1. SOME BASIC CONCEPTS IN CHEMISTRY

1 OMQ + 2 VSAQ + 1SAQ [1 M + 2M + 2M+ 4M = 9 M]

CONCEPTS & DEFINITIONS

1.0 In this first chapter, we study some basic concepts of chemistry.

The main contents of this chapter are i)Importance of Chemistry ii) Nature of Matter iii) Properties of Matter iv) Scientific Notation & significant figures v) 5 Laws of Chemical combination vi) Dalton's atomic theory viii) Atomic and molecular masses ix) Mole concept and molar masses x) Percentage composition xi) Stoichiometric calculations.

1.1 **Importance of Chemistry:** The study of chemistry touches almost every sphere of life such as weather patters, functioning of brain and operation of a computer, production in chemical industries, manufacturing fertilisers, alkalis, acids, salts, dyes, polymers, drugs, soaps, detergents, metals, alloys etc.,

Chemistry plays a big role in meeting human needs for food, healthcare etc.,

Chemistry principles are useful to develop entities which control environmental spoilers like CFCs, Ozone depletion etc.,

1.2 **Nature of Matter:** Matter exists mostly in 3 physical states, solids, liquids and gases.

1.3 **Properties of matter** are divided into i) Physical properties like mass, volume, density, temperature, etc., ii) Chemical properties like acidity, basicity, combustibility etc.,

1.4 **Measurements:** In one way, chemistry is the subject of extremely big numbers and extremely small numbers as well.

For example, number of molecules in 2g of Hydrogen gas is 602,200,000,000,000,000,000,000.

Mass of 1 hydrogen atom is 0.00000000000000000000000166g.

Similarly, speed of light, charges of particles, Planck's constant etc., involve very high numbers.

In order to deal with such extreme numbers, a **scientific notation** is developed and the concept of **significant figures** is introduced.

1.4.1 **Scientific Notation(Exponential Notation):** Any number (how much big, small or medium) can be written in the form $N \times 10^n$ where N called the digital term which varies from 1.000.....to 9.999....and n is the exponent which is an integer that can be positive, negative or 0.

Ex 1: Scientific Notation of 232.508 is 2.32508×10^2.

Ex 2: Scientific Notation of 0.00016 is 1.6×10^{-4}.

Ex 3: Scientific Notation of 5 is 5.0×10^0.

1.4.2 **Significant figures** are the 'number of significant digits' in a value that are known to be reliable and certain. They express the precision of a measurement.

The process of finding the 'number of significant figures' in a given number involves inclusion or exclusion of certain digits according to certain given rules.

Ex : Significant figures in 1.6×10^{-4} is 2.

1.5 Five laws of chemical combination :

	Laws	Illustration (s)
1.	**Law of conservation of mass :** In a chemical reaction, the total mass of reactants is equal to the total mass of products. **Proposed by : LAVOISIER**	1.00g C + 5.34 g Sulphur = 6.34 g CS_2 $NaCl_{(aq)} + AgNO_{3\,(aq)} \rightarrow AgCl_{(s)} + NaNO_{3\,(aq)}$ Mass of ($NaCl + AgNO_3$)=Mass of ($AgCl + NaNO_3$)
2.	**Law of definite proportion :** Elements always combine in fixed ratio of their weights. A given chemical compound always contains the same elements in a fixed ratio by mass, irrespective of source and method of preparation. **Proposed by : (Joseph Proust)**	CO_2 can be prepared by (i) combining C with oxygen (ii) Heating lime stone. But, in any process, mass ratio of C and O in CO_2 is 12 : 32 = 3:8
3.	**Law of multiple proportion(Dalton):** If two elements form two or more compounds then the weight of one element which combines with the fixed weight of other element maintain a simple multiple ratio.	Nitrogen and oxygen form N_2O, NO, N_2O_3,NO_2, N_2O_5. Here, the weight of O which combine with 28 g of N are 16,32,48,64,80 which are in the ratio 1:2:3:4:5
4.	**Gay - Lussac's gaseous law of combining volumes :** Gases combine in simple whole number ratio of their volumes under similar conditions of temperature and pressure.	$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$ 2 L of H + 1 L of O $\rightarrow$ 2 L of water vapour Ratio of volumes of H, O, H_2O is 2:1:2.
5.	**Avogadro's law :** Equal volumes of different gases contain equal number of molecules under similar conditions of temperature and pressure.	$2H + O \rightarrow H_2O$ 2 vol of H gas + 1 vol of O $\rightarrow$ 2 vol of water vapour 2n molecules + n molecules $\rightarrow$ 2n molecules

1.6 Dalton's Atomic Theory:

i) Matter consists of indivisible particles called atoms.

ii) Atoms of different elements differ in mass.

iii) Compounds are formed when atoms of different elements combined in a fixed ratio.

iv) Atoms are neither created nor destroyed but are reorganised in a chemical reaction.

1.7 Atomic and Molecular Masses:

1.7.1 Atomic mass: The mass of an atom is simply called atomic mass and it is measured relatively.

1.7.2 Atomic mass unit (a.m.u): Atomic masses are measured with respect to ^{12}C.

One a.m.u is defined as mass exactly equal to $1/2^{th}$ mass of one C-12 atom.

$$1\,amu = \frac{1}{12}\,\text{mass of }{}^{12}C = 1.66\times10^{-24}\,g$$

Ex : Mass of H atom = 1.008 amu=1.6736×10^{-24} g

Note: 'amu' is simply denoted by 'u'- unified mass

1.7.3 Molecular Mass: Molecular mass is the sum of atomic masses of the atoms of all the elements present in the molecule.

Ex: Molecular mass of water (H_2O) = 2(1.008 amu) + 16.00amu = 18.02 amu.

Ex: Molecular mass of methane (CH_4) = 12.011 u + 4(1.008 u) = 16.043 u.

1.8 Mole concept and molar masses:

1.8.1 Mole: One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g of the ^{12}C isotope.

The mass of ^{12}C atom is 1.992648×10^{-23} g.

The mass of one mole of ^{12}C is 12g.

$$\therefore \text{No. of atoms in one mole of }{}^{12}C = \frac{12\,g}{1.992648\times10^{-23}\,g} = 6.0221367\times10^{23} \cong 6.022\times10^{23}$$

The above number 6.022×10^{23} is called Avagadro Number (N_A).

Thus, the collection of 6.022×10^{23} molecules of an element or ions or a compound makes 1mole of that element or ions or compound.

The number of moles (n) in 'g' grams of a substance, with molecular weight M is $n = \frac{g}{M}$

Note: The unit of mole is denoted by 'mol'

1.9 Percentage composition :

Mass Percent (%):

Mass % of element is the Mass of element present in 100 grams of the sample of a compound.

Formula: Mass % of an element = $\frac{\text{mass of the element}}{\text{Molar mass}}\times100$

Ex: Percentage by weight of carbon in $CH_4 = \frac{12}{16}\times100 = 75\%$

1.9.1 Empirical formula: The formula showing the relative number of atoms of different elements present in one molecule of a compound is called emperical formula.

1.9.2 Molecular formula: The formula which represents the exact no. of atoms of each element present in one molecule of the substance is called molecular formula.

Ex: Molecular formula of ethane is C_2H_4; and its empirical formula is CH_2

1.9.3 Relations between empirical formula and molecular formula :

Molecular formula = n × (Empirical formula), n is an integer.

$$n = \frac{\text{Molecular weight}}{\text{Empirical formula weight}}$$

Molecular weight = n × Empirical formula weight

Note: Molecular weight = 2 × Vapour density

1.10 Stoichiometric Calculations: Stoichiometry deals with the calculation of masses (sometimes volumes) of the reactants and the products involved in a chemical reaction.

1.11 Various concentration methods of solute in a solution:

i) Mass Percent (%):

Mass % of solute is the mass of solute present in 100 grams of the solution.

Formula: Mass percentage of solute $= \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$

ii) Molarity (M):

Molarity of a solution is the number of 'moles of solute' dissolved in **one litre of solution**

Formula: Molarity $M = \frac{W}{GMW} \times \frac{1000}{V(mL)} = \frac{\text{No.of moles of solute}}{\text{Volume of solution in L}}$

iii) Molality(m):

Molality of a solvent is the number of 'moles of solute' present in **one kilogram of solvent**.

Formula: Molality $m = \frac{W}{GMW} \times \frac{1000}{W(g)} = \frac{\text{No.of moles of solute}}{\text{Mass of solvent in kg}}$

iv) Normality (N):

Normality is the number of gram equivalents of solute present in one litre of solution

Formula: Normality $N = \frac{W}{GEW} \times \frac{1}{V(L)}$

v) Mole Fraction(χ):

Mole fraction of a component is the ratio of the number of 'moles of the component' to the 'total number of moles of all the components' present in the solution.

Formula: Mole fraction $\chi_A = \frac{\text{No. of moles of component A}}{\text{Total no. of moles of all the components}} = \frac{n_A}{n_A + n_B}$

Imp. Formulae

1. Number of moles $n = \frac{W}{GMW} = \frac{\text{Weight in g}}{\text{Gram Molecular Weight}}$

2. Number of molecules $= n \times N_A = \text{Number of moles} \times 6.022 \times 10^{23}$

3. $n = \frac{PV}{RT}$ [From Ideal gas equation PV = nRT]

4. $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ [From ideal gas equation]

5. $\frac{V_1M_1}{n_1} = \frac{V_2M_2}{n_2}$ [From volumetric analysis principle]

6. $n = \frac{\text{Molecular weight}}{\text{Empirical formula weight}}$

7. Molecular Formula = (Empirical formula) × n

8. Mass % of an element $= \frac{\text{mass of the element}}{\text{Molar mass}} \times 100$

9. Mass % of solute $= \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$

10. Molarity $M = \frac{W}{GMW} \times \frac{1000}{V(mL)} = \frac{W}{GMW} \times \frac{1}{V(L)}$

11. Molality $m = \frac{W}{GMW} \times \frac{1000}{W(g)} = \frac{W}{GMW} \times \frac{1}{W(kg)}$

12. Normality $N = \frac{W}{GEW} \times \frac{1}{V(L)}$

13. Mole fraction $\chi_A = \frac{n_A}{n_A + n_B}$

14. Molar mass (or) Gram Molar weight (GMW) of certain important compounds:

$H_2 = 2g$; $N_2 = 28$ g ; $NH_3 = 17$ g ; $HCl = 36.5$ g; $H_2SO_4 = 98$ g; $CH_2O = 30$ g;

$CaCO_3 = 100g$; $Na_2CO_3 = 106$ g; $Na_2SO_4 = 142g$;

Glucose $C_6H_{12}O_6 = 180$ g; Sugar $C_{12}H_{22}O_{11} = 342$

☛ For **Chem Beats** of this chapter Refer Page **Nos. 196 to 198**