

Chapter 1

Some Basic Concepts of Chemistry

MOLE CONCEPT

1 mole is represented in the form of atoms, molecules and ions as :

For atoms → 1 gm atom

For molecules → 1 gm molecule

For ions → 1 gm ion

Moles can be calculated in following manner :

$$(a) \text{ Number of moles of molecules} = \frac{\text{Weight of substance (in g)}}{\text{Molecular weight}}$$

$$(b) \text{ Number of moles of atoms} = \frac{\text{Weight of substance (in g)}}{\text{Atomic weight}}$$

$$(c) \text{ Number of moles of gases} = \frac{\text{Volume of gas at NTP (in litres)}}{22.4}$$

(1 mole of any gas occupies a volume of 22.4 litres at N.T.P., N.T.P. Corresponds to 0°C and 1 atm pressure)

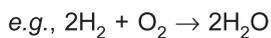
$$(d) \text{ Number of moles of atoms/molecules/ions} = \frac{\text{Number of atoms/molecules/ions}}{\text{Avogadro constant}}$$

(Avogadro constant is equal to 6.022×10^{23}).

LIMITING REAGENT

In the given reaction if number of quantities (either in gm/mole/molecules) are present with exact co-efficients, chemical reaction goes to completion without any reactant left unused.

However if exact proportion is not present then the one which gets totally consumed is known as **limiting reagent** (Limiting reagent decides the product quantity for given information).



In above e.g. 2 moles of H_2 reacts exactly with 1 mole of O_2 to give 2 moles of H_2O . If given moles of H_2 are 4 moles and that of O_2 are 0.5, then 0.5 moles O_2 will act as limiting reagent as it is in minimum amount and product formation is given w.r.t. O_2 i.e., 1 mole of H_2O .

EQUIVALENT WEIGHT

Equivalent weight of substance is defined as number of parts by weight of given substance which combines or displaces 1 part by weight of hydrogen (11.2 L of H₂ at STP), 8 parts by weight of oxygen (5.6 L of O₂ at STP), 35.5 parts by weight of chlorine (11.2 L of Cl₂ at STP).

- Equivalent weight of element = $\frac{\text{Atomic weight}}{\text{Valency}}$
- Equivalent weight of acids = $\frac{\text{Molecular weight of acid}}{\text{Basicity}}$
- Equivalent weight of bases = $\frac{\text{Molecular weight of base}}{\text{Acidity}}$
- Equivalent weight of salts = $\frac{\text{Molecular weight of salt}}{\text{Total + ve or - ve charge}}$
- Equivalent weight of reducing agent = $\frac{\text{Molecular weight}}{\text{Number of e}^- \text{ lost by one molecule}}$
- Equivalent weight of oxidising agent = $\frac{\text{Molecular weight}}{\text{Number of e}^- \text{ gained by one molecule}}$

n-FACTOR OR VALENCE FACTOR

n-factor is very important for both redox and non redox reactions through which we predict the following two informations:

- (a) It predicts the molar ratio of the species taking part in reactions i.e. reactants. The reciprocal of n-factor's ratio of the reactants is the molar ratio of the reactants.

For example : If X (having n-factor = a) reacts with Y (having n-factor = b) then its n-factor's ratio is a : b, so molar ratio of X to Y is b : a.

It can be represented as $bX + aY \longrightarrow \text{Products}$
(nf=a) (nf=b)

- (b) Equivalent weight = $\frac{\text{Molecular weight}}{\text{n - factor}}$ or $\frac{\text{Atomic weight}}{\text{n - factor}}$

LAW OF EQUIVALENCE

According to law of equivalence, for each and every reactant and product,
 Equivalents of each reactant reacted = Equivalents of each product formed.

For example :

Suppose, the reaction is taking place as under



Then according to law of equivalence,

$$\begin{aligned} \text{Equivalents of A reacted} &= \text{Equivalents of B reacted} \\ &= \text{Equivalents of C produced} \\ &= \text{Equivalents of D produced} \end{aligned}$$

$$\begin{aligned} \text{Equivalents of any substance} &= \frac{\text{Weight of substance (ng)}}{\text{Equivalent weight}} \\ &= \text{Normality (N)} \times \text{Volume (V) (In litre)} \end{aligned}$$

$$\text{Normality (N)} = n\text{-factor} \times \text{Molarity (M)}$$

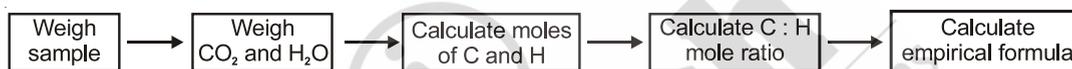
Normality and molarity are temperature dependent. Since on changing the temperature, the volume of solution changes, so normality and molarity change.

EMPIRICAL AND MOLECULAR FORMULA

- (a) **Empirical Formula of a compound** is the simplest whole number ratio of the atoms of elements constituting its one molecule. The sum of atomic masses of the atoms representing empirical formula is called **empirical** formula mass.
- (b) **Molecular Formula** of a compound shows the actual number of the atoms of the elements present in its one molecule. The sum of atomic masses of the atoms representing molecule is called **molecular mass**.
- (c) **Relationship between Empirical Formula and Molecular Formula**

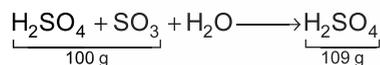
Molecular formula = $n \times$ empirical formula where n is a simple whole number having values of 1, 2, 3... etc.

Also, $n = \text{Molecular formula mass} / \text{Empirical formula mass}$.



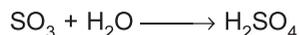
% of free SO₃ in Oleum

Strength of oleum is expressed in percentage e.g., oleum sample is 109%, It means 100 g of this oleum reacts with H₂O to form 109 g of H₂SO₄



Actually SO₃ react with H₂O to form H₂SO₄. Here 9 g H₂O or $\frac{1}{2}$ mole H₂O react with SO₃.

$$\text{So, moles of SO}_3 \text{ in sample} = \frac{1}{2} \text{ mole}$$



$$\text{So, wt. of SO}_3 = \frac{1}{2} \times 80 \text{ (molecular weight of SO}_3\text{)} = 40 \text{ g}$$

$$\% \text{ free SO}_3 = 40\%$$

$$\text{Note : Moles of H}_2\text{O} = \left(\frac{\text{Percentage labelling} - 100}{18} \right)$$

