Chapter 1

Some Basic Concepts of Chemistry

MOLE CONCEPT

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

For atoms $\rightarrow$ 1 gm atom

For molecules $\rightarrow$ 1 gm molecule

For ions $\rightarrow$ 1 gm ion

Moles can be calculated in following manner :

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

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

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

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

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

(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).

e.g., $2H_2 + O_2 \rightarrow 2H_2O$

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 H_2 at STP), 8 parts by weight of oxygen (5.6 L of O_2 at STP), 35.5 parts by weight of chlorine (11.2 L of Cl_2 at STP).

1. Equivalent weight of element = $\dfrac{\text{Atomic weight}}{\text{Valency}}$

2. Equivalent weight of acids = $\dfrac{\text{Molecular weight of acid}}{\text{Basicity}}$

3. Equivalent weight of bases = $\dfrac{\text{Molecular weight of base}}{\text{Acidity}}$

4. Equivalent weight of salts = $\dfrac{\text{Molecular weight of salt}}{\text{Total} + \text{ve or} - \text{ve charge}}$

5. Equivalent weight of reducing agent = $\dfrac{\text{Molecular weight}}{\text{Number of } e^- \text{ lost by one molecule}}$

6. Equivalent weight of oxidising agent = $\dfrac{\text{Molecular weight}}{\text{Number of } e^- \text{ gained by one molecule}}$

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 $\underset{(nf=a)}{bX} + \underset{(nf=b)}{aY} \longrightarrow \text{Products}$

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

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

$A + B \rightarrow C + D.$

Then according to law of equivalence,

Equivalents of A reacted = Equivalents of B reacted

= Equivalents of C produced

= Equivalents of D produced

$$\text{Equivalents of any substance} = \frac{\text{Weight of substance (in g)}}{\text{Equivalent weight}}$$

$$= \text{Normality (N)} \times \text{Volume (V) (In litre)}$$

Normality (N) = n-factor × 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 × empirical formula where n is a simple whole number having values of 1, 2, 3... etc.

Also, n = Molecular formula mass/Empirical formula mass.

Weigh sample → Weigh CO_2 and H_2O → Calculate moles of C and H → Calculate C : H mole ratio → Calculate empirical formula

% of free SO_3 in Oleum

Strength of oleum is expressed in percentage *e.g.*, oleum sample is 109%, It means 100 g of this oleum reacts with H_2O to form 109 g of H_2SO_4

$$\underbrace{H_2SO_4 + SO_3}_{100\text{ g}} + H_2O \longrightarrow \underbrace{H_2SO_4}_{109\text{ g}}$$

Actually SO_3 react with H_2O to form H_2SO_4. Here 9 g H_2O or $\frac{1}{2}$ mole H_2O react with SO_3.

So, moles of SO_3 in sample = $\frac{1}{2}$ mole

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

So, wt. of SO_3 = $\frac{1}{2} \times 80$ (molecular weight of SO_3) = 40 g

% free SO_3 = 40%

> ***Note :*** *Moles of* $H_2O = \left(\dfrac{\textit{Percentage labelling - 100}}{18}\right)$

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