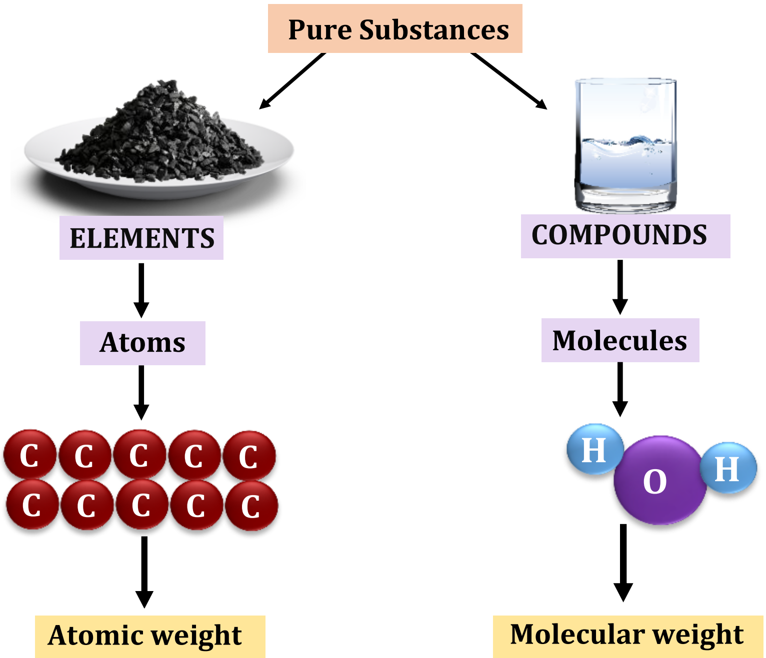
01 Some Basic Concepts of Chemistry- Mole Concept

Introduction :

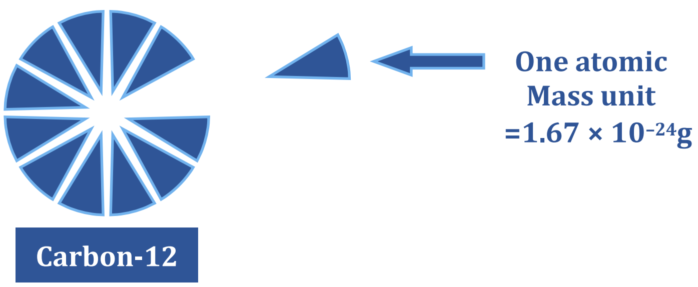
Chemistry deals with the composition, structure and properties of matter. These aspects can be best described and understood in terms of basic constituents of matter: **atoms** and **molecules**. That is why chemistry is called the science of atoms and molecules. Can we see weight and perceive these entities? Is it possible to count the number of atoms and molecules in a given mass of matter and have a quantitative relationship between the mass and number of these particles (atoms and molecules)? We will answer some of these questions in this unit. We would further describe how physical properties of matter can be quantitatively described using numerical values with suitable units.



Collectively atoms, molecules, ions, electrons, protons, neutrons etc. are called Fundamental Particles

Atomic Weight :

$$\text{Atomic weight (Relative Atomic weight)} = \frac{\text{Actual mass of one atom of element}}{1/12\text{th mass of one atom of Carbon}-12}$$



Relative Atomic Weight is a pure number, without any unit.

Example: Na = 23; H = 1; O = 16; N = 14, S = 32, Fe = 56, K = 39, C = 12

The quantity 1/12th mass of an atom of C^{12} is known as one atomic mass unit (amu).

1 atomic mass unit (amu) = 1.66×10^{-24} gram

= 1.66×10^{-27} kg

Molecular Weight :

$$\text{Molecular Weight} = \frac{\text{Actual mass of one molecule}}{\text{1/12th mass of one atom of Carbon – 12}}$$

Relative molecular weight is also a pure number, without any unit.

Example: H_2O = 18; NH_3 = 17; H_2SO_4 = 98; SO_2 = 64

Illustration 1:

How many atoms are there in 100 amu of He?

(1) 10 (2) 100 (3) 50 (4) 25

Solution:

The relative atomic weight of He is 4.

∴ The actual mass of one He atom = 4 amu.

No. of atoms in 100 amu = $\frac{100}{4}$ = 25

Atomicity :

Total number of atoms in one molecule of elementary substance is called as ATOMICITY.

Molecule	Atomicity
H_2	**2**
O_2	**2**
O_3	**3**
NH_3	**4** (1 N atom + 3 H atoms)
H_2SO_4	**7** (2 H atoms + 1 S atom + 4 O atoms)

Mole Concept :

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

It may be emphasised that the mole of a substance always contains the same number of entities, no matter what the substance may be. In order to determine this number precisely, the mass of a carbon-12 atom was determined by a mass spectrometer and found to be equal to 1.992648×10^{-23} g. Knowing that 1 mole of carbon weighs 12g, the number of atoms in it is equal to

$$\frac{12\text{g/mol } C^{12}}{1.992648 \times 10^{-23}\text{g/ } C^{12}\text{ atom}} = 6.022137 \times 10^{23}\text{ atoms/mol}$$

In a simple way, we can say that a mole has 6.022137 × 10^{23} entities (atoms, molecules or ions etc.)

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

Here entities may be atoms, ions, molecules or other subatomic entities. Chemists count the number of atoms and molecules by weighing. In a reaction we require these particles (atoms, molecules and ions) in a definite ratio. We make use of this relationship between numbers and masses of the particles for determining the stoichiometry of reactions.

Gram Atomic Mass and Gram Molecular Mass :

Gram Atomic Mass (or Mass of 1 Gram Atom)

When numerical value of atomic mass of an element is expressed in grams then the value becomes gram atomic mass.

Gram atomic mass = mass of 1 **gram atom** = mass of 1 **mole atom**

= mass of N_A atoms = mass of 6.023 × 10^{23} atoms.

Example:

Gram atomic mass of oxygen = mass of 1 **g atom** of oxygen = mass of 1 **mole atom** of oxygen.

= mass of N_A atoms of oxygen.

$$= \left(\frac{16}{N_A} g\right) \times N_A = 16 \text{ g}$$

Gram Molecular Mass (Mass of 1 Gram Molecule) :

When numerical value of molecular mass of the substance is expressed in grams then the value becomes gram molecular mass.

Gram molecular mass = mass of 1 **gram molecule** = mass of 1 **mole molecule**

= mass of N_A molecules = mass of 6.023 × 10^{23} molecules

Example:

Gram molecular mass of H_2SO_4 = mass of 1 **gram molecule** of H_2SO_4

= mass of 1 **mole molecule** of H_2SO_4

= mass of N_A molecules of H_2SO_4

$$= \left(\frac{98}{N_A} g\right) \times N_A = 98 \text{ g}$$

Formula to calculate moles are as following :

(i) Mole of atoms = $\dfrac{\text{weight(g)}}{\text{Gram atomic weight}}$

(ii) Mole of molecules = $\dfrac{\text{weight(g)}}{\text{Gram molecular weight}}$

(iii) Mole atoms = atomicity × mole molecules.

(iv) Mole of gas = $\dfrac{V}{22.4}$ (V = Volume of gas in litre at NTP or STP)

(v) Mole particles $n = \frac{N}{N_A}$ (N = Number of particles)

Mole of atoms = $\frac{\text{number of atoms}}{N_A}$ and Mole of molecules = $\frac{\text{number of molecules}}{N_A}$

Illustration 2:

Find the total number of H, S and O atoms in the following:

(1) 196 g H_2SO_4 (2) 196 amu H_2SO_4 (3) 5 mole $H_2S_2O_8$ (4) 3 molecules $H_2S_2O_6$

Solution:

(1) H = $4N_A$, S = $2N_A$, O = $8N_A$ atoms

(2) H = 4 atoms, S = 2 atoms, O = 8 atoms.

(3) H = $10N_A$, S = $10N_A$, O = 40 N_A atoms

(4) H = 6 atoms, S = 6 atoms, O = 18 atoms.

Illustration 3:

The charge on 1 gram ions of Al^{3+} is : (N_A = Avogadro number, e = charge on one electron)

(1) $\frac{1}{27} \times N_A$ e coulomb (2) $\frac{1}{3} \times N_A$ e coulomb (3) $\frac{1}{9} \times N_A$ e coulomb (4) $3 \times N_A$ e coulomb

Solution:

1 gram ion = 1 mole

charge on 1 mole Al^{3+} is = $3 \times e\ (N_A)$

Ans. (4)

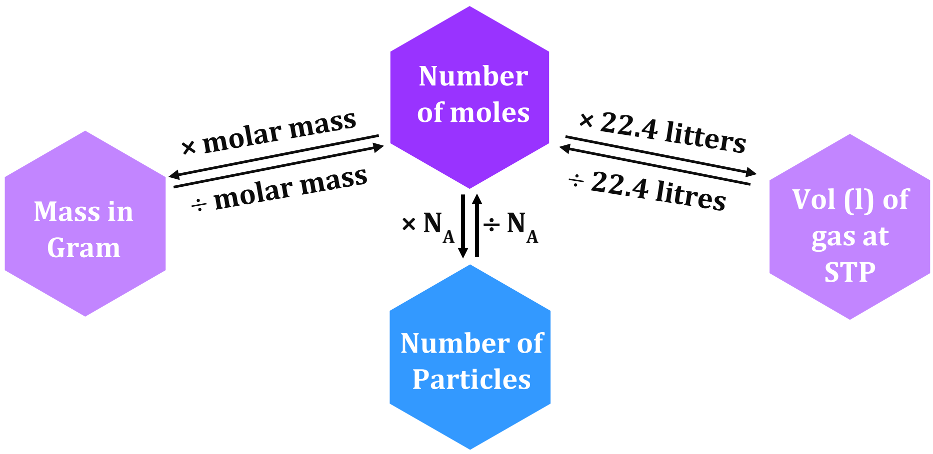


Illustration 4:

Calculate the number of moles in11.2 litre of Oxygen at STP.

Solution:

∵ At STP, 22.4 litre of Oxygen contains = 1 mol

∴ At STP, 11.2 litre of Oxygen contains = $\frac{1}{22.4} \times 11.2 = 0.5$ mol

OR

Mole of gas = $\frac{V}{22.4}$ (at NTP/STP)

$= \frac{11.2}{22.4} = 0.5$ mol

Illustration 5:

Find out the volume and moles of molecules in 56 g nitrogen at STP

Solution:

Gram molecular weight of N_2 is 28 g

(a) Calculation of volume: $\because$ 28 g of N_2 occupies = 22.4 litre at STP

$\therefore$ 56 g of N_2 occupies $= \frac{22.4}{28} \times 56$ litre = 44.8 litre at STP

(b) Calculation of mole: $\because$ 28 g of N_2 = 1 mol of N_2

$\therefore$ 56 g of $N_2 = \frac{1}{28} \times 56$ = 2 mol of N_2

Illustration 6:

Find out the moles & mass in 1.12 litre O_3 at STP.

Solution:

(a) Calculation of mole: $\because$ At STP, 22.4 litre of O_3 contain = 1 mol of O_3

$\therefore$ At STP, 1.12 litre of O_3 contain $= \frac{1}{22.4} \times 1.12 = 0.05$ mol of O_3

(b) Calculation of mass: Molecular weight of O_3 = 48 g

$\because$ weight of 22.4 litre of O_3 at STP is = 48 g

$\therefore$ weight of 1.12 litre of O_3 at STP is $= \frac{48}{22.4} \times 1.12 = 2.4$g

Illustration 7:

Calculate the volume and mass of 0.2 mol of O_3 at STP.

Solution:

(a) Calculation of volume: $\because$ volume of 1 mole of O_3 at STP = 22.4 litre

$\therefore$ volume of 0.2 mole of O_3 at STP = 22.4 × 0.2 = 4.48 litre

(b) Calculation of mass: $\because$ mass of 1 mol of O_3 = 48 g

$\therefore$ mass of 0.2 mol of O_3 = 48 × 0.2 g = 9.6 g

Illustration 8:

Calculate the number of molecules and number of atoms present in 1 g of nitrogen?

Solution:

Number of moles $(n) = \frac{\text{weight}}{M_w} = \frac{1}{28}$ and Number of molecules (N) = Mole of molecules × $N_A = \frac{1}{28} \times N_A$

$\because$ 1 molecule of N_2 gas contains = 2 atoms

$\therefore \frac{N_A}{28}$ molecules of N_2 gas contain $= 2 \times \frac{N_A}{28} = \frac{N_A}{14}$ atoms

Illustration 9:

Calculate the number of atoms present in one drop of water having mass 1.8 g.

Solution:

Number of moles of H_2O (n) in one drop of water $= \frac{\text{weight}}{M_w} = \frac{1.8}{18} = 0.1$ mol

Number of molecules of H_2O (N) = 0.1 N_A

$\because$ 1 molecule of H_2O contain = 3 atoms

$\therefore$ 0.1 N_A molecules of H_2O contain = 3 × (0.1 N_A) = 0.3 N_A atoms.

BEGINNER'S BOX-1

1. The modern atomic weight scale is based on.
(1) C^{12} (2) O^{16} (3) H^{1} (4) C^{13}
CMC152

2. Gram atomic weight of oxygen is
(1) 16 amu (2) 16 g (3) 32 amu (4) 32 g
CMC153

3. Molecular weight of SO_2 is :
(1) 64 g (2) 64 amu (3) 32 g (4) 32 amu
CMC154

4. 1 amu is equal to :-
(1) mass of $\frac{1}{12}$ of C^{12} (2) mass of $\frac{1}{14}$ of O^{16} (3) 1 g of H_2 (4) 1.66×10^{-24} kg
CMC155

Vapour Density :

Vapour density of a gas is the ratio of density of gas and that of hydrogen at the same temperature and pressure i.e. relative density of a gas with respect to the density of hydrogen gas at same temperature and pressure condition.

$$V.D. = \frac{d_{gas}}{d_{H_2}}$$

$$\because d = \frac{m}{v}$$

at STP, n = 1, m = mol weight, V = 22.4 L

$$d_{gas} = \frac{\text{mol wt.}_{gas}}{22.4\,L}$$

$$d_{H_2} = \frac{\text{mol wt.}_{H_2}}{22.4\,L}$$

$$V.D. = \frac{d_{gas}}{d_{H_2}} = \frac{\text{mol wt.}_{gas}}{\text{mol wt.}_{H_2}} = \frac{\text{mol wt.}_{gas}}{2}$$

$$\therefore \text{mol wt.}_{gas} = 2 \times V.D_{gas}$$

Some Important Points:

Vapour density is unitless term.

At NTP, the density of hydrogen gas = 0.000089 g/ml
$\simeq 9.0 \times 10^{-5}$ g/ml

$$\text{Relative density} = \frac{\text{Density of gas A}}{\text{Density of gas B}}$$

Illustration 10:

The formula of a hypothetical gas is $(N_2)_x$ and its vapour density is 70. Find x ?

Solution:

mol $wt._{gas}$ = 2 × Vapour density
= 2 × 70 = 140

Molecular weight of $(N_2)_x$ = x × molecular weight of N_2

Hence, 140 = x (28)
x = 5

Illustration 11:

A hydrocarbon C_nH_{2n+2} has Vapour density 29. What is the value of n?

Solution:

mol wt. = 2 × V.D = 2 × 29 = 58

Hence, 58 = n(12) + (2n + 2) 1

58 = 12n + 2n + 2

$\therefore$ n = 4

Illustration 12:

Vapour density of a gas is 11.2. The volume occupied by 11.2 g of this gas at STP is:

(1) 22.4 L (2) 11.2 L (3) 1 L (4) 2.25 L

Solution:

mol wt. = 2 × V.D.

= 2 × 11.2 = 22.4

$n_{gas} = \frac{11.2}{22.4} = \frac{1}{2}$

Volume occupied by gas at STP = mole ×22.4

Volume occupied by gas at STP = $\frac{1}{2}$ × 22.4 = 11.2 L

BEGINNER'S BOX-2

1. Calculate the number of atoms in 11.2 L of SO_2 gas at STP :

(1) $\frac{N_A}{2}$ (2) $\frac{3N_A}{2}$ (3) $3N_A$ (4) N_A

CMC156

2. Which of the following has maximum mass ?

(1) 0.1 gram atom of carbon (2) 0.1 mol of ammonia

(3) 6.02×10^{22} molecules of hydrogen (4) 1120 cc of carbon dioxide at STP

CMC157

3. The total number of electrons present in 18 mL of water :-

(1) 6.02×10^{22} (2) 6.02×10^{23} (3) 6.02×10^{24} (4) 6.02×10^{25}

CMC158

4. The volume of 1.0 g of hydrogen at NTP is :-

(1) 2.24 L (2) 22.4 L (3) 1.12 L (4) 11.2 L

CMC159

5. 11 grams of a gas occupy 5.6 litres of volume at STP. The gas is :-

(1) NO (2) N_2O_4 (3) CO (4) CO_2

CMC160

6. At NTP, 5.6 L of a gas weight 8 grams. The vapour density of gas is :-

(1) 32 (2) 40 (3) 16 (4) 8

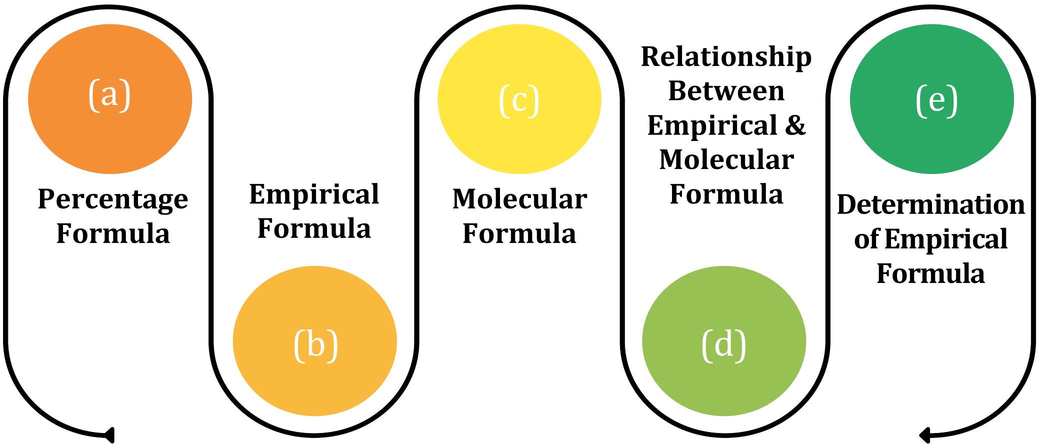
CMC161

7. The vapour densities of two gases are in the ratio of 1 : 3. Their molecular masses are in the ratio of :

(1) 1 : 3 (2) 1 : 2 (3) 2 : 3 (4) 3 : 1

CMC162

Percentage Formula, Empirical Formula & Molecular Formula :



(a) Percentage Formula

$$\% \text{ mass} = \frac{\text{Number of atoms} \times \text{atomic mass}}{\text{molecular mass}} \times 100$$

If number of atom (or atomicity) = 1; Molecular mass = minimum molecular mass

(b) Empirical Formula

The empirical formula of a compound expresses the simplest whole number ratio of atoms of various elements present in 1 molecule of the compound.

Molecular Formula	→	H_2O_2	CH_4	C_2H_6	$C_2H_4O_2$
		2 : 2	1 : 4	2 : 6	2 : 4 : 2
		1 : 1	1 : 4	1 : 3	1 : 2 : 1
Empirical Formula	→	$\boxed{HO}$	$\boxed{CH_4}$	$\boxed{CH_3}$	$\boxed{CH_2O}$

(c) Molecular Formula

The molecular formula of a compound represents the **actual number of atoms** present in 1 molecule of the compound i.e. it shows the real formula of its 1 molecule.

(d) Relationship between Empirical & Molecular Formula

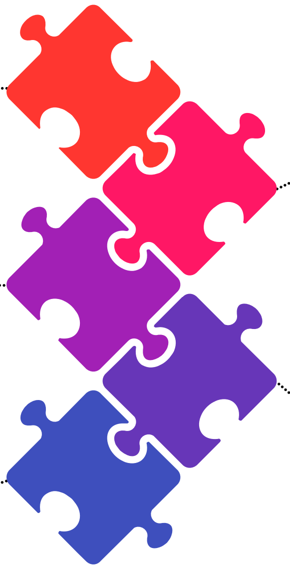
Molecular Formula = n × Empirical Formula

[Where n = natural no. (1, 2, 3, ...)]

or $\quad n = \dfrac{\text{Molecular Formula}}{\text{Empirical Formula}}$ $\quad$ or $\quad n = \dfrac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$

(e) Determination of Empirical Formula

Following steps are involved to determine the empirical formula of the compound –



(1) Firstly find the % by weight of each element present in 1 molecule of the compound

(2) The % by weight of each element is divided by its atomic weight. It gives atomic ratio of elements present in the

(3) Atomic ratio of each element is divided by the minimum value of atomic ratio so as to get the simplest ratio of atoms.

(4) If the value of simplest atomic ratio is fractional then raise the value to the nearest whole number or multiple with suitable coefficient to convert it into nearest whole number.

(5) Write the Empirical formula according to the simplest ratio of atoms.

Illustration 13:

In a compound, X is 75.8% and Y is 24.2% by weight. If atomic weight of X and Y are 24 and 16 respectively, then, calculate the empirical formula of the compound.

Solution:

Elements	%	Atomic weight	$\frac{\%}{\text{Atomic weight}}$	Simplest ratio	Ratio
X	75.8%	24	$\frac{75.8}{24}=3.1$	$\frac{3.1}{1.5}=2.06$	2
Y	24.2%	16	$\frac{24.2}{16}=1.5$	$\frac{1.5}{1.5}=1$	1

Empirical formula = X_2Y

Illustration 14:

In a compound Carbon = 52.2%, Hydrogen = 13%, Oxygen = 34.8% are present and molecular mass of the compound is 92. Calculate molecular formula of the compound?

Solution:

Elements	%	Atomic weight	$\frac{\%}{\text{Atomic weight}}$	Simplest ratio	Ratio
C	52.2	12	$\frac{52.2}{12} = 4.35 = 4.4$	$\frac{4.4}{2.2} = 2$	2
H	13	1	$\frac{13}{1} = 13$	$\frac{13}{2.2} = 5.9$	6
O	34.8	16	$\frac{34.8}{16} = 2.2$	$\frac{2.2}{2.2}=1$	1

Empirical formula = C_2H_6O

Empirical mass = 12 × 2 + 16 + 6 = 46

$$n = \frac{\text{Molecular mass}}{\text{Empirical mass}} = \frac{92}{46} = 2$$

Molecular formula = 2 × (C_2H_6O) = $C_4H_{12}O_2$

BEGINNER'S BOX-3

1. A hydrocarbon contain 80% C. The vapour density of compound is 30. Empirical formula of compound is :-

(1) CH_3 (2) C_2H_6 (3) C_4H_{12} (4) C_4H_8

CMC163

2. Two elements X (Atomic weight = 75) and Y (Atomic weight = 16) combine to give a compound having 75.8% of X. The empirical formula of compound is :

(1) XY (2) X_2Y (3) X_2Y_2 (4) X_2Y_3

CMC164

3. In a compound element A (Atomic weight = 12.5) is 25% and element B (Atomic weight = 37.5) is 75% by weight. The Empirical formula of the compound is :

(1) AB (2) A_2B (3) A_2B_2 (4) A_2B_3

CMC165

Stoichiometry Based Concept (Problems Based on Chemical Reaction) :

One of the most important aspects of a chemical equation is that when it is written in the balanced form, it gives quantitative relationships between the various reactants and products in terms of moles, molecules and volumes. This is called stoichiometry (Greek word, meaning 'to measure an element'). For example, a balanced chemical equation along with the quantitative information conveyed by it is given below:

$CaCO_{3\,(s)}$	+	$2HCl_{(aq.)}$	$\longrightarrow$	$CaCl_{2\,(aq.)}$	+	$H_2O_{(l)}$	+	$CO_{2\,(g)}$
1 Mole		2 Mole		1 Mole		1 Mole		1 Mole
40 + 12 + 3 × 16		2(1 + 35.5)		40 + 2 × 35.5		2 × 1 + 16		12 + 2 × 16
= 100 g		= 73 g		= 111 g		= 18 g		= 44 g

Thus,

(i) 1 mole of calcium carbonate reacts with 2 moles of hydrochloric acid to give 1 mole of calcium chloride, 1 mole of water and 1 mole of carbon dioxide.

(ii) 100 g of calcium carbonate react with 73 g hydrochloric acid to give 111 g of calcium chloride, 18 g of water and 44 g (or 22.4 litres at STP) of carbon dioxide.

1		3		2 Stoichiometry
$N_{2\,(g)}$	+	$3H_{2\,(g)}$	→	$2NH_{3\,(g)}$
1 mole	+	3 mole	→	2 mole
22.4 litre	+	3 × 22.4 litre	→	2 × 22.4 litre (at STP)
1 litre	+	3 litre	→	2 litre
1000 mL	+	3000 mL	→	2000 mL
1 mL	+	3 mL	→	2 mL
28 gm	+	6 gm	→	34g
				(According to the law of conservation of mass)

Mass is not expressed in the ratio of stoichiometric coefficient.

The quantitative information conveyed by a chemical equation helps in a number of calculations. The problems involving these calculations may be classified into the following different types:

(1) Mass - Mass Relationships

mass of one of the reactants or products is given and the mass of some other reactant or product is to be calculated.

(2) Mass - Volume Relationships

mass/volume of one of the reactants or products is given and the volume/mass of the other is to be calculated.

(3) Volume – Volume Relationships

volume of one of the reactants or the products is given and the volume of the other is to be calculated.

The general method of calculations for all the problems of the above types consists of the following steps:

(i) Write down the balanced chemical equation.

(ii) Write the relative number of moles or the relative masses (gram atomic or molecular masses) of the reactants and the products below their formula.

(iii) In case of a gaseous substance, write down 22.4 litres at STP below the formula in place of 1 mole

(iv) Apply unitary method to make the required calculations.

Quite often one of the reactants is present in larger amount than the other as required according to the balanced equation. The amount of the product formed then depends upon the reactant which has reacted completely. This reactant is called the limiting reactant. The excess of the other is left unreacted.

Type I Involving Mass-Mass Relationship

Illustration 15:

How much iron can be theoretically obtained in the reduction of 1 kg of Fe_2O_3?

Solution:

Writing the balanced equation for the reaction.

1	3	2	3

$$Fe_2O_3 \quad + \quad 3H_2 \longrightarrow 2Fe \quad + \quad 3H_2O$$

$$n = \frac{\text{weight}}{M_w} = \frac{1000}{160}\ \text{mol}$$

The equation shows that 2 mol of iron are obtained from 1 mol of ferric oxide.

Hence, the obtained no. of moles of Fe $= \dfrac{2 \times 1000}{160} = 12.5\ \text{mol} = \dfrac{\text{weight}}{\text{Atomic weight}} = \dfrac{\text{weight}}{56}$

Weight of iron obtained = 12.5 × 56 g = 700 g

Illustration 16:

What amount of silver chloride is formed by the action of 5.850 g of sodium chloride on an excess of silver nitrate?

Solution:

Writing the equation for the reaction

1	1	1	1

$$\underset{1}{NaCl} + \underset{1}{AgNO_3} \longrightarrow \underset{1}{AgCl} + \underset{1}{NaNO_3}$$

$$\text{n of NaCl} = \frac{\text{weight}}{M_w} = \frac{5.85}{58.5} = 0.1 \text{ mol}$$

1 mol of AgCl is obtained with 1 mol of NaCl

Hence, the number of moles of AgCl obtained with 0.1 mol of NaCl = 0.1 mol

$$\because \; n = \frac{\text{weight}}{M_W} \quad \therefore 0.1 \text{ mol} = \frac{\text{weight}}{M_w} = \frac{\text{weight}}{143.5}$$

$\therefore$ weight = 0.1 × 143.5 g = 14.35 g.

Type II Mass - Volume Relationship

Illustration 17:

For complete combustion of 3g ethane find required moles of O_2 and formed moles of CO_2. Also find volume of O_2 and CO_2 at STP.

Solution:

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

$$n = \frac{\text{weight}}{M_w} = \frac{3}{30} = \frac{1}{10} = 0.1 \text{ mole}$$

(a) Required moles of $O_2 = \frac{7}{2} \times 0.1 = 0.35$ mol

volume of O_2 at STP = 0.35 × 22.4 = 7.84 litre

(b) Produced moles of $CO_2 = \frac{4}{2} \times 0.1 = 0.2$ mol

volume of CO_2 at STP = 0.2 × 22.4 = 4.48 litre

Illustration 18:

In the following reaction, if 10 g of H_2, react with N_2. What will be the volume of NH_3 at STP.

$$N_2 + 3H_2 \rightarrow 2NH_3$$

Solution:

$$\underset{}{\overset{1}{N_2}} + \underset{10\text{ g}}{\overset{3}{3H_2}} \longrightarrow \overset{2}{2NH_3}$$

$$n = \frac{\text{weight}}{M_w} = \frac{10}{2} = 5 \text{ mol}$$

Produced moles of $NH_3 = \frac{2}{3} \times 5 = \frac{10}{3}$, Volume of NH_3 at STP = $\frac{10}{3} \times 22.4 = 74.67$ litre

Limiting Reagent (L.R.) Concept :

Limiting Reagent (L.R.): The reactant which is completely consumed in a reaction is called as L.R.

Ex.	1		2		1		2 $\leftarrow$ Stoichiometry
	A	+	2B	$\longrightarrow$	C	+	2D
given	3 mol		9 mol				
	3 – 3 = 0 mol		9 – 6 = 3 mol		3 mol		6 mol

L.R. = A

Formula for checking L.R. = $\dfrac{\text{Given value (moles, volume, or molecules)}}{\text{Stoichiometric Coefficient}}$

Least value indicate the L.R.

Ex. A B

$\frac{3}{1}=3$ $\frac{9}{2}= 4.5$

3 < 4.5 So, A is L.R.

Identification of L.R concept: More than 1 initial quantities of reactants are given

Illustration 19:

A + 5B $\rightarrow$ C + 3D. In this reaction identify the limiting reagent if the reaction starts with 10 mol of A and 10 mol of B.

Solution:

For A For B

$\frac{10}{1} = 10$ $\frac{10}{5} = 2$

2 < 10 So B is L.R.

Illustration 20:

4 g of H_2 (g) and 32 g of O_2 (g) are mixed and exploded. $H_{2(g)} + \frac{1}{2} O_{2(g)} \longrightarrow H_2O_{(g)}$; In the above reaction what is the volume of water vapour formed at STP?

Solution:

1		$\frac{1}{2}$		1
$H_{2(g)}$	+	$\frac{1}{2} O_{2(g)}$	$\rightarrow$	$H_2O_{(g)}$
4 g		32 g		
$n = \frac{4}{2} = 2\text{mol}$		$n = \frac{32}{32} = 1\text{ mol}$		
For H_2		For O_2		
$\frac{2}{1} = 2$		$\frac{1}{\frac{1}{2}} = 2\text{ mol}$	$\rightarrow$	identification of L.R.
				Both are L.R.

Moles of $H_2O_{(g)}$ = 2 mol = $\frac{V}{22.4}$

Volume of $H_2O_{(g)}$ at STP = 22.4 × 2 = 44.8 litre

Illustration 21:

At NTP, In a container 100 mL of N_2 and 100 mL of H_2 are mixed together. Find out the produced volume of NH_3.

Solution:

Balanced equation will be	N_2	+	$3H_2$	$\rightarrow$	$2NH_3$
Given:	100mL		100 mL		

For determination of Limiting reagent divide the given quantities by stoichiometric coefficients

$$\frac{100}{1} = 100 \qquad \frac{100}{3} = 33.3 \text{ (Limiting reagent)}$$

In this reaction, H_2 is limiting reagent so reaction will proceed according to H_2.

According to stoichiometry from 3 mL of H_2, produced volume of NH_3 = 2 mL

That is, from 100 mL of H_2, produced volume of $NH_3 = \frac{2}{3} \times 100 = 66.6$ mL

Combustion reaction: (Problem based on combustion reactions):

For balancing the combustion reaction: First of all balance C atoms, then balance H atoms, finally balance Oxygen atoms.

For Example:

Combustion reaction of C_2H_6	$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$ (skeleton equation)
Balance C atoms	$C_2H_6 + O_2 \longrightarrow 2CO_2 + H_2O$
Now balance H atoms	$C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$
Now balance Oxygen atoms	$C_2H_6 + 7/2\ O_2 \longrightarrow 2CO_2 + 3H_2O$

Illustration 22:

For complete combustion of 1.12 litre of butane (C_4H_{10}), the produced volume of $H_2O_{(g)}$ & $CO_2(g)$ will be:

Solution:

1		$\frac{13}{2}$		4		5
$C_4H_{10(g)}$	+	$\frac{13}{2}O_{2(g)}$	$\rightarrow$	$4CO_{2(g)}$	+	$5H_2O_{(g)}$
1.12 litre						

Volume of $H_2O_{(g)}$ at STP = 5 × 1.12 = 5.6 litre

Volume of $CO_{2(g)}$ at STP = 4 × 1.12 = 4.48 litre

Illustration 23:

For complete combustion of 5 mol propane (C_3H_8), find required moles of O_2. Also, find volume of O_2 used at STP.

Solution:

For C_3H_8, the combustion reaction is:

1		5		3		4
$C_3H_{8(g)}$	+	$5O_{2(g)}$	$\rightarrow$	$3CO_{2(g)}$	+	$4H_2O_{(g)}$
5 mol						

Required moles of $O_2 = 5 \times 5 = 25$ mole $= \frac{V}{22.4}$

Volume of O_2 gas at STP = 25 × 22.4 = 560 litre

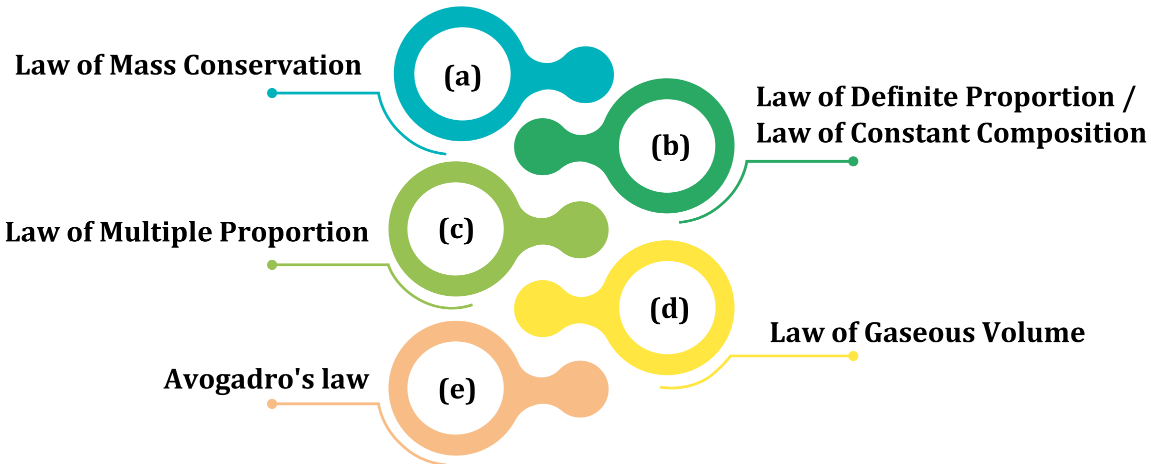
BEGINNER'S BOX-4

1. 1.5 mol of O_2 combine with Mg to form oxide MgO. The mass of Mg (At. mass 24) that has combined is :
(1) 72 g (2) 36 g (3) 24 g (4) 94 g
CMC166

2. What quantity of lime stone on heating will give 56 kg of CaO :-
(1) 1000 kg (2) 56 kg (3) 44 kg (4) 100 kg
CMC167

3. For reaction A + 2 B $\rightarrow$ C. The amount of product formed by starting the reaction with 5 mol of A and 8 mol of B is :
(1) 5 mol (2) 8 mol (3) 16 mol (4) 4 mol
CMC168

Laws of Chemical Combination :



(a) Law of Mass Conservation (Law of Indestructibility of Matter)

"It was given by **Lavoisier** and tested by **Landolt"**

According to this law, mass can neither be created nor be destroyed in a balanced chemical reaction or physical reaction. But one form is changed into another form. This is called as law of mass conservation.

If the reactants are completely converted into products, then the sum of the mass of reactants is equal to the sum of the mass of products.

$\boxed{\text{Total mass of reactants = Total mass of products.}}$

If reactants are not completely consumed then the relationship will be:

$\boxed{\text{Total mass of reactants = Total mass of products + Mass of unreacted reactants}}$

Illustration 24:

10g of $CaCO_3$ on heating gives 4.4 g of CO_2, then, determine weight of produced CaO in quintal.

Solution:

$CaCO_3 \rightarrow CaO + CO_2$

10 g   x g   4.4 g

According to law of conservation of mass

10 = 4.4 + x

10 – 4.4 = x

x = 5.6 g

$$\left\{\begin{array}{l} 1\text{quintal} = 100 \text{ kg} \\ 1\text{kg} = 1000\text{g} \end{array}\right.$$

Weight of CaO(x) = $5.6 \times \frac{kg}{1000} = 5.6 \times 10^{-3}$ kg $= 5.6 \times 10^{-3} \times \frac{1}{100}$ quintal $= 5.6 \times 10^{-5}$ quintal

(b) Law of Definite Proportion / Law of Constant Composition

"It was given by **Proust**."

According to this law, a compound can be obtained from different sources. But the ratio of each component (by weight) remain same i.e. it does not depend on the method of its preparation or the source from which it has been obtained.

For example: molecule of ammonia always has the formula NH_3. That is one molecule of ammonia always contains, one atom of nitrogen and three atoms of hydrogen or 17 g of NH_3 always contains 14 g of nitrogen and 3 g of hydrogen.

Ex. Water can be obtained from different sources but the ratio of weight of H and O remains same.

$2H_2O \leftarrow 2H_2 + O_2$
Sea water
Tap water
River water
Ganga Jal
Rain water

	H_2		O
	2	:	16
or	1	:	8

(c) Law of Multiple Proportion

"It was given by **John Dalton**"

According to law of Multiple proportion if two elements combine to form more than one compound then the different mass of one element which combine with a fixed mass of other element bear a simple ratio to one another.

The following examples illustrate this law.

Nitrogen and oxygen combine to form five oxides, which are: Nitrous oxide (N_2O), nitric oxide (NO), nitrogen trioxide (N_2O_3), nitrogen tetraoxide (N_2O_4) and nitrogen pentaoxide (N_2O_5).

Weight of oxygen which combine with the fixed weight of nitrogen in these oxides are calculated as under:

Oxide	N_2O	NO	N_2O_3	N_2O_4	N_2O_5
Ratio of weight	28 : 16	14 : 16	28 : 48	28 : 64	28 : 80

Number of parts by weight of oxygen which combine with 14 parts by weight of nitrogen from the above are 8, 16, 24, 32 and 40 respectively. Their ratio is 1 : 2 : 3 : 4 : 5, which is a simple ratio. Hence, the law is illustrated.

(d) Law of Gaseous Volume

"It was given by **Gay Lussac**"

According to this law, in the gaseous reaction, the reactants are always combined in a simple ratio by volume and form products, which is **simple ratio by volume** at same temperature and pressure.

Ex. One volume of nitrogen combines with 3 volumes of hydrogen to from 2 volumes of ammonia.

Simple ratio 1 : 3 : 2

1	3	2

$$N_{2(g)} \quad + \quad 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

1 Volume 3 Volume 2 Volume

Special Note: This law is used only for gaseous reaction. It relates volume to mole or molecules. But it doesn't relate with mass.

(e) Avogadro's law

"Equal volumes of all gases contain equal number of molecules at same temperature and pressure."

Ex.

1	1	2	← Stoichiometry
$H_{2(g)}$ +	$Cl_{2(g)}$ ⟶	$2HCl_{(g)}$	
1 Volume	1 Volume	2 Volume	
N molecules	N molecules	2N molecules	
$\boxed{\frac{1}{2}\text{molecule}}$ (1atom)	$\frac{1}{2}$ molecule (1 atom)	1 molecule	

It is correct as molecule is divisible.

Solution :

When two or more chemically non-reacting substances are mixed together forming homogeneous mixture the mixture is called as solution.

When the solution is composed of only two chemical substances, it is termed as binary solution, similarly, it is called ternary and quaternary if it is composed of three and four components respectively.

For binary solution: Solution = solute + solvent

- Generally the component present in lesser amount than other component in solution, is called solute.
- Generally the component present in greater amount than all other components, is called the solvent.
- Physical state of solvent and solution is same.

Solution	Solute + Solvent	
	(B)	(A)
moles	n	N
mass	$w_{(g)}$	$W_{(g)}$
molar mass	m	M
mole fraction	x_B	x_A

Ex.1: In a syrup (liquid solution) containing 60 g sugar (a solid) and 40 g water (a liquid) same aggregation as solution water is termed as the solvent.

Ex.2: In a solution of alcohol and water; having 10 mL alcohol and 20 mL water, water is solvent and alcohol will be solute.

- On the basis of amount of solute, solutions can be classified in two ways.

(a) Dilute Solution

A solution in which relatively a small amount of solute is dissolved in large amount of solvent is called a dilute solution.

(b) Concentrated Solution

A solution in which relatively a large amount of the solute is present is called a concentrated solution.

Concentration Terms :

Molarity (M)

The number of moles of solute present in one litre solution is called its molarity (M).

Molarity = $\frac{\text{Number of moles of solute}}{\text{volume of solution (L)}} = \frac{n}{V_{(L)}}$

Molality (m)

The number of moles of solute present in 1000 gram of the solvent is called molality of the solution.

Molality of a solution = $\frac{\text{Number of moles of solute}}{\text{weight of solvent (kg)}} = \frac{\text{Number of moles of solute} \times 1000}{\text{weight of solvent (g)}}$

Mole Fraction

The ratio of the number of moles of one component to the total number of moles of all the components present in the mixture, is called as the mole fraction of that component.

Mole fraction of solute X_B = $\frac{\text{moles of solute (n)}}{\text{moles of solute (n)} + \text{moles of solvent (N)}}$

Mole fraction of solvent X_A = $\frac{\text{moles of solvent (N)}}{\text{moles of solute (n)} + \text{moles of solvent (N)}}$

$$\boxed{X_A + X_B = 1}$$

Concentration in terms of percentage

(i) Percent By Mass (w/W)

Mass of solute (in g) present in 100 g of solution (g) is called mass percent of the solute.
Where 'w' gram of solute is dissolved in W gram of solvent.

Mass percent = $\frac{\text{Mass of solute (g)} \times 100}{\text{Mass of solution (g)}} = \frac{w \times 100}{w + W}$

Mass percent is independent of temperature.

(ii) Percent By Volume (v/V)

The volume of solute in mL present in 100 mL of solution is called volume percent.

Volume percent = $\frac{\text{Volume of solute} \times 100}{\text{Volume of solution}}$

(iii) Percent by strength /percentage mass by volume $\left(\frac{w}{V}\right)$: Mass of solute (in g) present in 100 mL solution is called percent mass by volume.

$$\% \left(\frac{w}{V}\right) = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100$$

Parts Per Million (ppm)

This method is used for expressing the concentration of very dilute solutions such as hardness of water, air pollution etc.

ppm of substance (by mass) $= \frac{\text{Mass of solute (g)} \times 10^6}{\text{Mass of solution (g)}}$

ppm (by volume) = $\frac{\text{Volume of solute(mL)} \times 10^6}{\text{Volume of solution(mL)}}$

$$\text{ppm}\left(\text{by } \frac{w}{V}\right) = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 10^6$$

BEGINNER'S BOX ANSWER KEY

BEGINNER'S BOX-1	Que.	1	2	3	4
	Ans.	**1**	**2**	**2**	**1**

BEGINNER'S BOX-2	Que.	1	2	3	4	5	6	7
	Ans.	**2**	**4**	**3**	**4**	**4**	**3**	**1**

BEGINNER'S BOX-3	Que.	1	2	3
	Ans.	**1**	**4**	**1**

BEGINNER'S BOX-4	Que.	1	2	3
	Ans.	**1**	**4**	**4**