4. STOICHIOMETRY

1 VSAQ+ 1SAQ = 6 MARKS

CONCEPTS & DEFINITIONS

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.

2. Five laws of chemical combination :

	Laws	Illustration (s)
2.1	**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 ($AgCl + NaNO_3$)
2.2	**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
2.3	**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
2.4	**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.
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

3. Measuring the masses of atoms/ molecules;Number of molecules in a substance:

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

3.1.1 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.

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

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

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

3.1.2 Average atomic mass: Average atomic mass is calculated by taking into account, the existance of isotopes of elements and their relative abundance (percent occurance) in nature.

3.2.1 Molecular mass/ weight: Molecular mass of a compound is the sum of the atomic masses of all the atoms 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 amu + 4(1.008 amu) = 16.043 amu.

3.2.2 Formula mass: The formula mass of an ionic substance is the sum of the atomic masses of all the atoms present in the molecule.

Ex: For Na*Cl* (Na^+Cl^-) ,we use formula mass instead of molecular mass.

Formula mass of Na*Cl*= Atomic mass of Na^+ atomic mass of C*l* = 23.0 + 35.5 = 58.5 u

3.3.1 Gram - atom: One gram atomic weight of a substance is called one gram atom.

Ex: Atomic weight of oxygen = 16 amu.
One gram atom of oxygen = 16 g of oxygen
Gram atomic weight of oxygen = 16 g

Formula: No. of gram atoms = $\dfrac{\text{Weight of substance in g}}{\text{Gram atomic weight}}$

3.3.2 Gram-molecule: One gram molecular weight of a substance is called one gram molecule.

Ex: Molecular weight of oxygen = 32 amu.
One gram molecule of oxygen = 32 g of oxygen
Gram molecular weight of oxygen = 32 g

3.3.3 Gram molecular mass: Molecular mass of a substance, expressed in grams is called its gram molecular mass. The amount of the substance is also called one gram molecule.

Ex: Molecular mass of $CaCO_3$ = 100 amu.
Gram molecular mass of $CaCO_3$ = 100 g

Formula: No. of gram molecules = $\dfrac{\text{Weight of substance in g}}{\text{Gram molecular weight}}$

3.3.4 Molar volume: One gram molecular weight of any gaseous substance at STP occupies 22.4L volume. This value is known as molar volume. (or) Gram molar volume.

3.4 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 x 10^{-23} g.

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

$$\therefore \text{No. of atoms in one mole of } ^{12}C = \frac{12\,g}{1.992648\times10^{-23}\,g} = 6.0221367\times10^{23} \cong 6.023\times10^{23}$$

The above number 6.023 x 10^{23} is called Avagadro Number (N_A).

Thus, the collection of 6.023 x 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 $n = \dfrac{g}{M}$

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

3.5 Molar mass : The mass of one mole of any substance in grams, is called its molar mass.

Ex: Molar mass of H_2SO_4 = 98 g.

4. Empirical and molecular formulae, Percentage compositions of compounds:

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

4.2 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

4.3 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

4.4 Percentage weight of an element : The weight of element present in 100 grams of the sample of a compound is called percentage weight of that element.

Formula: Percentage by weight of the element = $\frac{\text{weight of the element}}{\text{Molecular weight}} \times 100$

Ex : Percentage by weight of carbon in $CH_4 = \frac{12}{16} \times 100 = 75\%$

5. Oxidation - Reduction concepts :

5.1 Oxidation[LEO]: **L**oss of **E**lectrons is **O**xidation (or) Addition of oxygen is called oxidation.

5.2 Reduction [GER]: **G**ain of **E**lectrons is **R**eduction(or) Removal of oxygen is called reduction.

5.3 Redox reaction: It is the chemical reaction in which both oxidation and reduction occurs simultaneously.

5.4 Oxidation number: It is the number which indicates charge (or) possible charge that appears on an atom in a molecule (as per certain arbitrary rules) . It denotes the oxidation state of an element in a compound ascertained according to a set of rules. The rules are formulated on the basis that electron pair in a covalent bond belongs entirely to more electro negative element.

5.5 Oxidising agent: It is an acceptor of electrons. It increases the oxidation number of an element in a substance.

5.6 Reducing agent: It is a donor of electrons. It lowers the oxidation number of an element in a substance.

6. Types of Redox reactions:

i) Combination reactions: The redox reactions, in which two or more elements combine to form a new compound. These are in the form A+B→AB

Ex : $\overset{0}{C}_{(s)} + \overset{0}{O}_2 \rightarrow \overset{+4}{C}\,\overset{-2}{O}_{2(g)}$

ii) Decomposition reactions: The redox reactions, in which the chemical compounds chemically split into two or more simple substances. These are reverse of combination reactions. These are in the form AB → A + B.

Ex : $2\overset{+1}{H}_2\overset{-2}{O}(g) \rightarrow 2\overset{0}{H}_2(g) + \overset{0}{O}_2(g)$

iii) Displacement reactions : The redox reactions, in which an ion (atom) in a compound is replaced by an ion (atom) of another element. These are in the form A + BC $\rightarrow$ AC + B

Ex 1 : $\overset{0}{Zn}_{(s)} + \overset{+2}{Cu}\overset{-2}{SO}_{4(aq)} \rightarrow \overset{+2}{Zn}\overset{-2}{SO}_{4(aq)} + \overset{0}{Cu}_{(s)}$ (Metal displacement)

Ex 2 : $2\overset{0}{Na}(s) + 2\overset{+1}{H_2}\overset{-2}{O}(l) \rightarrow 2\overset{+1}{Na}\overset{-2}{O}\overset{+1}{H}(aq) + \overset{0}{H_2}(g)$ (Non-Metal displacement)

iv) Disproportionation reactions : The redox reactions, in which an element in one oxidation state is simultaneously oxidised and reduced.

Ex : $Cl_{2_{(g)}} + 2OH^-_{(aq)} \rightarrow \left[Cl^+O\right]^-_{aq} + Cl^{-1}_{(aq)} + H_2O$

7. Balancing of Redox Reactions is done in 2 ways.

a) Oxidation number method   b) Half reduction method

STEPS IN THE BALANCING OF REDOX EQUATION (IN ACIDIC MEDIUM)

STEP 1: Write the given skeletal ionic equation.

STEP 2: Write Oxidation half reaction and Reduction half reaction.

STEP 3: Balance atoms other than O and H.

STEP 4: Balance Oxygen atoms.

STEP 5: Balance Hydrogen atoms.

STEP 6: Balance charges.

STEP 7: Equalise electrons.

STEP 8: Apply Criss cross rule, if required.

STEP 9: Add the two Half reactions.

STEP 10: Write the Balanced Equation.

☛ For **Chem Beats** of this chapter Refer Page **No. 245 to 246**

Imp. Formulae

1. Number of moles $n = \dfrac{\text{Mass}}{\text{Molar mass}}$

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

3. $n = \dfrac{\text{Molecular weight}}{\text{Empirical formula weight}}$

4 Molecular Formula = (Empirical formula) $\times$ n

5. 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;