM_1}{mM_1 + 1000}$$

Hence $$m = \frac{1000 \times \chi_2}{(1 - X_2)}$$

Q.35 The molality and molarity of a solution of H_2SO_4 are 94.13 and 11.12 respectively. Calculate the density of the solution.

Ans. 1.2079 g/ml

Sol. $d = M\left(\frac{1}{m} + \frac{\text{mol.wt.}}{1000}\right) = 11.12\left(\frac{1}{94.13} + \frac{98}{1000}\right)$ **= 1.2079 g/ml**

Q.36 Calculate the (a) molarity (b) normality of the phosphoric acid solution (sp. gravity 1.426 and containing 60% by weight of pure H_3PO_4. Atomic mass of P = 31)

Ans. (a) 8.73 M (b) 26.19 N

Sol. Weight of H_3PO_4 in 100 g solution = 60 g; Weight of water = 100 – 60 = 40 g

(a) Calculation of molarity

Molecular mass of H_3PO_4 = 98 g mol^{-1}

$$\text{Molarity} = \frac{60}{98} \times \frac{1.426 \times 1000}{100} = \mathbf{8.73\ M}$$

(b) Normality = Molarity × Basicity = 8.73 × 3 = **26.19 N**

Q.37 Upon heating a litre of $\frac{M}{2}$ HCl solution, 2.675 g hydrogen chloride is lost due to evaporation and the volume of the solution shrinks to 750 ml. Calculate

(i) the molarity of the resulting solution and

(ii) the number of milli moles of HCl in 100 ml of the final solution.

Ans. (i) 0.569 M (ii) 56.9

Sol. Mol. wt. of HCl = 36.5

$$\text{Mass of HCl in 1L of } \frac{M}{2} \text{ HCl} = \frac{1}{2} \times 36.5 = 18.25 \text{ g}$$

Weight of HCl lost on heating = 2.675 g

Weight of HCl left in solution = 18.25 – 2.675 = 15.575 g

$$\text{Number of moles in 15.575 g HCl} = \frac{15.575}{36.5} = 0.4267$$

$$\text{Volume of the solution left after heating} = 750 \text{ ml} = \frac{750}{1000} = 0.75 \text{ L}$$

$$\therefore \quad \text{Molarity of solution} = \frac{\text{No. of gram moles}}{\text{Volume in litres}} = \frac{0.4267}{0.75} = 0.569 \text{ M}$$

Now number of milli moles = (volume in ml) × (Molarity)

= 100 × 0.569 = **56.9**

Q.38 Calculate the normality of a solution obtained by mixing 0.01 L of $\frac{N}{10}$ NaOH and 40 ml of seminormal KOH solution.

Ans. 0.42 N

Sol. $$\text{Total volume (V) of mixed solution} = \left(0.01\text{L} \times \frac{1000\text{mL}}{1\text{L}}\right) + 40 \text{ ml} = 50 \text{ m}l$$

Now $N_1V_1 + N_2V_2 = NV$

$$\frac{1}{10} \times 10 + \frac{1}{2} \times 40 = N \times 50$$

$$\therefore \quad N = \frac{21}{50} = \mathbf{0.42\ \ geq}\ \boldsymbol{L^{-1}}$$

Q.39 30 ml Na_2CO_3 solution is mixed in 20 ml 0.8 N H_2SO_4. The resulting solution required 20 ml 0.7 N HCl for neutralization. Calculate the strength of Na_2CO_3 solution in g/L, Take Na_2CO_3 as anhydrous.

Ans. 53 g/L

Sol. meq of acids = $(20 \times 0.8) + (20 \times 0.7) = 16 + 14 = 30$

meq of $Na_2CO_3 = 30$

$$\text{Normality of } Na_2CO_3 = \frac{30}{30} = 1N$$

$$\text{Strength} = N \times \text{eq. wt} = \left(1 \times \frac{106}{2}\right) = \mathbf{53\ g/L}$$

Q.40 1 g sample of H_2O_2 solution containing x % of H_2O_2 by weight requires x ml $KMnO_4$ solution for complete oxidation under acidic conditions. Calculate the normality of $KMnO_4$ solution.

Ans. 0.588 N

Sol. 1 g of solution contains $\frac{1 \times x}{100} = 0.01\,x$ g H_2O_2

17 g (eq. wt.) of H_2O_2 is present in 1000 ml 1N solution

$\therefore$ 0.01 x g (eq. wt.) of H_2O_2 is present in $\frac{1000 \times 0.01x}{17}$ ml 1N solution

$$\text{Now, } \frac{1000 \times 0.01x}{17} \times N = x \times N_{KMnO_4}$$

$$N_{KMnO_4} = \frac{1000 \times 0.01x \times N}{17 \times x} = \mathbf{0.588\ N}$$