SOME BASIC CONCEPT OF CHEMISTRY

SUBSTANCE (MATTER)

Anything having some mass and volume is called as substance.

Substance
- Physical classification
  - Solid
  - Liquid
  - Gas
- Chemical classification
  - Pure substance
    - Element
      - metal
      - nonmetal
      - metalloid
    - Compound
  - Impure substance
    - Mixture
      - Homogenous
        - Solution
      - Heterogenous
        - Colloid
        - Suspension

- Avogadro's no.: $N_A = 6.022 \times 10^{23}$
- Mole : A collection of Avogadro's number of any particle is called as one mole of that particle.
- Atomic weight of an element $= \dfrac{\text{mass of one atom of the element}}{\frac{1}{12} \times \text{mass of one } C^{12} - \text{atom}}$
- Molecular weight $= \dfrac{\text{mass of its one molecule}}{\frac{1}{12} \times \text{mass of one } C^{12} - \text{atom}}$ = Sum of atomic masses of all atoms

  Example : Molecular weight of $H_2SO_4 = (1 \times 2) + (32 \times 1) + (16 \times 4) = 98$
- Moles of atom $= \dfrac{\text{No of atoms}}{\text{Avagadro's no.}} = \dfrac{\text{weight of elemental sample in gm}}{\text{Gram atomic weight}} = \text{No. of gram} - \text{atoms}$

  Example : For sample of oxygen having weights 1.6 gm

  No of moles of O atoms $= \dfrac{1.6}{16} = 0.10$; No of O-atoms $= 0.10 \times 6.022 \times 10^{23} = 6.022 \times 10^{22}$
- Moles of molecule $= \dfrac{\text{No. of molecules}}{\text{Avogadro's no.}} = \dfrac{\text{weight of sample in gm}}{\text{Gram molecular wt.}} = \text{No. of gram} - \text{molecules}$

Example : For a sample of 3.011×10^{24} H_3PO_4 molecules

$$\text{Moles of } H_3PO_4 = \frac{3.01 \times 10^{24}}{6.022 \times 10^{23}} = 5 \quad ; \text{ weight of } H_3PO_4 = 5 \times 98 = 490 \text{ gm}$$

- Avogadro's hypothesis : At constant temperature and pressure, volume of the gaseous sample is directly proportional to number of gaseous moles, present.

$$\text{Moles of molecules} = \frac{\text{volume of gas in lit. at S.T.P.}}{22.4 \text{ lit.}}$$

Example : Calculate moles of molecule of NH_3 having 2.24 litre at S.T.P.

Solution : $\text{Moles of } NH_3 = \dfrac{2.24}{22.4} = 0.10$

- Loschmidt number : Number of molecules in 1 ml of gas at S.T.P. It is 2.687×10^{19}.

- Equivalent Weight $:= \dfrac{\text{Molecular weight}}{\text{n} - \text{factor}}$

- Calculation of n-factor :

(i) For any acid it is the number H^+ replaced or e-pair gained by its one molecule, i.e. its basicity.

Example : $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

n-factor for $H_2SO_4 = 2$

$H_3PO_4 + NaOH \rightarrow NaH_2PO_4 + H_2O$

n-factor for $H_3PO_4 = 1$

$$H_3BO_3 + \ddot{N}H_3 \rightarrow \begin{array}{c} OH \\ | \\ HO-B-OH \\ \uparrow \\ NH_3 \end{array}$$

n-factor for $H_3BO_3 = 1$

(ii) For any base, it is the no of OH^- lost or H^+ gained or electron pair donated by its one molecule, i.e. its acidity.

Example : $HCl + NH_3 \rightarrow NH_4^+ + Cl^-$

n-factor for $NH_3 = 1$

$H_2SO_4 + Mg(OH)_2 \rightarrow MgSO_4 + 2H_2O$

n-factor for $Mg(OH)_2 = 2$

(iii) For any salt it is the total cation or anionic charge.

Example : For $Al_2(SO_4)_3$ n-factor = 6

(iv) For any radical, it is the charge over it.

Example : For SO_4^{2-} ; n-factor = 2

Solution : For PO_4^{3-} ; n-factor = 3

(v) For any oxidising agent it is no of electrons lost by it's one molecule.

Example : $K_2Cr_2O_7 \longrightarrow Cr^{3+}$

n-factor = 6

(vi) For any reducing agent it is the no of electrons gained by it's one molecule.

Example : $C_2O_4^{2-} \longrightarrow 2CO_2$

n-factor = 2

Gram equivalent $= \dfrac{\text{weight in gm}}{\text{Eq.wt.in gm}}$ = moles × n factor = Normality × Vol in lit.

Laws of equivalence : Gram equivalent of every reactant used is the same and it is equal to the gram equivalent of every product, formed.

Concentration units :

It represent amount of solute present in a definite amount of solution.

(i) Molarity (M) : Moles of solute in one lit solution $= \dfrac{\text{moles of solute}}{\text{vol. of solution in lit}} = \dfrac{w \times 1000}{M^0 \times V(ml)}$

(where M° is molar mass of solute)

$$\boxed{\text{Molarity} \times \text{volume of solution in lit} = \text{moles of solute}}$$

Example : 4g NaOH is present in 100 ml of its aqueous solution. What is the molarity.

(1) 2M (2) 1 M (3) 10 M (4) 0.1 M

Solution : Molarity $= \dfrac{w}{M^0} \times \dfrac{1000}{\text{volume(ml)}} = \dfrac{4}{40} \times \dfrac{1000}{100} = 1$

(ii) Normality (N) : No. of equivalent of solute per litre of solution.

$$= \frac{\text{g.eq. of solute}}{\text{vol. of solution in lit}} = \frac{w \times 1000}{E \times V(ml)}$$

⇒ [Normality × volume in lit = g.eq. of solute]

⇒ [Normality = Molarity × n - factor]

Example : Find the number of milliequivalent of H_2SO_4 persent in 50 ml of N/20 H_2SO_4

Solution : meq. = N × V (mL) = 1/20 × 50 = 2.5

(iii) Molality (m) = Moles of solute per kg of solvent

$$= \frac{\text{Moles of solute}}{\text{Weight of solvent(kg)}} = \frac{w \times 1000}{M^0 \times w'}$$

Example : 10g HCl dissolved in 250 mL of its aqueous solution. If density of the solution is 1.2 g/ml then molarity of the solution will be

(1) 1 (2) 0.34 (3) 0.945 (4) 3.4

Solution : Weight of solute = 10 g Volume of solution = 250 mL

Density of solution = 1.2 g/mL Weight of solution = 250 × 1.2 = 300 g

∴ Weight of solvent = weight of solution – weight of solute = 300 – 10 = 290 g

$$\therefore \quad m = \frac{w}{M^0} \times \frac{1000}{W} = \frac{10}{36.5} \times \frac{1000}{290} = 0.945$$

(iv) % Composition :

(a) w/w = wt of solute in gm per 100 gm of solution = $\dfrac{\text{wt. of solute in gm}}{\text{wt. of solution in gm}} \times 100$

Example : What is the weight percentage of NaCl solution in which 20g NaCl is dissolved in 60 g of water.

(1) 10% (2) 5% (3) 25% (4) 15%

Solution : Weight percentage of NaCl = $\dfrac{\text{weight of NaCl}}{\text{weight of solution}} \times 100$

$$= \frac{20}{20+60} \times 100 = 25\% \text{ NaCl solution (w / w)}$$

(b) v/v = vol. of solute in ml per 100 ml of solution = $\dfrac{\text{Vol. of solute in ml}}{\text{Vol. of solution in ml}} \times 100$

Example : A solution is prepared by mixing of 10 ml ethanol with 120 ml of methanol. What is volume percentage of ethanol.

(1) 10% (2) 7.7% (3) 20% (4) 15%

Solution :Volume percentage of ethanol = $\dfrac{\text{volume of ethanol}}{\text{volume of solution}} \times 100 = \dfrac{10}{10+120} \times 100 = 7.7\%$

(c) w/v = wt of solute in gm per 100 ml of solution = $\dfrac{\text{wt. of solute in gm}}{\text{vol. of solution in ml}} \times 100$

Example :What is weight / volume percentage of a solution in which 7.5 g of KCl is dissolved in 100 ml of the solution.

(1) 7.5% (2) 92.5% (3) 50% (4) none

Solution : 7.5% of KCl (w/V). 7.50 g KCl present in 100 mL of the solution.

$$\frac{7.5}{100} \times 100 = 7.5\%$$

NOTE :

(i) If for any solution percentage (w/w) is 'x', density is 'd' g/ml and molar mass of the soute is 'M_0' then

$$\text{Molarity} = \frac{10x \,.\, d}{M_0}$$

Example : The solution of H_2SO_4 contains 80% by mass. Specific gravity (density) of solution is 1.71 g/cc. Find its Molarity.

Solution : $M = \dfrac{10 \times d \times \text{percent}}{\text{GMM}}$ $\qquad M = \dfrac{10 \times 1.71}{98} \times 80 = 13.95$

(ii) For any pure liquid having density 'd' g/ml and molar mass M_0.

$$\text{Molarity} = \frac{1000\,d}{M_0}$$

$$\text{Molarity of pure water} = \frac{1000 \times 1}{18} = 55.55\,\text{M}$$

5. Strength (g/lit) = $\dfrac{\text{wt. of solute in gm}}{\text{Vol. of solution in lit}}$

Molarity × Mol Mass of solute = g/lit

Normality × Eq. wt of solute = g/lit.

Note :

(i) Due to dilution of any solution moles of solute, present in solution remain the same but concentration is decreased

M_1 V_1	M_2 V_2
Before dilution	after dilution

$M_1V_1 = M_2\,V_2$ = moles of solute.

Example : We have 100 ml of 0.1 M NaOH solution which in diluted till its concentration become 0.01 M. Calculate volume of H_2O added.

Solution : $M_1V_1 = M_2V_2 \;\Rightarrow\; V_2 = \dfrac{100 \times 0.1}{0.01} = 1000\,\text{ml}$

Vol. of water added = 1000 – 100 = 900 ml

(ii) If from any solution any small amount is taken out, concentration of the small amount remain exactly the same.

Molarity = M
Volume = 100ml
original solution

→ Molarity = M
Volume = 90ml

→ Molarity = M
Volume = 10ml

Molecular formula : It represent actual no of constituent atoms present in any molecule.

Empirical formula : It represent atoms present in any molecule in the simplest ratio.

Molecular formula : (Empirical formula) × n

LAWS OF CHEMICAL COMBINATION

1. The Law of Conservation of Mass (Lavoiser 1744) : This law states "matter can neither be created nor destroyed or in a chemical reaction, the mass of the reactants is equal to the mass of the products".The exception to this law is nuclear reactions where Einstein equation is applicable.

Example :

$$2Ca + O_2 \rightarrow 2CaO$$

80 g   32 g   112 g

Total mass reactant = Total mass product = 112 g

* For a complete irreversible reaction total mass of reactants before reaction in equal to the total mass of Products after reaction.

2. The Law of Constant Composition or Definite Proportion (Proust in 1799) : This law states that " All pure sample of the same chemical compound contain the same elements combined in the same proportion by mass irrespective of the method of preparation".

   Example : Different samples of carbon dioxide contain carbon and oxygen in the ratio of 3 : 8 by mass. Similarly in water ratio of weight of hydrogen to oxygen is 1 : 8.

3. The Law of Multiple Proportion (Dalton)

   This law states that :when two elements A and B combine together to from more than one compound, then several, masses of A which separately combine with a fixed mass of B, are in a simple ratio.

   Example :

CO	and		CO_2
12 : 16			12 : 32
ratio =	16 : 32	i.e.	1 : 2

4. The Law of Reciprocal Proportions (Richer in 1792 - 94)

   This law states that "when two elements combines separately with third element and form different types of molecules their combining ratio is directly reciprocated if they combine directly.

   Example : C combines with O to form CO_2 and with H to form CH_4. In CO_2 12 g of C reacts with 32g of O, whereas in CH_4 12 g of reacts with 4g of H. Therefore when O combines with H, they should combine in the ratio of 32 : 4 (i.e. 8 : 1) or in simple multiple of it. The same is found to be true in H_2O molecules.The ratio of weight of H and O in H_2O is 1 : 8.

5. The Law of Gaseous Volume (Gay Lussac in 1808)

   This law states that "when gas combine, they do so in volume which bear a simple ratio to each other and also to the product formed provided all gases are measured under similar conditions. Or in other words volume of reacting gases and product gases have a simple numerical ratio to one another.

   Example

   $$H_2(g) + Cl_2(g) = 2HCl(g)$$

   1 unit vol.   1 unit vol   2 unit vol.   ratio = 1 : 1 : 2

6. The Avogadro Law

   This law states that "equal voume of all gaseous under similar conditions of temperature and pressure contain equal number of molecules".

   $$2H_2 + O_2 \rightarrow 2H_2O$$

   2 vol.   1 vol.   2 vol.   ratio 2 : 1 : 2

   It provides a relationship between vapour density and molecular mass of substances.

   2 × vapour density (VD) = molecular mass of gas.