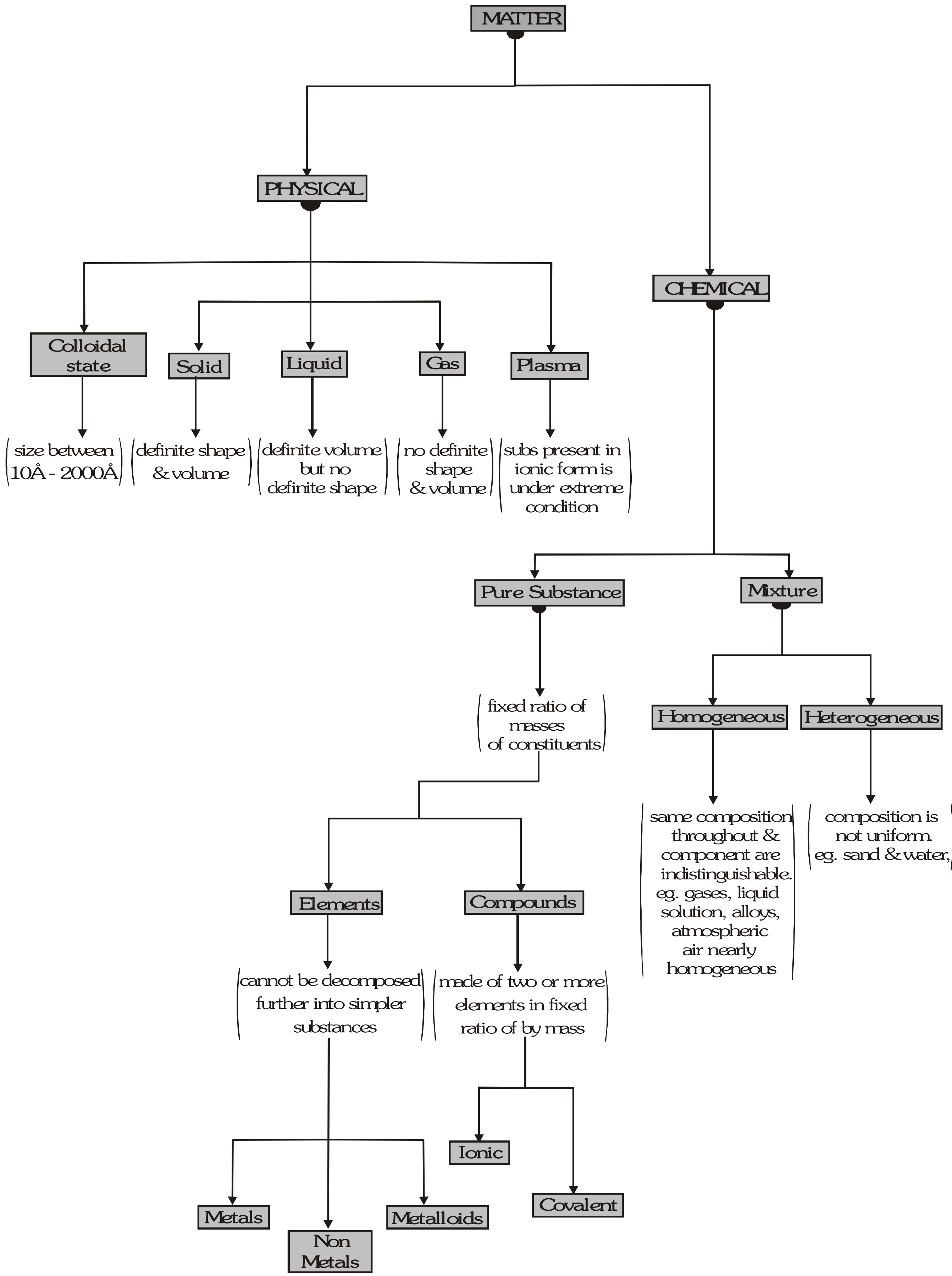
MOLE CONCEPT



1. CLASSIFICATION OF UNIVERSE :

(1) Matter

(2) Energy

(1) MATTER

The thing which occupy space and having mass which is feel by our five senses is called as matter.

◆ **2 types**

(I) Physical classification

(II) Chemical classification

(I) **Physical Classification :**

It is based on physical state under ordinary conditions of temperature and pressure, matter is classified into the following three types :

(a) Solid

(b) Liquid

(c) Gas

(a) **Solid :**

A substance is said to be solid if it possesses a definite volume and a definite shape

e.g. sugar, iron, gold, wood etc.

(b) **Liquid :**

A substance is said to be liquid if it possesses a definite volume but not definite shape. They take up the shape of the vessel in which they are put.

e.g. water, milk, oil, mercury, alcohol etc.

(c) **Gas :**

A substance is said to be gas if it neither possesses a definite volume nor a definite shape. This is because they fill up the whole vessel in which they are put.

e.g. hydrogen(H_2), oxygen(O_2), carbon dioxide(CO_2), etc.'

(II) **Chemical Classification :**

◆ **2 Types**

(A) Pure Substance

(B) Mixture

(A) **Pure Substance :**

A material containing only one type of substance. Pure Substance can not be seperated into simpler substance by physical method.

e.g. : Element = Na, Mg, Ca .... etc.

Compound = HCl, H_2O, CO_2, HNO_3 .... etc.

◆ **2 types**

(a) Element

(b) Compound

(a) **Element** : The pure substance containing only one kind of atoms .

3 types (depend on physical and chemical property)

(a') Metal

(b') Non-metal

(c') Metalloids

(b) Compound :

It is defined as pure substance containing more than one kind of atoms which are combined together in a fixed ratio by weight and which can be decomposed into simpler substance by the suitable chemical method. The properties of a compound are differnt from those of its components.

e.g. : $\boxed{H_2O}$, HCl, HNO_3 .... etc.

2 : 16

$\boxed{1 : 8}$ **by wt.**

- **2 types**
  - (a') Organic Compound
  - (b') Inorganic Compound

(B) Mixture :

A material which contain more than one type of substances and which is mixed any ratio by wt.is called as mixture.

- The property of the mixture is the property of its components
- The mixture is seperated by simple physical method.
- **2 types**

(a') Homogeneous mixture

(b') Hetrogeneous mixture

(a') Homogeneous mixture :

The mixture, in which all the components are present in **uniform** is called as homogeneous mixture.

e.g. : Water + Salt, Water + Sugar, Water + alcohol,

(b') Hetrogeneous mixture :

The mixture in which all the components are present in **nonuniform** is called as Hetrogeneous mixture.

e.g. : Water + Sand, Water + Oil,

❑ INTRODUCTION :

There are a large number of objects around us which we can see and feel.

Anything that occupies space and has mass is called matter.

It was **John Dalton** who firstly developed a theory on the structure of matter, later on which is known as **Dalton's atomic theory.**

1. DALTON'S ATOMIC THEORY :

1. Matter is made up of very small undivisible particle called atoms.
2. All the atoms of a given element is identical in all respect i.e. mass, shape, size, etc.
3. Atoms cannot be created nor destroyed by any chemical process.
4. Atoms of different elements are different in nature.

2. THE LAW OF CHEMICAL COMBINATION :

- **Atoine Lavoisier, John Dalton** and other scientists formulate certain law concerning the composition of matter and chemical reactions. These laws are known as the law of chemical combination.

3. THE LAW OF CONSERVATION OF MASS :

It is given by **Lavoisier.**

In a chemical change total mass remains conserved.

i.e. mass before reaction is always equal to mass after reaction.

◆ **Example** :

	H_2 (g) +	$\frac{1}{2}O_2$ (g) ⟶	H_2O (ℓ)
Before reaction initialy	1 mole	$\frac{1}{2}$ mole	
After the reaction	0	0	1 mole

mass before reaction = mass of 1 mole H_2 (g) + $\frac{1}{2}$ mole O_2 (g)

= 2 + 16 = 18 g

mass after reaction = mass of 1 mole water = 18 g

4. LAW OF CONSTANT OR DEFINITE PROPORTION :

It is given by **Proust.**

All chemical compounds are found to have constant composition irrespective of their method of preparation or sources.

◆ **Example** :

In water (H_2O), Hydrogen and Oxygen combine in 2 : 1 molar ratio, the ratio remains constant whether it is tap water, river water or sea water or produced by any chemical reaction.

Ex. 1.80 g of a certain metal burnt in oxygen gave 3.0 g of its oxide. 1.50 g of the same metal heated in steam gave 2.50 g of its oxide. Show that these results illustrate the law of constant proportion.

Sol. In the first sample of the oxide,

wt. of metal = 1.80 g,

wt. of oxygen = (3.0 – 1.80) g = 1.2 g

$$\therefore \quad \frac{\text{wt. of metal}}{\text{wt. of oxygen}} = \frac{1.80\text{g}}{1.2\text{g}} = 1.5$$

In the second sample of the oxide,

wt. of metal = 1.50 g,

wt. of oxygen = (2.50 – 1.50) g = 1 g

$$\therefore \quad \frac{\text{wt. of metal}}{\text{wt. of oxygen}} = \frac{1.50\text{g}}{1\text{g}} = 1.5$$

Thus, in both samples of the oxide the proportions of the weights of the metal and oxygen are fixed. Hence, the results follows the law of constant proportion.

5. THE LAW OF MULTIPLE PROPORTION :

It is given by **Dalton.**

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

Note : Simple ratio here means the ratio between small natural numbers, such as 1 : 1, 1 : 2, 1 : 3, Later on this simple ratio becomes the valency and then oxidation state of the element.

Example : Carbon and Oxygen when combine, can form two oxides viz CO (carbonmonoxide), CO_2 (Carbondioxides)

In CO, 12 g carbon combined with 16 g of oxygen.

In CO_2, 12 g carbon combined with 32 g of oxygen.

Thus, we can see the mass of oxygen which combine with a constant mass of carbon (12 g) bear simple ratio of 16 : 32 or 1 : 2.

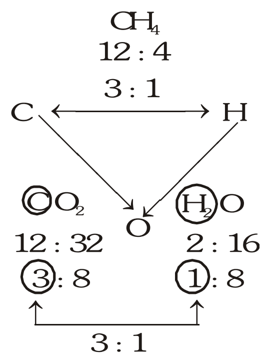
6. LAW OF RECIPROCAL PROPORTION (OR LAW OF EQUIVALENT WT.) :

- It is given by **Richter**.

Statement :

The ratio of the weights of two elements A and B which combine seperately with a fixed weight of the third element C is either the **same or simple ratio** of the weights in which A and B combine directly with each other.

e.g.



- **Special Note :** This law is also called as law of equivalent wt. due to each element combined in their equivalent wt. ratio.

$$E = \frac{M_w/\text{At.wt.}}{\text{V.F.}}$$

- **For ions**

 V.F. = Total no. of positive charge

 or V.F. = Total no. of negative charge

❑ EXAMPLE BASED ON LAW OF RECIPROCAL PROPORTION

Ex.1 Ammonia contains 82.35% of nitrogen and 17.65% of hydrogen. Water contains 88.90% of oxygen and 11.10% of hydrogen. Nitrogen trioxide contains 63.15% of oxygen and 36.85% of nitrogen. Show that these data illustrate the law of reciprocal proportions.

Sol. In NH_3, 17.65g of H combine with N = 82.35g

$\therefore$ 1 g of H combine with $N = \frac{82.35}{17.65}$ g $= 4.67$ g

In H_2O, 11.10 g of H combine with O = 88.90 g

$\therefore$ 1 g of H combine with $O = \frac{88.90}{11.10}$ g $= 8.01$ g

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

$= 4.67 : 8.01 = 1 : 1.7$

In N_2O_3, ratio of weights of N and O which combine with each other $= 36.85 : 63.15$

$= 1 : 1.7$

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

- Other Method

N

N_2		O_3
36.85	:	63.15
$\frac{36.85}{36.85}$	:	$\frac{63.15}{36.85}$
1	:	1.7

N		H_3
82.53	:	17.65
$\frac{82.53}{17.65}$	:	$\frac{17.65}{17.65}$
4.67	:	1

O

H_2		O
11.1	:	88.9
$\frac{11.1}{11.1}$	:	$\frac{88.9}{11.}$
1	:	8.01

N		O
4.67	:	8.01
$\frac{4.67}{4.67}$	:	$\frac{8.01}{4.67}$
1	:	1.7

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

7. RELATIVE ATOMIC MASS :

One of the most important concept come out from Dalton's atomic theory was that of relative atomic mass or relative atomic weight. This is done by expressing mass of one atom with respect to a fixed standard. Dalton used hydrogen as the standard (H = 1). Later on oxygen (O = 16) replaced hydrogen as the reference. Therefore relative atomic mass is given as

$$\text{Relative atomic mass (R.A.M.)} = \frac{\text{Mass of one atom of an element}}{\text{mass of one hydrogen atoms}} = \frac{\text{Mass of one atom of an element}}{\frac{1}{16}\times\text{mass of one Oxygen atom}}$$

- The present standard unit which was adopted internationally in 1961, is based on the mass of one carbon-12 atom.

8. ATOMIC MASS UNIT (OR AMU) :

The atomic mass unit (amu) is equal to one twelvth $\left(\frac{1}{12}\right)$ the mass of one atom of carbon-12 isotope.

$\therefore$ $\boxed{1\ \text{amu} = \frac{1}{12}\quad \text{mass of one C-12 atom}}$ $= 1.66 \quad 10^{-24}$ g or $1.66 \quad 10^{-27}$ kg

- **One amu is also called one Dalton (Da).**

Now the relative atomic mass is given as

$$\text{Relative atomic mass} = \frac{\text{mass of one atom of the element}}{\frac{1}{12}\times\text{mass of one C}-12\text{ atom}}$$

$$\text{R.A.M.} = \frac{\text{Atomic mass}}{1\ \text{amu}}$$

Atomic mass = R.A.M.   1 amu

$$\text{Relative molecular mass} = \frac{\text{mass of one molecule of the substance}}{\frac{1}{12}\times\text{mass of one C}-12\text{ atom}}$$

$\therefore$ Molecular mass = Relative molecular mass   1 amu

9. MOLE :

A mole is the amount of a substance that contains as many entities (atoms, molecules or other particles) as there are atoms exactly in 0.012 kg (or 12 g) of the carbon - 12 isotope.

From mass spectrometer we found that there are $6.023 \quad 10^{23}$ atoms are present in 12 g of C-12 isotope.

The number of entities in 1 mol is so important that it is given a separate name and symbol known as Avogadro constant denoted by N_A.

i.e. on the whole we can say that 1 mole is the collection of $6.02 \quad 10^{23}$ entities. Here entities may represent atoms, ions, molecules or even pens, chair, paper etc.

1 mole of atom is also termed as 1 g – atom

1 mole of ions is also termed as 1 g – ion

1 mole of molecule is also termed as 1 g – molecule

Methods of Calculations of mole :

(a) If no. of some species is given, then no. of moles $= \dfrac{\text{Given no.}}{N_A}$

(b) If weight of a given species is given, then no. of moles $= \dfrac{\text{Given wt.}}{\text{Atomic wt.}}$ (for atoms),

or $= \dfrac{\text{Given wt.}}{\text{Molecular wt.}}$ (for molecules)

(c) If volume of a gas is given along with its temperature (T) and pressure (P).

use $n = \dfrac{PV}{RT}$

where R = 0.0821 lit-atm/mol-K (when P is in atmosphere and V is in litre)

1 mole of any gas at STP occupies 22.4 litre.

Ex. Chlorophyll the green colouring material of plants contains 3.68 % of magnesium by mass. Calculate the number of magnesium atom in 5.00 g of the complex.

Sol. Mass of magnesium in 5.0 g of complex $= \dfrac{3.68}{100} \quad 5.00 = 0.184$ g

Atomic mass of magnesium = 24

24 g of magnesium contain = $6.023 \quad 10^{23}$ atoms

$\therefore$ 0.184 g of magnesium would contain $= \dfrac{6.023 \times 10^{23}}{24} \quad 0.184 = 4.617 \quad 10^{21}$ atom

Therefore, 5.00 g of the given complex would contain $4.617 \quad 10^{21}$ atoms of magnesium.

10. GRAM ATOMIC MASS :

The atomic mass of an element expressed in gram is called gram atomic mass of the element.

OR

It is also defined as mass of $6.02 \quad 10^{23}$ atoms.

OR

It is also defined as the mass of one mole atoms.

For example for oxygen atom :

Atomic mass of 'O' atom = mass of one 'O' atom = 16 amu

gram atomic mass = mass of $6.02 \quad 10^{23}$ 'O' atoms

$= 16 \text{ amu} \quad 6.02 \quad 10^{23}$

$= 16 \quad 1.66 \quad 10^{-24} \text{ g} \quad 6.02 \quad 10^{23} = 16 \text{ g}$

($\because$ $1.66 \quad 10^{-24} \quad 6.02 \quad 10^{23} \simeq 1$)

◆ Now see the table given below and understand the definition given before.

Element	R.A.M. (Relative Atomic Mass)	Atomic mass (mass of one atom)	Gram Atomic mass/weight
N	14	14 amu	14 gm
He	4	4 amu	4 gm
C	12	12 amu	12 gm

Average atomic weight = Σ % of isotopes X molar mass of isotopes.

11. GRAM MOLECULAR MASS :

The molecular mass of a substance expressed in gram is called the gram-molecular mass of the substance.

OR

It is also defined as mass of 6.02 10^{23} molecules

OR

It is also defined as the mass of one mole molecules.

For example for 'O_2' molecule :

Molecular mass of 'O_2' molecule = mass of one 'O_2' molecule
= 2 mass of one 'O' atom
= 2 16 amu
= 32 amu

gram molecular mass = mass of 6.02 10^{23} 'O_2' molecules
= 32 amu 6.02 10^{23}
= 32 1.66 10^{-24}g 6.02 10^{23}
= 32 gm

Average molecule wt. $= \dfrac{\Sigma n_i M_i}{\Sigma n_i}$ where n_i = no. of moles of compound, m_i = molecular mass of compound

12. GAY-LUSSAC'S LAW OF COMBINING VOLUME :

According to him elements combine in a simple ratio of atoms, gases combine in a simple ratio of their volumes provided all measurements should be done in the same temperature and pressure

$$H_2\,(g) + Cl_2\,(g) \longrightarrow 2HCl$$

1 vol 1 vol 2 vol

13. AVOGADRO'S HYPOTHESIS :

Equal volume of all gases have equal number of molecules (not atoms) at same temperature and pressures conditions.

S.T.P. (Standard Temperature and Pressure)

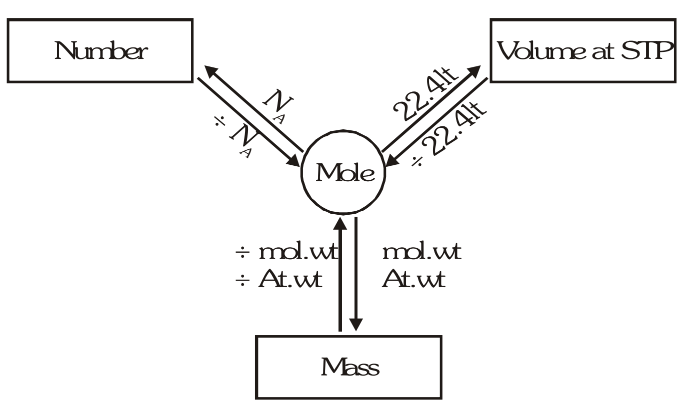
At S.T.P. condition :

temperature = 0 C or 273 K

pressure = 1 atm = 760 mm of Hg

and volume of one mole of gas at STP is found to be experimentaly equal to 22.4 litres which is known as molar volume.

14. Y-MAP :



15. PERCENTAGE COMPOSITION AND MOLECULAR FORMULA :

Here we are going to find out the percentage of each element in the compound by knowing the molecular formula of compound.

We known that according to law of definite proportions any sample of a pure compound always possess constant ratio with their combining elements.

◆ **Example :**

Every molecule of ammonia always has formula NH_3 irrespective of method of preparation or sources. i.e. 1 mole of ammonia always contains 1 mol of N and 3 mole of H. In other wards 17 g of NH_3 always contains 14 g of N and 3 g of H. Now find out % of each element in the compound.

Mass % of N in NH_3 = $\dfrac{\text{Mass of N in 1 mol } NH_3}{\text{Mass of 1 mol of } NH_3}\quad 100 = \dfrac{14\text{ g}}{17}\quad 100 = 82.35\ \%$

Mass % of H in NH_3 = $\dfrac{\text{Mass of H is 1 mol } NH_3}{\text{Mass of 1 mol of } NH_3}\quad 100 = \dfrac{3}{17}\quad 100 = 17.65\%$

16. EMPIRICAL AND MOLECULAR FORMULA :

We have just seen that knowing the molecular formula of the compound we can calculate percentage composition of the elements. Conversely if we know the percentage composition of the elements initially, we can calculate the relative number of atoms of each element in the molecules of the compound. This gives as the empirical formula of the compound. Further if the molecular mass is known then the molecular formula can be easily determined.

Thus, the empirical formula of a compound is a chemical formula showing the relative number of atoms in the simplest ratio, the molecular formula gives the actual number of atoms of each element in a molecule.

i.e. **Empirical formula :** Formula depicting constituent atom in their simplest ratio.

Molecular formula : Formula depicting actual number of atom in onemolecule of the compound.

The molecular formula is generally an integral multiple of the empirical formula.

i.e. $\boxed{\text{molecular formula = empirical formula}\quad n}$

where n = $\dfrac{\text{molecular formula mass}}{\text{empirical formula mass}}$

Ex. An organic substance containing carbon, hydrogen and oxygen gave the following percentage composition.

C = 40.687 % ; H = 5.085 % and O = 54.228 %

The molecular weight of the compound is 118. Calculate the molecular formula of the compound.

Sol. **Step-1**

To calculate the empirical formula of the compound.

Element	Symbol	Percentage of element	At. mass of element	Relative no. of atoms = $\dfrac{\text{Percentage}}{\text{At.mass}}$	Simplest atomic ratio	Simplest whole no. atomic ratio
Carbon	C	40.687	12	$\dfrac{40.687}{12}$=3.390	$\dfrac{3.390}{3.389}$=1	2
Hydrogen	H	5.085	1	$\dfrac{5.085}{1}$=5.035	$\dfrac{5.085}{3.389}$=1.5	3
Oxygen	O	54.228	16	$\dfrac{54.228}{16}$=3.389	$\dfrac{3.389}{3.389}$=1	2

∴ Empirical formula is $C_2H_3O_2$

- **Step - 2**

  To calculate the empirical formula mass.

  The empirical formula of the compound is $C_2H_3O_2$.

  $\therefore$ Empirical formula mass

  = (2 12) + (3 1) + (2 16) = 59.

- **Step - 3**

  To calculate the value of 'n'

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

- **Step - 4**

  To calculate the molecular formula of the salt

  Molecular formula = n (Empirical formula)

  = 2 $C_2H_3O_2$ = $C_4H_6O_4$

  Thus the molecular formula is $C_4H_6O_4$.

17. DENSITY :

It is of two type.

1. Absolute density 2. Relative density

For liquid and solids

$$\text{Absolute density} = \frac{\text{mass}}{\text{volume}}$$

$$\text{Relative density or specific gravity} = \frac{\text{density of the subs tan ce}}{\text{density of water at } 4°\text{C}}$$

For gases :

$$\text{Absolute density (mass / volume)} = \frac{PM}{RT}$$

where P is pressure of gas, M = mol. wt. of gas, R is the gas constant, T is the temperature.

18. RELATIVE DENSITY OR VAPOUR DENSITY :

Vapour density is defined as the density of the gas with respect to hydrogen gas at the same temperature and pressure.

$$\text{Vapour density} = \frac{d_{gas}}{d_{H_2}} = \frac{PM_{gas/RT}}{PM_{H_2/RT}}$$

$$\text{V.D.} = \frac{M_{gas}}{M_{H_2}} = \frac{M_{gas}}{2} \Rightarrow \boxed{M_{gas} = 2\,\text{V.D.}}$$

19. CHEMICAL EQUATION :

All chemical reaction are represented by chemical equations by using formule of reactant and products. Qualitatively a chemical equation simply describes what the reactants and products are. However, a balanced chemical equation gives us a lot of quantitative information mainly the molar ratio in which reactants combine and the molar ratio in which products are formed.

Example :

When potassium chlorate ($KClO_3$) is heated it gives potassium chloride (KCl) and oxygen (O_2).

$KClO_3 \xrightarrow{\Delta} KCl + O_2$ (unbalanced chemical equation)

$2KClO_3 \xrightarrow{\Delta} 2\,KCl + 3\,O_2$ (balanced chemical equation)

- Remember a balanced chemical equation is one which contains an equal number atoms of each element on both sides of equation.

20. GRAVIMETRIC ANALYSIS :

Once we get a balanced chemical equation then we can interpret a chemical equation by following ways

1. Mass - mass analysis
2. Mass - volume analysis
3. Mole - mole analysis
4. Vol - Vol analysis (separately discussed as eudiometry or gas analysis)

Now you can understand the above analysis by following example

1. Mass - mass analysis :

Consider the reaction $2\ KClO_3 \longrightarrow 2KCl + 3O_2$ According to stoichiometry of the reaction

mass-mass ratio : 2 122.5 : 2 74.5 : 3 32

or $\dfrac{\text{Mass of } KClO_3}{\text{Mass of } KCl} = \dfrac{2 \times 122.5}{2 \times 74.5}$ | $\dfrac{\text{Mass of } KClO_3}{\text{Mass of } O_2} = \dfrac{2 \times 122.5}{3 \times 32}$

Ex. Calculate the weight of iron which will be converted into its oxide by the action of 36 g of steam. (Given : $3Fe + 4H_2O \longrightarrow Fe_3O_4 + H_2$)

Sol. Mole ratio of reaction suggests,

$$\frac{\text{Mole of Fe}}{\text{Mole of } H_2O} = \frac{3}{4}$$

$$\therefore \text{ Mole of Fe} = \frac{3}{4} \quad \text{mol of } H_2O$$

$$= \frac{3}{4} \times \frac{36}{18} = \frac{3}{2}$$

$$\text{wt. of Fe} = \frac{3}{2} \quad 56 = 84 \text{ g}$$

Ex. In a gravimetric determination of P of an aqueous solution of dihydrogen phosphate in $H_2PO_4^-$ is treated with a mixture of ammonium and magnesium ions to precipitate magnesium ammonium phosphate, $Mg(NH_4)PO_4.6H_2O$. This is heated and decomposed to magnesium pyrophosphate, $Mg_2P_2O_7$. A solution of $H_2PO_4^-$ yielded 2.054 g of ($Mg_2P_2O_7$) which is weighed. What weight of NaH_2PO_4 was present originally?

Sol. $NaH_2PO_4 + Mg^{2+} + NH_4^+ \longrightarrow Mg(NH_4)PO_4.6H_2O \xrightarrow{\Delta} Mg_2P_2O_7$

As P atoms are conserved, applying POAC for P atoms, moles of P in NaH_2PO_4 = Moles of P in $Mg_2P_2O_7$

$\Rightarrow$ 1 Moles of NaH_2PO_4 = 2 Moles of $Mg_2P_2O_7$

$$\therefore \quad \frac{W_{NaH_2PO_4}}{M_{NaH_2PO_4}} = 2 \quad \frac{W_{Mg_2P_2O_7}}{M_{Mg_2P_2O_7}} \Rightarrow \frac{W_{NaH_2PO_4}}{120} = 2 \quad \frac{2.054}{222}$$

$$\therefore \quad W_{NaH_2PO_4} = 2.22 \text{ g}$$

Ex. A solid mixture weighing 5.00 g containing lead nitrate and sodium nitrate was heated below 600 C until the mass of the residue was constant. If the loss of mass is 30 %, find the mass of lead nitrate and sodium nitrate in mixture.

(At. wt. of Pb = 207, Na = 23, N = 14, O = 16)

Sol. $2Pb(NO_3)_2 \xrightarrow{\Delta} 2PbO + 4NO_2 \uparrow + O_2 \uparrow$

2 331   2 223

$2NaNO_3 \xrightarrow{\Delta} 2NaNO_2 + O_2 \uparrow$

2 85   2 69

Let, wt. of $Pb(NO_3)_2$ in mixture = x

wt. of $NaNO_3$ = (5 − x) g

662 g of $Pb(NO_3)_2$ will give residue = 446

$\therefore$ xg of $Pb(NO_3)_3$ will give residue = $\frac{446}{662} \times (x) = 0.674x$ g

170 g of $NaNO_3$ give residue = 138 g

$\therefore$ (5 − x), g $NaNO_3$ will give residue $= \frac{138}{170} \times (5 - x) = 0.812 \quad (5 - x)$

Actual wt. of residue obtained $= \left(5 - 5 \times \frac{30}{100}\right) = 3.5$ g

$\therefore$ $0.674x + 0.812 \quad (5 - x) = 3.5 \Rightarrow 0.138\ x = 0.56$

$\Rightarrow$ x = 4.05 g = wt. of $Pb(NO_3)_2$

$\therefore$ wt. of $NaNO_3$ in the mixture = (5 − 4.05) = 0.95 g

Ex. 3.0 g an impure sample of sodium sulphate dissolved in water was treated with excess of barium chloride solution when 1.74 g of **$BaSO_4$** were obtained as dry precipitate. Calculate the percentage purity of sample.

Sol. $Na_2SO_4 + BaCl_2 \longrightarrow BaSO_4 + 2NaCl$

142g   223g

223 g of $BaSO_4$ are produced from 142 g of Na_2SO_4

$\therefore$ 1.74 g of $BaSO_4$ would be produced by $= \frac{142}{233} \quad 1.74 = 1.06$ g of Na_2SO_4

% purity of $Na_2SO_4 = \frac{1.06}{3.0} \times 100 = 35.33$ %

2. Mass - volume analysis :

Now again consider decomposition of $KClO_3$

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

mass volume ratio : 2 122.5 g : 2 74.5 g : 3 22.4 L at STP

we can use two relation for volume of oxygen

$$\frac{\text{Mass of } KClO_3}{\text{volume of } O_2 \text{ at STP}} = \frac{2 \times 122.5\,g}{3 \times 22.4\,L} \qquad ....(i)$$

and $$\frac{\text{Mass of } KCl}{\text{volume of } O_2 \text{ at STP}} = \frac{2 \times 74.5\,g}{3 \times 22.4\,L} \qquad ......(ii)$$

Ex. How much marble of 90.5 % purity would be required to prepare 10 litres of CO_2 at STP when the marble is acted upon by dilute HCl ?

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

100 g   22.4litre

22.4 L of CO_2 at STP will be obtained from 100 g of $CaCO_3$

$\therefore$ 10 L of CO_2 at STP will be obtained from pure $CaCO_3 = \frac{100}{22.4} \times 10 = 44.64$ g

$\therefore$ Impure marble required $= \frac{100}{90.5} \quad 44.64 = 49.326$ g

3. **Mole - mole analysis :**

This analysis is very much important for quantitative analysis point of view.

Now consider again the decomposition of $KClO_3$.

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

In very first step of mole-mole analysis you should read the balanced chemical equation like **2 moles $KClO_3$ on decomposition gives you 2 moles KCl and 3 moles O_2** and from the stoichiometry of reaction we can write

$$\frac{\text{Moles of } KClO_3}{2} = \frac{\text{Moles of KCl}}{2} = \frac{\text{Moles of } O_2}{3}$$

Now for any general balance chemical equation like

$$a\ A + b\ B \longrightarrow c\ C + d\ D$$

you can write.

$$\frac{\text{Moles of A reacted}}{a} = \frac{\text{Moles of B reacted}}{b} = \frac{\text{Moles of C reacted}}{c} = \frac{\text{Moles of D reacted}}{d}$$

21. **PRINCIPLE OF ATOM CONSERVATION (POAC) :**

Infact POAC is nothing but the conservation of mass, expressed before in the concepts of atomic theory. And if atoms are conserved, moles of atoms shall also be conserved.

The principle is fruitful for the students when they don't get the idea of balanced chemical equation in the problem.

This principle can be under stand by the following example.

Consider the decomposition of $KClO_3$ (s) $\rightarrow$ KCl (s) + O_2 (g) (unbalanced chemical reaction)

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

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

or 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, moles of K atoms in $KClO_3$ = 1 moles of $KClO_3$

and moles of K atoms in KCl = 1 moles of KCl

$\therefore$ moles of $KClO_3$ = moles of KCl

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

- The above equation gives the mass-mass 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 $KClO_3$ = 3 moles of $KClO_3$

moles of O in O_2 = 2 moles of O_2

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

$$\text{or} \quad 3 \quad \frac{\text{wt. of } KClO_3}{\text{mol. wt. of } KClO_3} = 2 \quad \frac{\text{vol. of } O_2 \text{ at NTP}}{\text{standard molar vol. (22.4 lt)}}$$

- The above equations thus gives the mass-volume relationship of reactants and products.

Ex. 27.6 g K_2CO_3 was treated by a series of reagents so as to convert all of its carbon to $K_2Zn_3\,[Fe(CN)_6]_2$. Calculate the weight of the product.

[mol. wt. of K_2CO_3 = 138 and mol. wt. of $K_2Zn_3\,[Fe(CN)_6]_2$ = 698]

Sol. Here we have not knowledge about series of chemical reactions but we known about initial reactant and final product accordingly

$$K_2CO_3 \xrightarrow[\text{Steps}]{\text{Several}} K_2Zn_3\,[Fe(CN)_6]_2$$

Since C atoms are conserved, applying POAC for C atoms,

moles of C in K_2CO_3 = moles of C in $K_2Zn_3\,[Fe(CN)_6]_2$

1 moles of K_2CO_3 = 12 moles of $K_2Zn_3\,[Fe(CN)_6]_2$

($\because$ 1 mole of K_2CO_3 contains 1 moles of C)

$$\frac{\text{wt. of } K_2CO_3}{\text{mol. wt. of } K_2CO_3} = 12 \quad \frac{\text{wt. of the product}}{\text{mol. wt. of product}}$$

wt. of $K_2Zn_3\,[Fe(CN)_6]_2$ = $\frac{27.6}{138}$ $\frac{698}{12}$ = 11.6 g

Ex. A sample of 3 g containing Na_2CO_3 and $NaHCO_3$ loses 0.248 g when heated to 300 C, the tempera ture at which $NaHCO_3$ decomposes to Na_2CO_3, CO_2 and H_2O. What is the percentage of Na_2CO_3 in the given mixture?

Sol. The loss in weight is due to removal of CO_2 and H_2O which escape out on heating.

wt. of Na_2CO_3 in the product = 3.00 – 0.248 = 2.752 g

Let wt. of Na_2CO_3 in the mixture be x g

$\therefore$ wt. of $NaHCO_3$ = (3.00 – x) g

Since Na_2CO_3 in the products contains x g of unchanged reactant Na_2CO_3 and rest produced from $NaHCO_3$.

The wt. of Na_2CO_3 produced by $NaHCO_3$ = (2.752 – x)g

$$NaHCO_3 \longrightarrow Na_2CO_3 + (H_2O + CO_2)\uparrow$$

(3.0 – x) (2.752 – x)

Applying POAC for Na atom

1 moles of $NaHCO_3$ = 2 moles of Na_2CO_3 $\Rightarrow \frac{(3-x)}{84} = 2x\frac{(2.752-x)}{106}$

$\therefore$ x = 2.3244 g

$\therefore$ % of Na_2CO_3 = $\frac{2.3244}{3}$ 100 = 77.48 %

22. LIMITING REAGENT :

The reactant which consumed first into the reaction

When we are dealing with balance chemical equation then if number of moles of reactants are not in the ratio of stoichiometric coefficient of balanced chemical equation, then there should be one reactant which should be limiting reactant.

Ex. Three mole of Na_2CO_3 is reacted with 6 moles of HCl solution. Find the volume of CO_2 gas produced at STP. The reaction is

$$Na_2CO_3 + 2\,HCl \longrightarrow 2\,NaCl + CO_2 + H_2O$$

Sol.

From the reaction :	Na_2CO_3 + 2 HCl $\longrightarrow$ 2 NaCl + CO_2 + H_2O		
gives moles	3 mol		6 mol
given mole ratio	1	:	2
Stoichiometric coefficient ratio	1	:	2

❑ See here given moles of reactant are in stoichiometric coefficient ratio therefore non reactant left over. Now use Mole-mole analysis to calculate volume of CO_2 produced at STP

$$\frac{\text{Moles of } Na_2CO_3}{1} = \frac{\text{Mole of } CO_2 \text{ produced}}{1}$$

Moles of CO_2 produced = 3

volume of CO_2 produced at STP = 3 22.4 L = 67.2 L

Ex. 6 moles of Na_2CO_3 is reacted with 4 moles of HCl solution. Find the volume of CO_2 gas produced at STP. The reaction is

$$Na_2CO_3 + 2\ HCl \longrightarrow 2\ NaCl + CO_2 + H_2O$$

Sol.

From the reaction :	Na_2CO_3	+ 2 HCl	$\longrightarrow 2\ NaCl + CO_2 + H_2O$
gives moles of reactant	6	: 4	
given molar ratio	3	: 2	
Stoichiometric coefficient ratio	1	: 2	

❑ See here given number of moles of reactants are not in stoichiometric coefficient ratio. Therefore there should be one reactant which consumed first and becomes limiting reagent.

But the question is how to find which reactant is limiting, it is not very difficult you can easily find it. According to the following method.

23. HOW TO FIND LIMITING REAGENT :

Step : I

Divided the given moles of reactant by the respective stoichiometric coefficient of that reactant.

Step : II

See for which reactant this division come out to be minimum. The reactant having minimum value is limiting reagent for you.

Step : III

Now once you find limiting reagent then your focus should be on limiting reagent

From Step I & II $\quad Na_2CO_3 \qquad$ HCl

$\frac{6}{1} = 6 \qquad \frac{4}{2} = 2$ (division in minimum)

∴ **HCl is limiting reagent**

From Step III

From $\frac{\text{Mole of HCl}}{2} = \frac{\text{Mole of } CO_2 \text{ produced}}{1}$

∴ mole of CO_2 produced = 2 moles

∴ volume of CO_2 produced at S.T.P. = 2 22.4 = 44.8 L

Ex. Calculate the weight of FeO from 4 g VO and 5.75 g of Fe_2O_3. Also report the limiting reactant.

$VO + Fe_2O_3 \longrightarrow FeO + V_2O_5$

Sol.

Balanced equation	2VO +	$3Fe_2O_3 \longrightarrow$	6FeO +	V_2O_5
Moles before reaction	$\frac{4}{67}$	$\frac{5.75}{160}$	0	0
=	0.05970	0.03590		
Moles after reaction	(0.05970–0.0359)	0	$\left(\frac{6}{3}\times 0.0359\right)$	$\left(\frac{1}{3}\times 0.0359\right)$

As 2 moles of VO react with 3 moles of Fe_2O_3

∴ 0.05970 g moles of VO = $\frac{3}{2}$ 0.05970 = 0.08955 moles of Fe_2O_3

Moles of Fe_2O_3 available = 0.0359 only

Hence, Fe_2O_3 is the limiting reagent.

Moles of FeO formed = $\frac{6}{3}$ 0.0359

$\therefore$ Weight of FeO formed = 0.0359 2 72 = 5.17 g

$$\left(\frac{n_{FeO}}{n_{Fe_2O_3}} = \frac{6}{3}\right) \Rightarrow n_{FeO} = \frac{6}{3} \times n_{Fe_2O_3}$$

$$W_{FeO} = \frac{6}{3} \times n_{Fe_2O_3} \times M_{Fe_2O_3}$$

Ex. A mixture of KBr, NaBr weighing 0.56 g was treated with aqueous solution of Ag^+ and the bromide ion was recovered as 0.97 g of pure AgBr. What was the weight of KBr in the sample ?

Sol.

KBr	+	NaBr	+	Ag^+	$\longrightarrow$	AgBr
a g		(0.56 – a)g				0.97 g

Applying POAC for Br atoms,

Moles of Br in KBr + Moles of Br in NaBr = Moles of Br in AgBr

or 1 Moles of KBr + 1 Moles of NaBr = 1 Moles of AgBr

$$\Rightarrow \frac{a}{119} + \frac{(0.56 - a)}{103} = \frac{0.97}{188} \quad (M_{KBr} = 199,\ M_{NaBr} = 103,\ M_{AgBr} = 188)$$

$\therefore$ a = 0.2124 g

Percentage of KBr in the sample $= \frac{0.2124}{0.560} \times 100 =$ 37.93

Ex. The reaction

$2C + O_2 \longrightarrow 2CO$

is carried out by taking 24 g of carbon and 128 g of O_2.

Find out :

(i) Which reactant is left in excess ?

(ii) How much of it is left ?

(iii) How many moles of CO are formed ?

(iv) How many grams of other reactant should be taken so that nothing is left at the end of reaction ?

Sol.

	2C	+	O_2	$\longrightarrow$	2CO
Mole before reaction	$\frac{24}{12}$		$\frac{128}{32}$		
Mole after reaction	0		3		2

$\because$ Mole ratio of C : O_2 : CO :: 2 : 1 : 2

(i) O_2 is left in excess.

(ii) 3 moles of O_2 or 96 g of O_2 is left.

(iii) 2 moles of CO or 56 g of CO is formed.

(iv) To use O_2 completely, total 8 moles of carbon or 96 g of carbon is needed.

PERCENTAGE YIELD : The percentage yield of product $= \frac{\text{actual yield}}{\text{the theoretical maximum yield}}$ 100

24. SOLUTIONS :

A mixture of two or more substances can be a solution. We can also say that a solution is a homogeneous mixture of two or more substances 'Homogeneous' means 'uniform throughout'. Thus a homogeneous mixture, i.e., a solution, will have uniform composition throughout.

25. CONCENTRATION TERMS :

The following concentration terms are used to expressed the concentration of a solution. These are :

1. Molarity (M)
2. Molality (m)
3. Mole fraction (x)
4. % calculation
5. ppm

❑ Remember that all of these concentration terms are related to one another. By knowing one concentration term you can also find the other concentration terms. Let us discuss all of them one by one.

1. **Molarity (M) : The number of moles of a solute dissolved in 1 L (1000 ml) of the solution is known as the molarity of the solution.**

i.e., Molarity of solution = $\dfrac{\text{number of moles}}{\text{volume of solution in litre}}$

Let a solution is prepared by dissolving w g of solute of mol. wt. M in V mL water.

$\therefore$ Number of moles of solute dissolved = $\dfrac{w}{M}$

$\therefore$ V mL water have $\dfrac{w}{M}$ mole of solute

$\therefore$ 1000 mL water have $\dfrac{w \times 1000}{M \times V(\text{in mL})}$ $\Rightarrow$ $\therefore$ Molarity (M) $= \dfrac{w \times 1000}{(\text{Mol. wt of solute}) \times V(\text{in mL})}$

2. **Molality (m) : The number of moles of solute dissolved in 1000 g (1 kg) of a solvent is known as the molality of the solution.**

i.e., molality = $\dfrac{\text{number of moles of solute}}{\text{mass of solvent in gram}}$ 100

Let y g of a solute is dissolved in x g of a solvent. The molecular mass of the solute is m. Then y/m mole of the solute are dissolved in x g of the solvent. Hence

Molality $= \dfrac{y}{m \times x}$ 1000

3. **Mole fraction (x) : The ratio of number of moles of the solute or solvent present in the solution and the total number of moles present in the solution is known as the mole fraction of substances concerned.**

Let number of moles of solute in solution = n

Number of moles of solvent in solution = N

$\therefore$ Mole fraction of solute $(x_1) = \dfrac{n}{n+N}$

$\therefore$ Mole fraction of solvent $(x_2) = \dfrac{N}{n+N}$ $\Rightarrow$ also $x_1 + x_2 = 1$

4. **% Calculation** : The concentration of a solution may also expressed in terms of percentage in the following way.

(i) % weight by weight (w/w) : It is given as mass of solute present in per 100 g of solution.

i.e. % w/w = $\dfrac{\text{mass of solute in g}}{\text{mass of solution in g}}$ 100

[X % by mass means 100 g solution contains X g solute ; $\therefore$ (100 – X) g solvent]

(ii) % weight by volume (w/v) : It is given as mass of solute present in per 100 mL of solution.

i.e. $\% \text{ w/v} = \dfrac{\text{mass of solute in g}}{\text{volume of solution in mL}} \quad 100$

[X % $\left(\dfrac{w}{V}\right)$ means 100 mL solution contains X g solute]

(iii) % volume by volume (V/V) : It is given as volume of solute present in per 100 mL solution.

i.e. $\% \text{ V/V} = \dfrac{\text{Volume of solute}}{\text{Volume of solution in mL}} \quad 100$

5. **Parts per million (ppm)** : $\dfrac{\text{Mass of solute}}{\text{Mass of solvent}} \quad 10^6 \cong \dfrac{\text{Mass of solute}}{\text{Mass of solution}} \quad 10^6$

26. **EUDIOMETRY OR GAS ANALYSIS :**

Gaseous reactions are carried out in a special type of tube known as eudiometer tube. The tube is graduated in millimeters for volume measurement. The reacting gases taken in the eudiometer tube are exploded by sparks. The volumes of the products of a gases are determined by absorbing them in suitable reagents,

Eg.

Solvent	**gas (es) absorb**
KOH	CO_2, SO_2, Cl_2
Ammonical Cu_2Cl_2	CO
Turpentine oil	O_3
Alkaline pyrogallol	O_2
Water	NH_3, HCl
$CuSO_4$	H_2O

Eudiometry is mainly bases on Avogadro's law i.e. $V \propto n$ at the same temperature and pressure.

$\therefore$ The mole concept may be applied in solving the problems, keeping in mind that in a gaseous reaction the relative volumes (measured under identical conditions) of each reactant and product represent their relative numbers of moles.

eg :	A(g)	+	B(g)	$\rightarrow$	C(g)	+	D(g)
	a volumes		b volumes		c volumes		d volumes
	a moles		b moles		c moles		d moles

❑ Generally, explosions are carried out at STP and H_2O is assumed to be in liquid state, means its volume is negligible as compared to product gases.

Burning of hydrocarbon :

1. Hydrocarbon containing carbon and hydrogen only.

$$C_xH_y\,(g) + \left(x + \frac{y}{4}\right) O_2\,(g) \longrightarrow xCO_2\,(g) + \frac{y}{2} H_2O\,(\ell)$$

2. Hydrocarbon containing carbon and hydrogen and oxygen.

$$C_xH_yO_z\,(g) + \left(x + \frac{y}{2} - \frac{z}{2}\right) O_2\,(g) \longrightarrow xCO_2\,(g) + \frac{y}{2} H_2O\,(\ell)$$

Ex. What volume of oxygen at STP is required to effect complete combustion of 400 cm^3 of acetylene and what would be the volume of carbon dioxide formed?

Sol. $2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O$

2 volume of C_2H_2 require O_2 for complete combustion = 5 vol.

$\therefore$ 400 cm^3 of C_2H_2 will require O_2 for complete combustion $= \dfrac{5}{2} \times 400$

$= 1000\ cm^3$ at STP

2 volume of C_2H_2 produce CO_2 = 4 volume

$\therefore$ 400 cm^3 of C_2H_2 at STP will produce $CO_2 = \frac{4}{2} \times 400 = 800\ cm^3$

Thus, volume of CO_2 produced = 800 cm^3 at STP.

Ex. A gaseous hydrocarbon requires 6 times its own volume of O_2 for complete oxidation and produces 4 times its volume of CO_2. What is its formula ?

Sol. The balanced equation for combustion

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

$$1 \text{ volume} \left(x + \frac{y}{4}\right) \text{volume}$$

$\therefore\ x + \frac{y}{4} = 6$ (by equation)

or $4x + y = 24$ ....(1)

Again $x = 4$ since evolved CO_2 is 4 times that of hydrocarbon

$\therefore$ $16 + y = 24$ or $y = 8$ $\therefore$ formula of hydrocarbon C_4H_8

Ex. A 30 c.c. mixture of CO, CH_4 and He gases is exploded by an electric discharge at room temperature with excess of oxygen. The decrease in volume is found to be 13 c.c. A further contraction of 14 c.c. occurs when the residual gas is treated with KOH solution. Find out the composition of the gaseous mixture in terms of volume percentage.

Sol. Let the volume of CO be 'a' c.c. and CH_4 be 'b' c.c

$\therefore$ Volume of He = $(30 - a - b)$

on explosion with oxygen

$$CO\ (g) + \frac{1}{2}\ O_2\ (g) \longrightarrow CO_2\ (g)$$

$$CH_4\ (g) + 2O_2\ (g) \longrightarrow CO_2\ (g) + 2H_2O\ (\ell)$$

'a' c.c. of CO give 'a' c.c. of CO_2 and 'b' c.c. of CH_4 gives 'b' c.c. of CO_2.

Therefore the volume decrease is due to the consumption of O_2. O_2 consumed for 'a' c.c. of CO is $\frac{a}{2}$ c.c. and O_2 consumed for 'b' c.c. of CH_4 is '2b' c.c.

$\therefore\ \frac{a}{2} + 2b = 13$

The further contraction occurs because of the absorption of CO_2 by KOH, $a + b = 14$

$\therefore$ $b = 4$ c.c.

$\therefore$ $a = 10$ c.c.

$\therefore$ Percentage composition of CO $= \frac{10}{30}\ \ 100 = 33.33\ \%$

Percentage composition of $CH_4 = \frac{4}{30}\ \ 100 = 13.33\ \%$

Percentage composition of He $= \frac{(30 - 10 - 4)}{30}\ \ 100 = 53.33\ \%$