CHAPTER 1- SOME BASIC CONCEPTS IN CHEMISTRY CLASS XI

Calculate the number of atoms in each of the following:

(i) 52 moles of Ar

(ii) 52 u of He

(iii) 52 g of He.

Calculate the number of moles in each of the following:

i. 49 grams of sulphuric acid
ii. 44.8 litres of carbon dioxide at STP
iii. 3.011×10^{23} molecules of oxygen

(i) 1 mole of Ar contains 6.022×10^{23} atoms

52 moles of Ar contains $52 \times 6.022 \times 10^{23}$
$= 3.13 \times 10^{25}$ atoms

(ii) 4 u of He = 1 atom

52 u of He $= \frac{1}{4} \times 52 = 13$ atoms

(iii) 4 g of He contains 6.022×10^{23} atoms

52 g of He contains $\frac{6.022 \times 10^{23}}{4} \times 52$
$= 7.83 \times 10^{24}$ atoms

i. 1 mole of H_2SO_4=98g
x mole = 49g
x=0.5moles

ii. 1mole of gas at STP=22.4L
x mole= 44.8L
x= 2moles

iii. 1mole of oxygen = 6.022×10^{23}molecules
x moles = 3.011×10^{23}molecules
x = 0.5moles

Questions

(i) Calculate the mass of a carbon atom in Kg

(ii) Calculate the mass of CO_2 that contains same number of molecules as there are in 3.4 g of ammonia, NH_3

Percentage composition

The mass percentage of each constituent element present in any compound is called its percentage composition

$$\% \text{ mass of element} = \frac{\text{mass of element}}{\text{total mass of compound}} \times 100$$

Q. Calculate the mass percent of different elements present in sodium sulphate (Na_2SO_4).
Solution: Molecular mass of Na_2SO_4 = 142 u.

$$\text{Mass \% of sodium} = \frac{2 \times 23}{142} \times 100 = \frac{46}{142} \times 100 = 32.39\ \%$$

$$\text{Mass \% of sulphur} = \frac{32}{142} \times 100 = 22.53\ \%$$

$$\text{Mass \% of oxygen} = \frac{4 \times 16}{142} \times 100 = 45.07\ \%$$

Molecular Formula

The **molecular formula** shows the exact number of different types of atoms present in a molecule of a compound. E.g. C_6H_6 is the molecular formula of benzene.

Empirical Formula

An **empirical formula** represents the simplest whole number ratio of various atoms present in a compound. E.g. CH is the empirical formula of benzene.

Compound	Molecular Formula	Empirical Formula
Water	H_2O	H_2O
Hydrogen peroxide	H_2O_2	HO
Ethane	C_2H_6	CH_3
Glucose	$C_6H_{12}O_6$	CH_2O

Relationship between empirical and molecular formulae: **MF= n (EF)**

Vapour density-The ratio of the density of a gas to the density of hydrogen at the same temperature and pressure. **Molar mass = 2 × vapour density**

Q. Calculate the empirical formula of a compound containing 54.54 % C, 9.09 % H and rest oxygen. Also calculate the molecular formula ,if its molar mass is 88 u.

Element	%	Relative number of moles	Simple ratio
C	54.54	54.54/12=4.53	4.53/2.27=2
H	9.09	9.09/1=9.09	9.09/2.27=4
O	36.36	36.36/16=2.27	2.27/2.27=1

$EF=C_2H_4O$

EF mass= 12x2 + 1x4 + 16x1=44

n=Molar mass/EF mass =44/44=1

MF=n x (EF)= 2 x (C_2H_4O)

$MF=C_4H_8O_2$

Questions

(i) An organic compound on analysis was found to contain 57.8% of carbon, 3.6% of hydrogen and the rest oxygen. If its vapour density is 83 find its empirical and molecular formula.

(ii) An organic compound on analysis was found to contain1.8 g of carbon, 0.6 g of hydrogen and 2.4 g of oxygen. Find its empirical formula.

Q. A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Calculate
(i) empirical formula,
(ii) molar mass of the gas, and
(iii) molecular formula.

Element	Amount	Ratio	Simple ratio
C	0.92/12	0.076	1
H	0.076/1	0.076	1

44g of CO_2 contains 12g carbon
3.38 g of CO_2 will contain xg of C
x=0.92g

18 g of H_2O contains 2g H
0.690g of H_2O will contain xg of H
x=0.076g

EF=CH EF mass= 12+ 1=13

10L weighs 11.6g
22.4L will weigh xg x=26g [molar mass]

n=EF mass/Molar mass=26/13=2

MF=n(EF)= 2(CH) thus: MF=C_2H_2

Stoichiometry

$$CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2\,H_2O\,(g)$$

- One mole of CH_4 (g) reacts with two moles of O_2 (g) to give one mole of CO_2 (g) and two moles of H_2O(g)
- One molecule of CH_4 (g) reacts with 2 molecules of O_2 (g) to give one molecule of CO_2 (g) and 2 molecules of H_2O(g)
- 22.4 L of CH_4 (g) reacts with 44.8 L of O_2 (g) to give 22.4L of CO_2 (g) and 44.8 L of H_2O(g)
- 16 g of CH_4 (g) reacts with 64 g of O_2 (g) to give 44 g of CO_2 (g) and 36 g of H_2O (g).

Q. Calculate the amount of water (g) produced by the combustion of 16 g of methane.
Solution : 16 g of CH_4 corresponds to one mole.
1 mol of CH_4 (g) gives 2 mol of $H_2\,O$ (g)
16 g of CH_4 = 2 × 18 = 36 g
Therefore 36 g $H_2\,O$

Q. How many moles of methane are required to produce 22g CO_2 (g) after combustion?
Solution : 16g methane gives 44g of CO_2
xg methane will give 22g of CO_2
x=8g ie 8/16moles=0.5moles

Q. How much copper can be obtained from 100 g of copper sulphate ($CuS0_4$) ?

Solution:

Molar mass of $CuSO_4$ = 63.5 + 32 + 4 × 16 =63.5 + 32 + 64 = 159.5 g

159.5 g of $CuSO_4$ contains copper = 63.5 g

100 g of $CuSO_4$ contains copper = x g

x= 39.81 g

Q. Chlorine is prepared in the laboratory by treating manganese dioxide ($Mn0_2$) with aqueous hydrochloric acid according to the reaction

$4HCI_{(aq)} + MnO_{2(s)} \rightarrow 2H_2O_{(I)} + MnCI_{2(aq)} + Cl_{2(g)}$

How many grams of HCI react with 5.0 g of manganese dioxide? [Mn = 55, O=16]

Solution:

$$\underset{\substack{\text{4 mole}\\ \text{or 146 g}}}{4HCl_{(aq)}} + \underset{\substack{\text{1 mole}\\ \text{or 87 g}}}{MnO_{2(s)}} \longrightarrow 2H_2O_{(l)} + MnCl_{2(aq)} + Cl_{2(g)}$$

87 g of MnO_2 reacts with HCl = 146 g

$$5 \text{ g of } MnO_2 \text{ reacts with HCl} = \frac{146 \times 5}{87} = 8.39$$

$$\simeq 8.40 \text{ g}$$

Limiting Reagent

The reactant which gets consumed completely or limits the amount of product formed is known as limiting reagent.

Q. Calculate the amount of carbon dioxide that could be produced when

(i) 1 mole of carbon is burnt in air.

(ii) 1 mole of carbon is burnt in 16 g of dioxygen.

(iii) 2 moles of carbon are burnt in 16 g of dioxygen.

Solution:

(i) Hence, 1 mole of C produces 44 g of CO_2

$$\underset{\substack{1\ \text{mole} \\ (12\ \text{g})}}{C} + \underset{\substack{1\ \text{mole} \\ (32\ \text{g})}}{O_2} \longrightarrow \underset{\substack{1\ \text{mole} \\ (44\ \text{g})}}{CO_2}$$

(ii)Hence, O_2 is the limiting reagent.

32 g O_2 reacts with C to produce 44 g of CO_2

16 g O_2 reacts with C to produce

$$\underset{12\ \text{g}}{C} + \underset{32\ \text{g}}{O_2} \longrightarrow \underset{44\ \text{g}}{CO_2}$$

(iii) 64 g O_2 reacts with C to produce 88 g of CO_2

16 g O_2 reacts with C to produce = xg

x=22g

$$\underset{24\ \text{g}}{2C} + \underset{64\ \text{g}}{2O_2} \longrightarrow \underset{88\ \text{g}}{2CO_2}$$

Q. Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation:

$N_2(g) + H_2(g) \rightarrow 2NH_3(g)$

(i) Calculate the mass of ammonia produced if 2.00×10^3 g dinitrogen reacts with 1.00×10^3 g of dihydrogen.

(ii) Will any of the two reactants remain unreacted?

(iii) If yes, which one and what would be its mass?

$$\text{Moles of } N_2 = \frac{2.00 \times 10^3}{28} = 71.43,$$

$$\text{Moles of } H_2 = \frac{1.00 \times 10^3}{2} = 500$$

1 mole of N_2 required 3 moles of H_2 from above equation.

$\therefore$ 71.43 mole of N_2 will require 3×71.43 = 214.29 mole of H_2

But moles of H_2 actually present = 500 moles

Thus, H_2 is in excess and will remain unreacted and N_2 is limiting reagent.

(i) 1 mole of N_2 reacts with H_2 to form NH_3 = 2 moles

71.43 moles of N_2 react with H_2 to form

$$NH_3 = \frac{2}{1} \times 71.43 = 142.86 \text{ moles}$$

Mass of NH_3 produced = 142.86×17 = 2428.6 g

(ii) Hydrogen will remain unreacted.

(iii) Moles of H_2 remaining unreacted = 500 – 214.29 = 285.71 moles

Mass of H_2 left unreacted = 285.71×2 = 571.42 g

Reactions in Solutions

Concentration

The concentration of a solution can be expressed in the following ways-

1. Mass Percent
2. Volume percent
3. Molarity
4. Molality
5. Mole Fraction

1. Mass Percent is the mass of the solute in grams per 100 grams of the solution.

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

2. Volume percent is the number of units of volume of the solute per 100 units of the volume of solution.

$$\text{Volume \% of the solute} = \frac{\text{Volume of the solute}}{\text{Volume of the solution}} \times 100$$

3. **Molarity**

- the number of moles of solute dissolved per litre of the solution.

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{volume of solution (L)}} = \frac{W}{M \times V}\ \text{mol/L}$$

- It is denoted by the symbol M.
- Measurements in Molarity can change with the change in temperature because solutions expand or contract accordingly
- To calculate the volume of a definite solution required to prepare solution of other molarity, the following equation is used

$$\boxed{\mathbf{M_1V_1 = M_2V_2}}$$

where M_1= initial molarity, M_2= molarity of the new solution,
V_1=initial volume and V_2= volume of the new solution.

- When density and molar mass is known molarity is calculated as

$$M = \frac{\% \times d \times 10}{\text{mass}}$$

Q. Calculate the molarity of NaOH in the solution prepared by dissolving its 4 g in enough water to form 250 mL of the solution.

$$M = \frac{W}{M \times V}\ \text{mol/L} = \frac{4 \times 1000}{40 \times 250} = 0.4\ \text{mol/L}$$

Q. Commercially available concentrated HCl contains 38%HCl by mass. What is its molarity, if its density is 1.19g/cm^3? [H=1,Cl=35.5]

Solution:

M = %.d.10/Molar mass

=38x1.19x10/36.5

= 12.38M

Q. Calculate the molarity of 30 mL of 0.5 M H_2SO_4 when diluted to 500 mL

Solution:

$\mathbf{M_1V_1 = M_2V_2}$

0.5 x 30/1000=M_2 x 500/1000

M_2= 0.03 moles /L

4.Molality

- the number of moles of solute dissolved per 1000 g (1 kg) of solvent.
- Molality is expressed as 'm'.
- It is independent of temperature as mass of solvent is independent of temperature.

$$\text{molality} = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}} \qquad \text{Molality} = \frac{\text{Moles of the solute}}{\text{Wt. of Solvent (in gm)}} \times 1000$$

Q. What is the molality of a solution containing 5.0 g NaCl dissolved in 25.0 g water?

Solution:

$$\text{Molality} = \frac{\text{Moles of the solute}}{\text{Wt. of Solvent (in gm)}} \times 1000 \qquad m = \frac{5 \times 1000}{58.5 \times 25} = 3.42 \text{ moles/kg}$$

Q. Calculate the mass of urea (NH_2CONH_2) required in making 2.5 kg of 0.25 molal aqueous solution.[Atomic mass of N=14,O=16,C=12,H=1]

Solution:

$$\text{molality} = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}} \qquad 0.25 = \frac{\text{moles}}{2.5} \qquad \text{moles} = 0.625 \text{ mole}$$

Mass of urea = moles x molar mass = 0.625 (60) = 37.5g

5. Mole Fraction

- The ratio of number of moles of one component to the total number of moles of all the components present in the solution.
- It is expressed as 'χ'.
- It has no units.
- $\chi_A + \chi_B = 1$

$$\text{Mole fraction of A} = x_A = \frac{n_A}{n_A + n_B}$$

$$\text{Mole fraction of B} = x_B = \frac{n_B}{n_A + n_B}$$

Q.A solution is prepared by mixing 25.0 g of water, H_2O, and 25.0 g of ethanol, C_2H_5OH. Determine the mole fractions of each substance.

Solution:

Moles of H_2O : 25.0 g / 18.0 g/mol = 1.34 mol

moles of C_2H_5OH : 25.0 g / 46.07 g/mol = 0.543 mol

χ_{H2O} : 1.34 mol / (1.34 mol + 0.543 mol) = 0.71

χ_{C2H5OH} : 0.543 mol / (1.34 mol + 0.543 mol) = 0.29

Q. A tank is charged with a mixture of 1.0×10^3 mol of oxygen and 4.5×10^3 mol of helium. Calculate the mole fraction of each gas in the mixture.

Solution:

Mole fraction can be calculated as

$\chi_{He} = 4.5 \times 10^3 \text{ mol} / (4.5 \times 10^3 \text{mol} + 1.0 \times 10^3 \text{ mol})$

$\chi_{He} = 4.5 \text{ mol} / 5.5 \text{ mol}$

$\chi_{He} = 0.82$

$\chi_{O2} = 1.0 \times 10^3 \text{ mol} / (4.5 \times 10^3 \text{ mol} + 1.0 \times 10^3 \text{ mol})$

$\chi_{O2} = 1.0 \times 10^3 / 5.5 \times 10^3$

$\chi_{O2} = 0.18$