

## Fill Ups of Some Basic Concepts of Chemistry

**Q.1. The modern atomic mass unit is based on ..... (1980)**

**Ans.** Carbon-12.

**Sol.** Carbon-12.

**Q.2. The total number of electrons present in 18 ml of water is ..... (1980)**

**Ans.**  $6.02 \times 10^{24}$

**Sol.**  $6.02 \times 10^{24}$

18 ml  $\text{H}_2\text{O}$  = 18 g  $\text{H}_2\text{O}$  ( $\because$  density of water = 1 g/cc)

= 1 mole of  $\text{H}_2\text{O}$ .

1 Mole of  $\text{H}_2\text{O}$  =  $10 \times 6.02 \times 10^{23}$  electrons

( $\because$  Number of electrons present in one molecule of water

=  $2 + 8 = 10$ )

=  $6.02 \times 10^{24}$  electrons

**Q.3. 3 g of a salt of molecular weight 30 is dissolved in 250 g of water. The molality of the solution is ..... (1983 - 1 Mark)**

**Ans.** 0.4m

**Sol. TIPS/Formulae :**  $\text{Molality} = \frac{\text{Moles of solute}}{\text{Mass of solvent in kg}}$

$$= \frac{\text{wt. of solute in gram} / \text{M. wt. of solute}}{\text{Mass of solvent in kg}}$$

$$\text{Molality} = \frac{3/30}{250/1000} = 0.4\text{m}$$



**Q.4. The weight of  $1 \times 10^{22}$  molecules of  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  is ..... (1991 - 1 Mark)**

**Ans.** 4.14 g

**Sol.** TIPS/Formulae : 1 Mole =  $6.023 \times 10^{23}$  molecules = Molecular weight in gms.

Weight of  $6.023 \times 10^{23}$  (Avogadro's number) molecules of  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  = Molecular wt. of  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  = 249 g.

$\therefore$  Weight of  $1 \times 10^{22}$  molecules of  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

$$= \frac{249 \times 1 \times 10^{22}}{6.023 \times 10^{23}} = 4.14 \text{ g}$$

**Q.5. The compound  $\text{YBa}_2\text{Cu}_3\text{O}_7$ , which shows super conductivity, has copper in oxidation state....., assume that the rare earth element yttrium is in its usual + 3 oxidation state. (1994 - 1 Mark)**

**Sol.** NOTE : Sum of oxidation states of all atoms (elements) in a neutral compound is zero.

TIPS/Formulae : As  $\text{YBa}_2\text{Cu}_3\text{O}_7$  is neutral.  $(+3) + 2(+2) + 3(x) + 7(-2) = 0$  or  $3 + 4 + 3x - 14 = 0$

$$\Rightarrow 3x + 7 - 14 = 0 \text{ or } x = +\frac{7}{3}$$



## Integar Type ques of Some Basic Concepts of Chemistry

Q. 1. A student performs a titration with different burettes and finds titre values of 25.2 mL, 25.25 mL, and 25.0 mL. The number of significant figures in the average titre value is (2010)

**Ans. Sol.** The least significant figure in titre values is 3.

$$\text{Average titre value} = \frac{25.2 + 25.25 + 25}{3} = \frac{75.4}{3} = 25.1$$

The number of significant figures in average titre value will also be 3.

Q. 2. Silver (atomic weight =  $108 \text{ g mol}^{-1}$ ) has a density of  $10.5 \text{ g cm}^{-3}$ . The number of silver atoms on a surface of area  $10^{-12} \text{ m}^2$  can be expressed in scientific notation as  $y \times 10^x$ . The value of x is : (2010)

**Ans. Sol.**  $d = \frac{\text{mass}}{V}$

$\Rightarrow 10.5 \text{ g/cc}$  means in 1 cc  $\Rightarrow 10.5 \text{ g}$  of Ag is present.

$$\text{Number of atoms of Ag in 1 cc} \Rightarrow 1 \frac{10.5}{108} \times N_A$$

$$\text{In 1 cm, number of atoms of Ag} = \sqrt[3]{\frac{10.5}{108} N_A}$$

$$\text{In 1 cm}^2, \text{ number of atoms of Ag} = \left(\frac{10.5}{108} N_A\right)^{2/3}$$

In  $10^{-12} \text{ m}^2$  or  $10^{-8} \text{ cm}^2$ , number of atoms of Ag

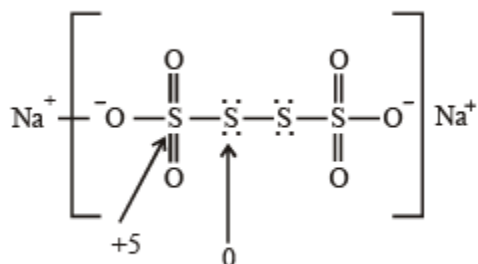
$$= \left(\frac{10.5}{108} N_A\right)^{2/3} \times 10^{-8} = \left(\frac{10.5 \times 6.022 \times 10^{23}}{108}\right)^{2/3} \times 10^{-8} = 1.5 \times 10^7$$

Hence  $x = 7$

Q. 3. The difference in the oxidation numbers of the two types of sulphur atoms in  $\text{Na}_2\text{S}_4\text{O}_6$  is (2011)



**Ans. Sol.**



Difference in oxidation number =  $5 - 0 = 5$

**Q. 4.** If the value of Avogadro number is  $6.023 \times 10^{23} \text{ mol}^{-1}$  and the value of Boltzmann constant is  $1.380 \times 10^{-23} \text{ J K}^{-1}$ , then the number of significant digits in the calculated value of the universal gas constant is (JEE Adv. 2014)

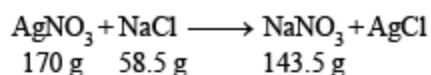
**Ans. Sol.**  $R = N_A \times k = 6.023 \times 10^{23} \times 1.380 \times 10^{-23}$   
 $= 8.312$  which has 4 significant figures



## Subjective questions of Some Basic Concepts of Chemistry (Part -1)

**Q. 1.** What weight of AgCl will be precipitated when a solution containing 4.77 g of NaCl is added to a solution of 5.77 g of AgNO<sub>3</sub>? (1978)

**Ans. Sol.** TIPS/Formulae : Write the balance chemical equation and use mole concept for limiting reagent.



From the given data, we find AgNO<sub>3</sub> is limiting reagent as NaCl is in excess.

∴ 170.0 g of AgNO<sub>3</sub> precipitates AgCl = 143.5 g

∴ 5.77 g of AgNO<sub>3</sub> precipitates AgCl

$$= \frac{143.5}{170.0} \times 5.77 = 4.87 \text{ g}$$

**Q. 2.** One gram of an alloy of aluminium and magnesium when treated with excess of dil. HCl forms magnesium chloride, aluminium chloride and hydrogen. The evolved hydrogen, collected over mercury at 0°C has a volume of 1.20 litres at 0.92 atm. pressure. Calculate the composition of the alloy. [H = 1, Mg = 24, Al = 27] (1978)

**Ans. Sol.** TIPS/Formulae : (i) Find volume of H<sub>2</sub> at N.T.P. (ii) Total amount of H<sub>2</sub> liberated = H<sub>2</sub> liberated by Mg & HCl + H<sub>2</sub> liberated by Al & HCl.

Conversion of volume of H<sub>2</sub> to N.T.P

Given conditions

N.T.P conditions

P<sub>1</sub> = 0.92 atm.

P<sub>2</sub> = 1 atm.

V<sub>1</sub> = 1.20 litres

V<sub>2</sub> = ?

T<sub>1</sub> = 0 + 273 = 273 K

T<sub>2</sub> = 273 K

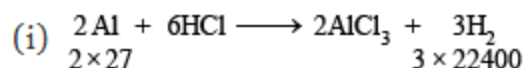


Applying ideal gas equation,  $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

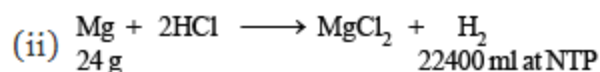
$$\frac{0.92 \times 1.20}{273} = \frac{1 \times V_2}{273}, \quad V_2 = \frac{0.92 \times 1.20 \times 273}{273 \times 1} \text{ litres}$$

$$= 1.104 \text{ litres} = 1104 \text{ ml}$$

The relevant chemical equations are



$$= 54 \text{ g} \qquad \qquad \qquad = 67200 \text{ ml at NTP}$$



Wt. of alloy = 1 g Let the wt. of aluminium in alloy = x g

∴ Wt. of magnesium in alloy = (1 - x) g According to equation

$$(i) \quad 54 \text{ g of Al} = 67200 \text{ ml of H}_2 \text{ at N.T.P} \therefore x \text{ g of Al} = 6 \frac{67200}{54} x = 1244.4 x \text{ ml of H}_2 \text{ at N.T.P}$$

$$(ii) \quad 24 \text{ g of Mg} = 22400 \text{ ml of H}_2 \text{ at N.T.P} \quad (1 - x) \text{ g of Mg} = \frac{22400}{24} \times (1 - x) = 933.3 (1 - x) \text{ ml of H}_2$$

Hence total vol. of H<sub>2</sub> collected at N.T.P = 1244.4 x + 933.3 (1 - x) ml

But total vol. of H<sub>2</sub> as calculated above = 1104 ml

$$\therefore 1244.4 x + 933.3 (1 - x) = 1104 \text{ ml}$$

$$1244.4 x - 933.3 x = 1104 - 933.3$$

$$311.1 x = 170.7, \quad x = 0.5487$$

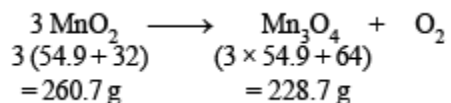
Hence 1 g of alloy contains Al = 0.5487 g

$$\therefore \text{Percentage of Al in alloy} = \frac{0.5487 \times 100}{1} = 54.87\%$$

$$\% \text{ of Mg in alloy} = 100 - 54.87 = 45.13\%$$

**Q. 3. Igniting MnO<sub>2</sub> converts it quantitatively to Mn<sub>3</sub>O<sub>4</sub>. A sample of pyrolusite is of the following composition : MnO<sub>2</sub> 80%, SiO<sub>2</sub> and other inert constituents 15%, rest being water. The sample is ignited in air to constant weight. What is the percentage of Mn in the ignited sample? (1978) [O = 16, Mn = 54.9]**

**Ans. Sol.**



Let the amount of pyrolusite ignited = 100.00 g

∴ Wt. of MnO<sub>2</sub> = 80 g (80% of 100 g = 80 g) Wt. of SiO<sub>2</sub> and other inert substances = 15 g  
Wt. of water = 100 – (80 + 15) = 5 g

According to equation, 260.7 g of MnO<sub>2</sub> gives = 228.7 g of Mn<sub>3</sub>O<sub>4</sub>

$$\therefore 80 \text{ g of MnO}_2 \text{ gives} = \frac{228.7}{260.7} \times 80 = 70.2 \text{ g of Mn}_3\text{O}_4$$

NOTE :

During ignition, H<sub>2</sub>O present in pyrolusite is removed while silica and other inert substances remain as such.

∴ Total wt. of the residue = 70.2 + 15 = 85.2 g

Calculation of % of Mn in ignited Mn<sub>3</sub>O<sub>4</sub>



$$3 \times 54.9 = 164.7 \text{ g} \qquad 3 \times 54.9 + 64 = 228.7 \text{ g}$$

Since, 228.7 g of Mn<sub>3</sub>O<sub>4</sub> contains 164.7 g of Mn

$$70.2 \text{ g of Mn}_3\text{O}_4 \text{ contains} = \frac{164.7}{228.7} \times 70.2 = 50.55 \text{ g of Mn}$$

Weight of residue = 85.2 g

Hence, percentage of Mn in the ignited sample

$$= \frac{50.55}{85.2} \times 100 = 59.33\%$$



**Q. 4.** 4.215 g of a metallic carbonate was heated in a hard glass tube and the CO<sub>2</sub> evolved was found to measure 1336 ml at 27°C and 700 mm pressure. What is the equivalent weight of the metal? (1979)

**Ans. Sol.** TIPS/Formulae : (i) Find the volume of CO<sub>2</sub> at NTP

(ii) Find molecular wt. of metal carbonate

(iii) Find the wt. of metal

(iv) Calculate equivalent weight of metal

Given P<sub>1</sub> = 700 mm, P<sub>2</sub> = 760 mm, V<sub>1</sub> = 1336 ml, V<sub>2</sub> = ?

T<sub>1</sub> = 300 K, T<sub>2</sub> = 273 K

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}, \text{ or } V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{700 \times 1336 \times 273}{760 \times 300}$$

$$= 1119.78 \text{ ml} = 1.12 \text{ L at NTP}$$

∴ 1.12 L of CO<sub>2</sub> is given by carbonate = 4.215 g

$$\text{Molecular weight of metal carbonate} = \frac{4.215}{1.12} \times 22.4$$

$$= 84.3$$

$$\text{Metal carbonate is } MCO_3 = M + 12 + 48 = M + 60$$

$$\text{Atomic weight of } M = 84.3 - 60 = 24.3$$

$$\text{Eq. wt. of metal} = \frac{1}{2} \times \text{M. wt.} = \frac{1}{2} \times 24.3 = 12.15$$

**Q. 5. (a)** 5.5 g of a mixture of FeSO<sub>4</sub> · 7H<sub>2</sub>O and Fe<sub>2</sub>(SO<sub>4</sub>)<sub>3</sub> · 9H<sub>2</sub>O requires 5.4 ml of 0.1 N KMnO<sub>4</sub> solution for complete oxidation. Calculate the number of gram mole of hydrated ferric sulphate in the mixture.

**(b)** The vapour density (hydrogen = 1) of a mixture consisting of NO<sub>2</sub> and N<sub>2</sub>O<sub>4</sub> is 38.3 at 26.7°C. Calculate the number of moles of NO<sub>2</sub> in 100 g of the mixture. (1979)

**Ans. Sol.** (a) Equivalents of KMnO<sub>4</sub> = Equivalents of FeSO<sub>4</sub> · 7H<sub>2</sub>O





5.4 ml 0.1 N  $\text{KMnO}_4 = \frac{5.4 \times 0.1}{1000} = 5.4 \times 10^{-4}$  equivalents  
 Amount of  $\text{FeSO}_4 = 5.4 \times 10^{-4} \times$   
 Mol wt. of

$\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$

$$= 5.4 \times 10^{-4} \times 278 = 0.150 \text{ g}$$

Total weight of mixture = 5.5 g

Amount of ferric sulphate =  $5.5 - 0.150 \text{ g} = 5.35 \text{ g}$

$$\text{Hence Moles of ferric sulphate} = \frac{\text{Mass}}{\text{M. wt.}} = \frac{5.35}{562}$$

$$= 9.5 \times 10^{-3} \text{ gram-mole}$$

(b) Using the relation, Mol. wt. =  $2 \times$  vapour density, we get Mol. wt. =  $2 \times 38.3 = 76.6$

$$\text{No. of moles} = \frac{\text{Mass}}{\text{Mol. wt.}} = \frac{100}{76.6} = 1.30 \quad \dots(\text{i})$$

Let weight of  $\text{NO}_2$  in mixture = x g

$$\text{Then weight of } \text{N}_2\text{O}_4 \text{ in mixture} = 100 - x \text{ No. of moles of } \text{NO}_2 = \frac{\text{Mass}}{\text{Mol. wt.}} = \frac{x}{46} \quad \dots(\text{ii})$$

$$\text{No. of moles of } \text{N}_2\text{O}_4 = \frac{\text{Mass}}{\text{Mol. wt.}} = \frac{100 - x}{92} \quad \dots(\text{iii})$$

- According to problem

$$1.30 = \frac{x}{46} + \frac{100 - x}{92}$$

On solving the equation we find,  $x = 20.1$

$$\therefore \text{weight of } \text{NO}_2 = 20.1 \text{ g Moles of } \text{NO}_2 = \frac{\text{Mass}}{\text{M. wt.}} = \frac{20.1}{46} = 0.437 \text{ moles.}$$



**Q. 6.** 5 ml of a gas containing only carbon and hydrogen were mixed with an excess of oxygen (30 ml) and the mixture exploded by means of an electric spark. After the explosion, the volume of the mixed gases remaining was 25 ml. On adding a concentrated solution of potassium hydroxide, the volume further diminished to 15 ml of the residual gas being pure oxygen. All volumes have been reduced to N.T.P.

Calculate the molecular formula of the hydrocarbon gas. (1979)

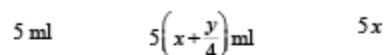
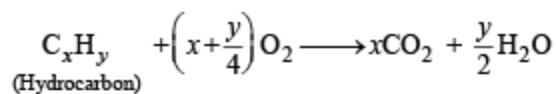
**Ans. Sol.** Volume of oxygen taken = 30 ml, Volume of unused oxygen = 15 ml

Volume of O<sub>2</sub> used = Volume of O<sub>2</sub> added – Volume of O<sub>2</sub> left

$$= 30 - 15 = 15 \text{ ml}$$

Volume of CO<sub>2</sub> produced = Volume of gaseous mixture after explosion – Volume of unused oxygen or Volume of CO<sub>2</sub> produced = 25 – 15 = 10 ml

Volume of hydrocarbon = 5 ml General equation for combustion of a hydrocarbon is as follows –



∴ Volume of CO<sub>2</sub> produced = 5x, Since Volume of CO<sub>2</sub> = 10 ml

∴ 5x = 10 ⇒ x = 2, Volume of O<sub>2</sub> used = 15 ml

$$\therefore 5\left(x + \frac{y}{4}\right) = 15 \Rightarrow x + \frac{y}{4} = 3$$

$$\Rightarrow 2 + \frac{y}{4} = 3 \quad (\because x = 2) \Rightarrow 8 + y = 12 \therefore y = 4$$

Hence Molecular formula of hydrocarbon is C<sub>2</sub>H<sub>4</sub>.



**Q. 7. In the analysis of 0.500 g sample of feldspar, a mixture of chlorides of sodium and potassium is obtained which weighs 0.1180g. Subsequent treatment of mixed chlorides with silver nitrate gives 0.2451g of silver chloride. What is the percentage of sodium oxide and potassium oxide in feldspar. (1979)**

**Ans. Sol.** TIPS/Formulae : (i) Equate given mass of AgCl against mass obtained from NaCl and KCl

(ii)  $2\text{NaCl} \equiv \text{Na}_2\text{O}$  &  $2\text{KCl} \equiv \text{K}_2\text{O}$  Let amount of NaCl in mixture = x gm

$\therefore$  amount of KCl in mixture =  $(0.118 - x)$  gm

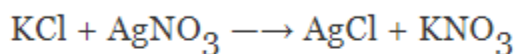


143.5 g

$\therefore$  58.5 g NaCl gives AgCl = 143.5g

$$\therefore x \text{ g NaCl gives AgCl} = \frac{143.5}{58.5} \times x \text{ g}$$

Again



74.5 g

143.5 g

$\therefore$  74.5 g KCl gives AgCl = 143.5g

$$\therefore (0.118 - x) \text{ g KCl gives AgCl} = \left( \frac{143.5}{74.5} \times 0.118 - x \right) \text{ g}$$

Total weight of AgCl = 0.2451g

$$\therefore \left( \frac{143.5}{58.5} \times x \right) + \left[ \frac{143.5}{74.5} \times (0.118 - x) \right] = 0.2451$$

$\therefore x = 0.0338\text{g}$

$\therefore$  Amount of NaCl in mixture = 0.0338g

$\therefore$  Amount of KCl in mixture =  $0.118 - 0.0338 = 0.0842\text{g}$



Since



$$2 \times 58.5 \quad 62$$

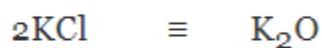
$$= 117.0$$

$\therefore$  117g NaCl is equivalent to = 62.0g  $\text{Na}_2\text{O}$

$$\therefore 0.0338\text{g NaCl is equivalent to} = \frac{62.0}{117} \times 0.0338 \text{ g Na}_2\text{O}$$

$$= 0.0179\text{g}$$

$$\% \text{ of Na}_2\text{O in 0.5g of feldspar} = \frac{0.0179}{0.500} \times 100 = 3.58\%$$



$$2 \times 74.5 = 149 \quad 94$$

$\therefore$  149g of KCl is equivalent to = 94g  $\text{K}_2\text{O}$

$$\therefore 0.0842\text{g of KCl is equivalent to} = \frac{94}{149} \times 0.0842 = 0.0531\text{g K}_2\text{O}$$

$$\therefore \% \text{ of K}_2\text{O in 0.5g of feldspar} = \frac{0.0531}{0.5} \times 100 = 10.62\%$$

$$\% \text{ of Na}_2\text{O in feldspar} = 3.58\%$$

$$\% \text{ of K}_2\text{O in feldspar} = 10.62\%$$

**Q. 8. A compound contains 28 percent of nitrogen and 72 percent of metal by weight. 3 atoms of metal combine with 2 atoms of N. Find the atomic weight of metal. (1980)**

**Ans. Sol.** According to problem, three atoms of M combine with 2 atoms of N

$\therefore$  Formula of compound is  $\text{M}_3\text{N}_2$  (Where M is the metal) Equivalent wt of N =  $\frac{14}{3}$  ( $\therefore$  valency of N in compound is 3)

$\therefore$  28 g N combines with = 72g metal

$$\therefore \frac{14}{3} \text{ N combines with} = \frac{72}{28} \times \frac{14}{3} = 12$$



∴ Eq. wt. of metal = 12

At wt of metal = Eq. wt × valency = 12 × 2 = 24 [Valency of metal = 2]

**Q. 9. (i) A sample of  $\text{MnSO}_4 \cdot 4\text{H}_2\text{O}$  is strongly heated in air. The residue is  $\text{Mn}_3\text{O}_4$ .**

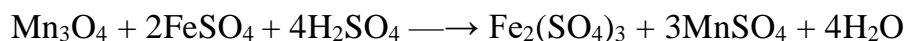
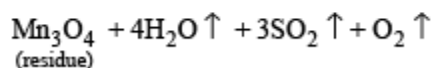
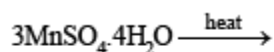
**(ii) The residue is dissolved in 100 ml of 0.1 N  $\text{FeSO}_4$  containing dilute  $\text{H}_2\text{SO}_4$ .**

**(iii) The solution reacts completely with 50 ml of  $\text{KMnO}_4$  solution.**

**(iv) 25 ml of the  $\text{KMnO}_4$  solution used in step (iii) requires 30 ml of 0.1 N  $\text{FeSO}_4$  solution for complete reaction.**

**Find the amount of  $\text{MnSO}_4 \cdot 4\text{H}_2\text{O}$  present in the sample. (1980)**

**Ans. Sol.** Following reactions take place



Milliequivalents of  $\text{FeSO}_4$  in 30 ml of 0.1N  $\text{FeSO}_4 = 30 \times 0.1 = 3$  m. eq.

According to problem step (iv) 25 ml of  $\text{KMnO}_4$  reacts with = 3 m eq of  $\text{FeSO}_4$

Thus in step (iii) of the problem, 50 ml of  $\text{KMnO}_4$  reacts with =  $\frac{3}{25} \times$  m.eq. of  $\text{FeSO}_4$

= 6 meq of  $\text{FeSO}_4$

Milli eq. of 100 ml of 0.1N  $\text{FeSO}_4 = 100 \times 0.1 = 10$  m eq.

$\text{FeSO}_4$  which reacted with  $\text{Mn}_3\text{O}_4 = (10 - 6) = 4$  m eq.

Milli eq of  $\text{FeSO}_4 =$  Milli eq. of  $\text{Mn}_3\text{O}_4$  (∵ Milli eq of oxidising agent and reducing agent are equal)

∴  $\text{Mn}_3\text{O}_4 \equiv 3\text{MnSO}_4 \cdot 4\text{H}_2\text{O}$



∴ 1 Meq of  $\text{Mn}_3\text{O}_4 = 3 \text{ Meq of MnSO}_4 \cdot 4\text{H}_2\text{O}$

∴ 4 Meq of  $\text{Mn}_3\text{O}_4 = 12 \text{ Meq of MnSO}_4 \cdot 4\text{H}_2\text{O}$

$$\text{Eq. wt of MnSO}_4 \cdot 4\text{H}_2\text{O} = \frac{\text{Molwt.}}{2} = \frac{223}{2} = 111.5$$

Wt of  $\text{MnSO}_4 \cdot 4\text{H}_2\text{O}$  in sample =  $12 \times 111.5 = 1338 \text{ mg} = 1.338\text{g}$ .

**Q. 10. (a) One litre of a sample of hard water contains 1 mg of  $\text{CaCl}_2$  and 1 mg of  $\text{MgCl}_2$ . Find the total hardness in terms of parts of  $\text{CaCO}_3$  per 100 parts of water by weight.**

**(b) A sample of hard water contains 20 mg of  $\text{C}^{a++}$  ions per litre. How many milli-equivalent of  $\text{Na}_2\text{CO}_3$  would be required to soften 1 litre of the sample?**

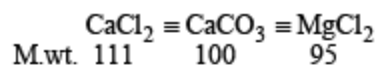
**(c) 1 gm of Mg is burnt in a closed vessel which contains 0.5 gm of  $\text{O}_2$ .**

**(i) Which reactant is left in excess?**

**(ii) Find the weight of the excess reactants?**

**(iii) How many milliliters of 0.5 N  $\text{H}_2\text{SO}_4$  will dissolve the residue in the vessel. (1980)**

**Ans. Sol. (a)**



From this it is evident, that 111 mg  $\text{CaCl}_2$  will give  $\text{CaCO}_3 = 100\text{mg}$

$$\therefore 1 \text{ mg CaCl}_2 \text{ will give CaCO}_3 = \frac{100}{111} \text{ mg} = 0.90 \text{ mg}$$

95 mg  $\text{MgCl}_2$  gives  $\text{CaCO}_3 = 100 \text{ mg}$

$$\therefore 1 \text{ mg MgCl}_2 \text{ gives CaCO}_3 = \frac{100}{95} \text{ mg} = 1.05 \text{ mg}$$

∴ Total  $\text{CaCO}_3$  formed by 1 mg  $\text{CaCl}_2$  and 1 mg  $\text{MgCl}_2 = 0.90 + 1.05 = 1.95 \text{ mg}$

∴ Amount of  $\text{CaCO}_3$  present per litre of water = 1.95mg

∴ wt of 1 ml of water =  $1\text{g} = 10^3 \text{ mg}$



∴ wt of 1000 ml of water =  $10^3 \times 10^3 = 10^6$ mg

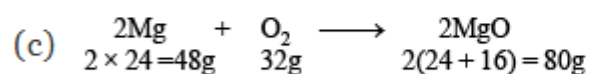
∴ Total hardness of water in terms of parts of  $\text{CaCO}_3$  per  $10^6$  parts of water by weight = 1.95 parts.

$$(b) \text{ Eq wt of } \text{Ca}^{++} = \frac{\text{Mol.wt}}{\text{Charge}} = \frac{40}{2} = 20$$



1 milliequivalent of  $\text{Ca}^{2+} = 20$  mg

1 milliequivalent of  $\text{Na}_2\text{CO}_3$  is required to soften 1 litre of hard water.



∴ 32g of  $\text{O}_2$  reacts with = 48g Mg

∴ 0.5g of  $\text{O}_2$  reacts with =  $\frac{48}{32} \times 0.5 = 0.75$ g

Weight of unreacted Mg =  $1.00 - 0.75 = 0.25$ g

Thus Mg is left in excess.

Weight of MgO formed =  $\frac{80}{48} \times 0.75 = 1.25$ g



(40g) According to reaction

∴ 40g MgO is dissolved it gives 1000 ml of 1 N.  $\text{H}_2\text{SO}_4$

∴ 40 g MgO is dissolved it gives 2000 ml 0.5 N  $\text{H}_2\text{SO}_4$

∴ 1.25 MgO is dissolved it gives

$$= \frac{2000 \times 1.25}{40} \text{ml of } 0.5 \text{ N } \text{H}_2\text{SO}_4$$

= 62.5ml of 0.5N  $\text{H}_2\text{SO}_4$



**Q. 11.** A hydrocarbon contains 10.5g of carbon per gram of hydrogen. 1 litre of the vapour of the hydrocarbon at 127°C and 1 atmosphere pressure weighs 2.8g. Find the molecular formula. (1980)

**Ans. Sol.**

Given  $P = 1 \text{ atm}$ ,  $V = 1\text{L}$ ,  $T = 127^\circ\text{C} = 127 + 273 = 400 \text{ K}$

$PV = nRT$  (Ideal gas equation)

$$\text{or } n = \frac{PV}{RT} = \frac{1 \times 1}{0.082 \times 400} = 0.0304$$

$$\text{Mol. wt} = \frac{\text{Mass}}{\text{Moles}} = \frac{2.8}{0.0304} = 92.10$$

Element	wt. of element	Relative no. of atoms	Ratio of atoms	Wholen. of atoms
C	10.5	$\frac{10.5}{12} = 0.875$	$\frac{0.875}{0.875} = 1$	$1 \times 7 = 7$
H	1.0	$\frac{1.0}{1} = 1$	$\frac{1}{0.875} = 1.14$	$1.14 \times 7 = 8$

$\therefore$  Empirical formula =  $\text{C}_7\text{H}_8$

Empirical formula, wt =  $12 \times 7 + 1 \times 8 = 92$

$$n = \frac{\text{Molecular wt}}{\text{Empirical formula wt}} = \frac{92.10}{92} = 1$$

Molecular formula =  $n \times$  empirical formula

$$= 1(\text{C}_7\text{H}_8) = \text{C}_7\text{H}_8$$

**Q. 12.** Find (1980) (i) The total number of neutrons and (ii) The total mass of neutron in 7 mg of  $^{14}\text{C}$ . (Assume that mass of neutron = mass of hydrogen atom)

**Ans. Sol.**

(i) No. of C atoms in 14g of  $^{14}\text{C} = 6.02 \times 10^{23}$

$\therefore$  No. of C atom in 7 mg (7/1000g) of  $^{14}\text{C}$





$$= \frac{6.02 \times 10^{23} \times 7}{14 \times 1000} = 3.01 \times 10^{20}$$

No. of neutrons in 1 carbon atom = 7

$$\therefore \text{Total no. of neutrons in 7 mg of } ^{14}\text{C} = 3.01 \times 10^{20} \times 7 \\ = 21.07 \times 10^{20}$$

$$\text{Wt of 1 neutron} = \text{wt of 1 hydrogen atom} = \frac{1}{6.02 \times 10^{23}} \text{ g}$$

$\therefore$  Wt of  $3.01 \times 10^{20} \times 7$  neutrons

$$= \frac{3.0 \times 10^{20} \times 7}{6.02 \times 10^{23}} = 3.5 \times 10^{-3} \text{ g}$$

**Q. 13. A mixture contains NaCl and unknown chloride MCl.**

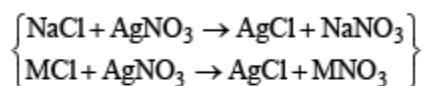
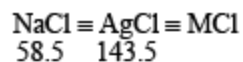
**(i) 1 g of this is dissolved in water. Excess of acidified  $\text{AgNO}_3$  solution is added to it. 2.567 g of white ppt. is formed.**

**(ii) 1 g of original mixture is heated to  $300^\circ\text{C}$ . Some vapours come out which are absorbed in acidified  $\text{AgNO}_3$  solution, 1.341 g of white precipitate was obtained. Find the molecular weight of unknown chloride. (1980)**

**Ans. Sol.** Weight of AgCl formed = 2.567 g

Amount of AgCl formed due to MCl = 1.341 g ( $\because$  NaCl does not decompose on heating to  $300^\circ\text{C}$ )

$\therefore$  Weight of AgCl formed due to NaCl =  $2.567 - 1.341 = 1.226\text{g}$



$\therefore$  143.5g of AgCl is obtained from NaCl = 58.5g

$\therefore$  1.226 g of AgCl is obtained from NaCl

$$= \frac{58.5}{143.5} \times 1.226 = 0.4997 \text{ g}$$



∴ Wt of MCl in 1 g of mixture = 1.000 - 0.4997 = 0.5003g

∴ 1.341 g of AgCl is obtained from MCl = 0.5003g

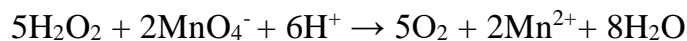
∴ 143.5g of AgCl is obtained from MCl

$$= \frac{0.5003}{1.341} \times 143.5 = 53.53 \text{ g}$$

∴ Molecular weight of MCl = 53.53

**Q. 14. A 1.00 gm sample of H<sub>2</sub>O<sub>2</sub> solution containing X per cent H<sub>2</sub>O<sub>2</sub> by weight requires X ml of a KMnO<sub>4</sub> solution for complete oxidation under acidic conditions. Calculate the normality of the KMnO<sub>4</sub> solution. (1981 - 3 Marks)**

**Ans. Sol.** The complete oxidation under acidic conditions can be represented as follows:



Since 34 g of H<sub>2</sub>O<sub>2</sub> = 2000 ml of 1N . H<sub>2</sub>O<sub>2</sub>

$$\left( \because \text{Eq. wt of H}_2\text{O}_2 = \frac{34}{2} \right)$$

∴ 34 g of H<sub>2</sub>O<sub>2</sub> = 2000 ml of 1N KMnO<sub>4</sub> [∵ N<sub>1</sub>V<sub>1</sub> = N<sub>2</sub>V<sub>2</sub>]

$$\text{or } \frac{X}{100} \text{ g H}_2\text{O}_2 = \frac{2000 \times X}{100 \times 34} \text{ ml of 1N KMnO}_4$$

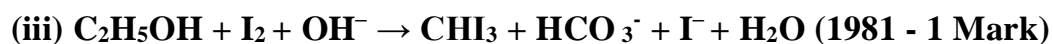
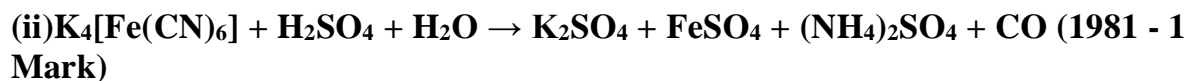
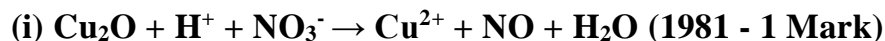
$$\text{Therefore the unknown normality} = \frac{2000 \times X}{34 \times 100 \times X}$$

$$= \frac{10}{17} \text{ or } 0.588 \text{ N}$$

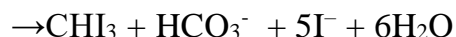
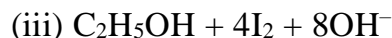
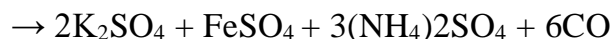
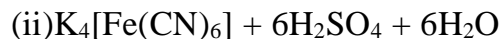
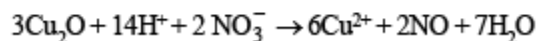
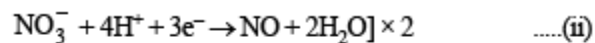
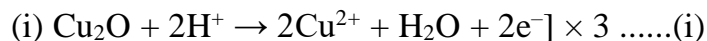


## Subjective questions of Some Basic Concepts of Chemistry (Part -2)

Q. 15. Balance the following equations.



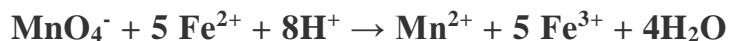
Ans. Sol. Balance the reactions by ion electron method.



Q. 16. Hydroxylamine reduces iron (III) according to the equation:



Iron (II) thus produced is estimated by titration with a standard permanganate solution. The reaction is :



A 10 ml. sample of hydroxylamine solution was diluted to 1 litre. 50 ml. of this diluted solution was boiled with an excess of iron (III) solution. The resulting solution required 12 ml. of 0.02 M  $\text{KMnO}_4$  solution for complete oxidation of iron (II). Calculate the weight of hydroxylamine in one litre of the original solution. (H = 1, N = 14, O = 16, K = 39, Mn = 55, Fe = 56) (1982 - 4 Marks)



**Ans. Sol.** Given  $2\text{NH}_2\text{OH} + 4\text{Fe}^{3+} \rightarrow \text{N}_2\text{O} + \text{H}_2\text{O} + 4\text{Fe}^{2+} + 4\text{H}^+ \dots$ (i) and  $\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O} \dots$ (ii)

$\therefore 10\text{NH}_2\text{OH} + 4\text{MnO}_4^- + 12\text{H}^+ \rightarrow 5\text{N}_2\text{O} + 21\text{H}_2\text{O} + 4\text{Mn}^{2+}$  [On multiplying (i) by 5 and (ii) by 4 and then adding the resulting equations]

Molecular weight of  $\text{NH}_2\text{OH} = 33$

Thus 4000 ml of 1M  $\text{MnO}_4^-$  would react with  $\text{NH}_2\text{OH} = 330\text{g}$

$\therefore 12\text{ ml}$  of 0.02 M  $\text{KMnO}_4$  would react with  $\text{NH}_2\text{OH}$

$$= \frac{330 \times 12 \times 0.02}{400} \text{g}$$

$\therefore$  Amount of  $\text{NH}_2\text{OH}$  present in 1000 ml of diluted solution

$$= \frac{330 \times 12 \times 0.02 \times 1000}{4000 \times 50} \text{g}$$

Since 10 ml of sample of hydroxylamine is diluted to one litre

$\therefore$  Amount of hydroxyl amine in one litre of original solution

$$= \frac{330 \times 0.02 \times 12 \times 1000}{4000 \times 50} \times \frac{1000}{10} \text{g} = 39.6 \text{ g}$$

**Q. 17.** The density of a 3 M sodium thiosulphate solution ( $\text{Na}_2\text{S}_2\text{O}_3$ ) is 1.25 g per ml. Calculate (i) the percentage by weight of sodium thiosulphate, (ii) the mole fraction of sodium thiosulphate and (iii) the molalities of  $\text{N}^{\text{a}+}$  and  $\text{S}_2\text{O}_3^{2-}$  ions. (1983 - 5 Marks)

**Ans. Sol.** TIPS/Formulae :

(i) Mole fraction =  $\frac{\text{Moles of substance}}{\text{Total moles}}$

(ii) 1 mole of  $\text{Na}_2\text{S}_2\text{O}_3$  gives 2 moles of  $\text{Na}^+$  and 1 mole of  $\text{S}_2\text{O}_3^{2-}$  Molecular wt. of sodium thiosulphate solution ( $\text{Na}_2\text{S}_2\text{O}_3$ )

$$= 23 \times 2 + 32 \times 2 + 16 \times 3 = 158$$

(i) The percentage by weight of  $\text{Na}_2\text{S}_2\text{O}_3 = \frac{\text{wt of Na}_2\text{S}_2\text{O}_3}{\text{wt of solution}} \times 100 = \frac{3 \times 158 \times 100}{1000 \times 1.25} = 37.92$



[Wt. of  $\text{Na}_2\text{S}_2\text{O}_3 = \text{Molarity} \times \text{Mol wt}$ ] (ii) Mass of 1 litre solution =  $1.25 \times 1000 \text{ g} = 1250 \text{ g}$

[ $\therefore$  density = 1.25g/l]

Mole fraction of  $\text{Na}_2\text{S}_2\text{O}_3$

$$= \frac{\text{Number of moles of } \text{Na}_2\text{S}_2\text{O}_3}{\text{Total number of moles}}$$

$$\text{Moles of water} = \frac{1250 - 158 \times 3}{18} = 43.1$$

$$\text{Mole fraction of } \text{Na}_2\text{S}_2\text{O}_3 = \frac{3}{3 + 43.1} = 0.065$$

(iii) 1 mole of sodium thiosulphate ( $\text{Na}_2\text{S}_2\text{O}_3$ ) yields 2 moles of  $\text{Na}^+$  and 1 mole of  $\text{S}_2\text{O}_3^{2-}$ .

$$\text{Molality of } \text{Na}_2\text{S}_2\text{O}_3 = \frac{3 \times 1000}{776} = 3.87$$

$$\text{Molality of } \text{Na}^+ = 3.87 \times 2 = 7.74\text{m}$$

$$\text{Molality of } \text{S}_2\text{O}_3^{2-} = 3.87\text{m}$$

**Q. 18.** 4.08 g of a mixture of BaO and an unknown carbonate  $\text{MCO}_3$  was heated strongly. The residue weighed 3.64 g. This was dissolved in 100 ml of 1 N HCl. The excess acid required 16 ml of 2.5 N NaOH solution for complete neutralization.

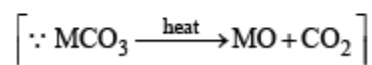
Identify the metal M. (1983 - 4 Marks) (At. wt. H = 1, C = 12, O = 16, Cl = 35.5, Ba = 138)

**Ans. Sol.** Weight of  $\text{MCO}_3$  and BaO = 4.08 g (given)

Weight of residue = 3.64 g (given)

$\therefore$  Weight of  $\text{CO}_2$  evolved on heating =  $(4.08 - 3.64) \text{ g} = 0.44 \text{ g}$

$$= \frac{0.44}{44} = 0.01 \text{ mole}$$



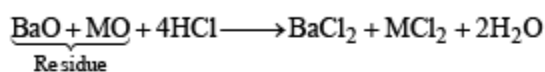
Volume of 1N HCl in which residue is dissolved = 100 ml



Volume of 1N HCl used for dissolution =  $(100 - 2.5 \times 16)$  ml = 60 ml

$$= \frac{60}{1000} = 0.06 \text{ equivalents}$$

The chemical equation for dissolution can be written as



[Number of moles of BaO and MO =  $1 + 1 = 2$ ]

$$\text{Number of moles of BaO} + \text{Number of moles of MO} = \frac{0.06}{2}$$

$$= 0.03$$

Number of moles of BaO =  $(0.03 - 0.01) = 0.02$  moles

Molecular weight of BaO =  $138 + 16 = 154$

$$\therefore \text{Weight of BaO} = (0.02 \times 154) \text{ g} = 3.08 \text{ g}$$

Weight of  $\text{MCO}_3 = (4.08 - 3.08) = 1.0$  g

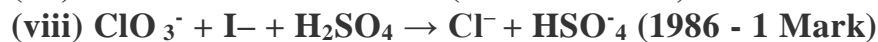
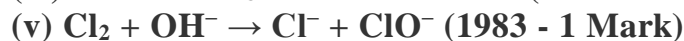
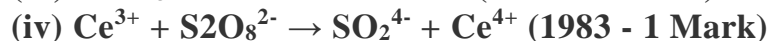
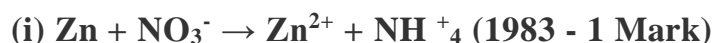
Since weight of 0.01 mole of  $\text{MCO}_3 = 1.0$  g

$$\therefore \text{Mol. wt. of } \text{MCO}_3 = \frac{1}{0.01} = 100$$

Hence atomic weight of unknown M =  $(100 - 60) = 40$

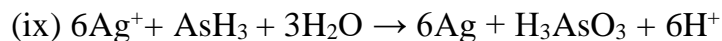
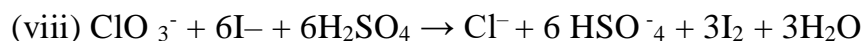
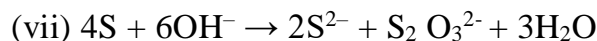
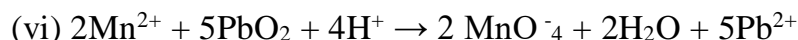
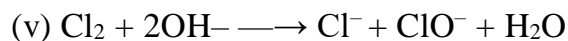
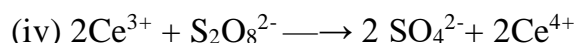
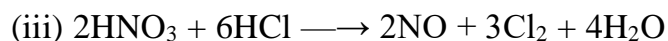
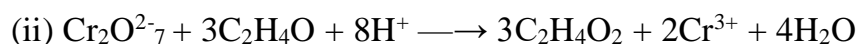
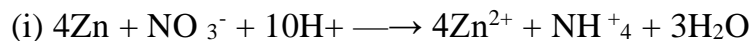
The atomic weight of metal is 40 so the metal M is Ca.

**Q. 19. Complete and balance the following reactions :**



**Ans. Sol. TIPS/Formulae :** Balance the atoms as well as charges by ion electron/oxidation number method.

While balancing the equations, both the charges and atoms must balance.



**Q. 20.**  $2.68 \times 10^{-3}$  moles of a solution containing an ion  $\text{A}^{n+}$  require  $1.61 \times 10^{-3}$  moles of  $\text{MnO}_4^-$  for the oxidation of  $\text{A}^{n+}$  to  $\text{AO}_3^-$  in acid medium. What is the value of  $n$ ? (1984 - 2 Marks)

**Ans. Sol. TIPS/Formulae :** Equivalents of A oxidised = Equivalents of A reduced.  
Since in acidic medium,  $\text{A}^{n+}$  is oxidised to  $\text{AO}_3^-$ , the change in oxidation state from

$$(+5) \text{ to } (+n) = 5 - n \quad [\because \text{O.S. of A in } \text{AO}_3^- = +5]$$

$\therefore$  Total number of electrons that have been given out during oxidation of  $2.68 \times 10^{-3}$  moles of  $\text{A}^{n+}$

$$= 2.68 \times 10^{-3} \times (5 - n)$$

Thus the number of electrons added to reduce  $1.61 \times 10^{-3}$  moles of  $\text{MnO}_4^-$  to  $\text{Mn}^{2+}$ ,  
i.e. (+7) to (+2) =  $1.61 \times 10^{-3} \times 5$

$$[\text{Number of electrons involved} = +7 - (+2) = 5]$$



$$\therefore 1.61 \times 10^{-3} \times 5 = 2.68 \times 10^{-3} \times (5 - n)$$

$$5 - n = \frac{1.61 \times 5}{2.68} \text{ or } n = 5 - \frac{8.05}{2.68} \approx 2$$

**Q. 21.** Five ml of 8N nitric acid, 4.8 ml of 5N hydrochloric acid and a certain volume of 17M sulphuric acid are mixed together and made upto 2litre. Thirty ml. of this acid mixture exactly neutralise 42.9 ml of sodium carbonate solution containing one gram of  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$  in 100 ml. of water. Calculate the amount in gram of the sulphate ions in solution. (1985 - 4 Marks)

**Ans. Sol.** TIPS/Formulae : (i) Find normality of acid mixture and  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ . Equate them to find volume of  $\text{H}_2\text{SO}_4$ .

$$(ii) \text{ Meq. of } \text{H}_2\text{SO}_4 = V \times N = \frac{V \times N}{1000} \text{ eq.}$$

(iii) Equivalent of  $\text{SO}_4^{2-}$  = equivalents of  $\text{H}_2\text{SO}_4 \times \text{Eq. wt. of } \text{SO}_4^{2-}$   
 $N \times V \text{ (ml.)} = \text{meq.}$

Acid mixture contains 5 ml of 8N,  $\text{HNO}_3$ , 4.8 ml of 5N,  $\text{HCl}$  and say, 'V' ml of 17 M  $\equiv 34 \text{ N}$ ,  $\text{H}_2\text{SO}_4$ . [ $1\text{MH}_2\text{SO}_4 = 2\text{N} \cdot \text{H}_2\text{SO}_4$ ]

$$N \text{ of the acid mixture} = \frac{\text{meq. (total) of acid}}{\text{ml of solution}}$$

$$= \frac{5 \times 8 + 4.8 \times 5 + V \times 34}{2000} \text{ [Total volume} = 2 \text{ L} = 2000 \text{ ml]}$$

$$\text{or, } N_{\text{mixture}} = \frac{64 + 34V}{2000}$$

$$\therefore \text{ Eq. of wt. of } \text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} = \frac{\text{Mol. wt.}}{2}$$

$$= \frac{106 + 180}{2} = 143$$

$$N \text{ of } \text{Na}_2\text{CO}_3 = \frac{\text{Meq. of } \text{Na}_2\text{CO}_3}{\text{Volume of solution (ml)}}$$

$$= \frac{\frac{1}{100}}{\frac{100}{1000}} = \frac{1}{143} \times \frac{1000}{100} = 0.069\text{N}$$

$$N_1V_1 = N_2V_2$$





$$\text{or } 30 \times N_{\text{mixture}} = 42.9 \times 0.069$$

(acid)                      (sod. carbonate)

$$\therefore N_{\text{mixture}} = \frac{42.9 \times 0.069}{30}$$

$$\text{Hence } \frac{64 + 34V}{2000} = 0.0986$$

$$64 + 34V = 0.0986 \times 2000, \quad 64 + 34V = 197.2$$

$$34V = 197.2 - 64.0 = 133.2 \quad \therefore \text{or } V = \frac{133.2}{34} = 3.9 \text{ ml.}$$

$$\text{Hence meq. of } \text{H}_2\text{SO}_4 = V \times N \text{ of } \text{H}_2\text{SO}_4$$

$$= 3.9 \times 34 = 132.6 \text{ meq.}$$

$$= 0.1326 \text{ eq. of } \text{H}_2\text{SO}_4$$

$$= 0.1326 \text{ eq. of } \text{SO}_4^{2-}$$

$$= 0.1326 \times 48 \text{ g of } \text{SO}_4^{2-}$$

$$\left( \because \text{Eq. wt. of } \text{SO}_4^{2-} = \frac{32 + 64}{2} = 48 \right)$$

$$= 6.3648 \text{ g of } \text{SO}_4^{2-} \text{ are in } 3.9 \text{ ml of } 17\text{M } \text{H}_2\text{SO}_4$$

**Q. 22. Arrange the following in increasing oxidation number of iodine. (1986 - 1 Mark)**

**I<sub>2</sub>, HI, HIO<sub>4</sub>, ICl**

**Ans. Sol.** HI < I<sub>2</sub> < ICl < HIO<sub>4</sub>; O.N. of I in I<sub>2</sub> = 0, HI = -1, ICl = +1, HIO<sub>4</sub> = +7.

**Q. 23. (i) What is the weight of sodium bromate and molarity of solution necessary to prepare 85.5 ml of 0.672 N solution when the half-cell reaction is -**



**(ii) What would be the weight as well as molarity if the half-cell reaction is : -**



**Ans. Sol.** (i) From the given half-cell reaction, Here Eq. wt. of NaBrO<sub>3</sub>



$$= \frac{\text{Mol. wt.}}{6} = \frac{151}{6} = 25.17$$

[∵ number of electron involved = 6]

Now we know that Meq. = Normality × Vol. in ml. =  $85.5 \times 0.672 = 57.456$

$$\text{Also Meq.} = \frac{W_{\text{NaBrO}_3}}{\text{Eq. wt.}_{\text{NaBrO}_3}} \times 1000$$

$$= \frac{W_{\text{NaBrO}_3}}{25.17} \times 1000$$

$$\frac{W_{\text{NaBrO}_3}}{25.17} \times 1000 = 57.456 \text{ g}$$

$$\therefore W_{\text{NaBrO}_3} = 1.446 \text{ g}$$

$$\text{Molarity of NaBrO}_3 = \frac{\text{Normality}}{\text{Valence factor}}$$

$$= \frac{0.672}{6} = 0.112 \text{ M}$$

(ii) From the given half-cell reaction, Eq. wt. of  $\text{NaBrO}_3 = \frac{\text{Mol. wt.}}{5} = \frac{151}{5} = 30.2$

[Number of electron involved per  $\text{BrO}_3^- = \frac{10}{2} = 5$ ]

Thus, the amount of  $\text{NaBrO}_3$  required for preparing 1000 ml. of 1 N  $\text{NaBrO}_3 = 30.2 \text{ g}$

∴ The amount of  $\text{NaBrO}_3$  required for preparing 85.5 ml of 0.672 N  $\text{NaBrO}_3$ .

$$= \frac{30.2 \times 0.672 \times 85.5}{1000} = 1.7532 \text{ g}$$

Hence, Molarity =  $\frac{0.672}{5} = 0.1344 \text{ M}$



**Q. 24.** A sugar syrup of weight 214.2 g contains 34.2 g of sugar ( $C_{12}H_{22}O_{11}$ ). Calculate : (i) molal concentration and (ii) mole fraction of sugar in the syrup. (1988 - 2 Marks)

**Ans. Sol.** (i) Weight of sugar syrup = 214.2 g Weight of sugar in the syrup = 34.2 g

$\therefore$  Weight of water in the syrup = 214.2 – 34.2 = 180.0 g Mol. wt. of sugar,  $C_{12}H_{22}O_{11}$  = 342

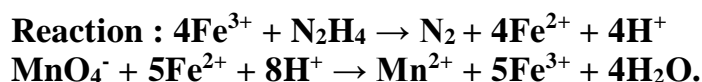
$$\therefore \text{Molal concentration} = \frac{34.2 \times 1000}{342 \times 180} = 0.56$$

(ii) Mol. wt. of water,  $H_2O$  = 18

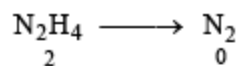
$$\therefore \text{Mole fraction of sugar} = \frac{34.2/342}{180/18 + 34.2/342}$$

$$= \frac{0.1}{10 + 0.1} = \frac{0.1}{10.1} = 0.0099$$

**Q. 25.** A sample of hydrazine sulphate ( $N_2H_6SO_4$ ) was dissolved in 100 ml. of water, 10 ml of this solution was reacted with excess of ferric chloride solution and warmed to complete the reaction. Ferrous ion formed was estimated and it required 20 ml. of M/50 potassium permanganate solution. Estimate the amount of hydrazine sulphate in one litre of the solution. (1988 - 3 Marks)



**Ans. Sol.** TIPS/Formulae : No. of equivalents of  $KMnO_4$  = No. of equivalents of hydrazine sulphate.



Change in oxidation state for each  $N_2H_4 = 2 \times 2 = 4$

$$\text{Equivalent weight of } N_2H_6SO_4 = \frac{130}{4} = 32.5$$

Normality of  $KMnO_4 = 5 \times 450$  ( $\because$  valence factor = 5)



Number of equivalents of  $\text{KMnO}_4 = 20 \times \frac{5}{50 \times 1000} = \frac{1}{500}$  and if weight of hydrazine sulphate be x gm then

$$\text{equivalents of hydrazine sulphate} = \frac{x}{32.5}$$

$$\therefore \frac{1}{500} = \frac{x}{32.5} \text{ or } x = \frac{32.5}{500} = 0.065 \text{ g}$$

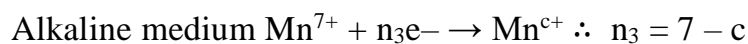
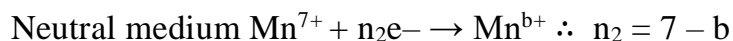
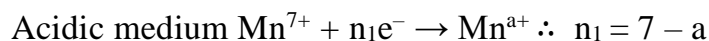
Hence wt. of  $\text{N}_2\text{H}_6\text{SO}_4$  in 10 ml solution = 0.065 g

$\therefore$  Wt. of  $\text{N}_2\text{H}_6\text{SO}_4$  in 1000 ml solution = 6.5 g

**Q. 26. An equal volume of a reducing agent is titrated separately with 1M  $\text{KMnO}_4$  in acid neutral and alkaline media. The volumes of  $\text{KMnO}_4$  required are 20 ml. in acid, 33.4 ml. neutral and 100 ml. in alkaline media. Find out the oxidation state of manganese in each reduction product. Give the balanced equations for all the three half reactions. Find out the volume of 1M  $\text{K}_2\text{Cr}_2\text{O}_7$  consumed; if the same volume of the reducing agent is titrated in acid medium.(1989 - 5 Marks)**

**Ans. Sol.** TIPS/Formulae : No. of equivalents of  $\text{KMnO}_4$  in neutral medium = No. of equivalents of reducing agent.

Assuming that  $\text{KMnO}_4$  shows the following changes during its oxidising nature.



Let V ml. of reducing agent be used for  $\text{KMnO}_4$  in different medium.

$\therefore$  Meq. of reducing agent

= Meq. of  $\text{KMnO}_4$  in acid medium Meq. of  $\text{KMnO}_4$  in neutral medium

$$= \text{Meq. of } \text{KMnO}_4 \text{ in alkaline medium} = 1 \times n_1 \times 20 = 1 \times n_2 \times 33.4 = 1 \times n_3 \times 100 = n_1 = 1.667 n_2 = 5 n_3$$



Since  $n_1$ ,  $n_2$  and  $n_3$  are integers and  $n_1$  is not greater than 7

$\therefore n_3 = 1$  Hence  $n_1 = 5$  and  $n_2 = 3$

$\therefore$  Different oxidation states of Mn in Acidic medium  $\text{Mn}^{7+} + 5e^- \rightarrow \text{Mn}^{a+}$  or  $a = + 2$

Neutral medium  $\text{Mn}^{7+} + 3e^- \rightarrow \text{Mn}^{b+}$  or  $b = + 4$

Alkaline medium  $\text{Mn}^{7+} + 1e^- \rightarrow \text{Mn}^{c+}$  or  $c = + 6$

Further, same volume of reducing agent is treated with  $\text{K}_2\text{Cr}_2\text{O}_7$ , and therefore Meq. of reducing agent = Meq. of  $\text{K}_2\text{Cr}_2\text{O}_7$

$$1 \times 5 \times 20 = 1 \times 6 \times V [\text{Cr}^{+6} + 6e^- \rightarrow 2\text{Cr}^{+3}]$$

$$V = 16.66 \text{ mL} \therefore 1\text{M} = 6 \times 1\text{N}$$

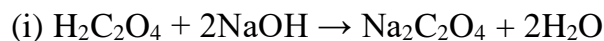
**Q. 27. A mixture of  $\text{H}_2\text{C}_2\text{O}_4$  (oxalic acid) and  $\text{NaHC}_2\text{O}_4$  weighing 2.02 g was dissolved in water and solution made upto one litre. Ten millilitres of the solution required 3.0 ml. of 0.1 N sodium hydroxide solution for complete neutralization. In another experiment, 10.0 ml. of the same solution, in hot dilute sulphuric acid medium, require 4.0 ml. of 0.1 N potassium permanganate solution for complete reaction.**

**Calculate the amount of  $\text{H}_2\text{C}_2\text{O}_4$  and  $\text{NaHC}_2\text{O}_4$  in the mixture. (1990 - 5 Marks)**

**Ans. Sol.** TIPS/Formulae : No. of equivalents of  $\text{KMnO}_4$

= No. of equivalents of reducing agents.

Case I. Reaction of  $\text{NaOH}$  with  $\text{H}_2\text{C}_2\text{O}_4$  and  $\text{NaHC}_2\text{O}_4$ .



Number of milliequivalents of  $\text{NaOH} = N \times V = 3.0 \times 0.1 = 0.3$

$\therefore$  Combined normality of the mixture titrated with  $\text{NaOH}$

$$= \frac{0.3}{10} = 0.03$$



Case II. Reaction of  $C_2O_4^-$  ion and  $KMnO_4$

(iii)  $5C_2O_4^- + MnO_4^- + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$   $KMnO_4$  will react in same manner with both  $NaHC_2O_4$  and  $H_2C_2O_4$  as it can be seen from the above reaction.

Number of milliequivalents of  $KMnO_4 = 4.0 \times 0.1 = 0.4$

$\therefore$  Combined normality of the mixture titrated with  $KMnO_4$

$$= \frac{0.4}{10} = 0.04$$

The difference ( $0.04\text{ N} - 0.03\text{ N} = 0.01\text{ N}$ ) is due to  $NaHC_2O_4$  The total normality of  $NaHC_2O_4$  will be  $= 0.01 + 0.01 = 0.02\text{ N}$  From equation (ii) in case I.

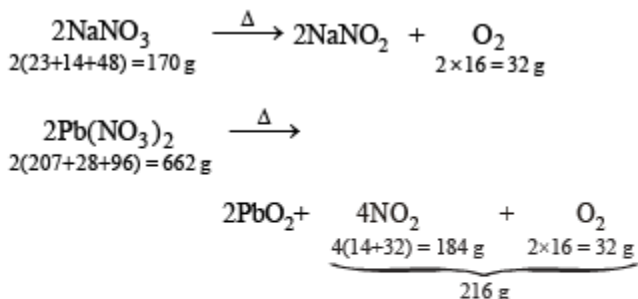
Eq. wt. of  $NaHC_2O_4 = 112$  Amount of  $NaHC_2O_4$  in one litre of solution formed  $= 0.01 \times 112 = 1.12\text{ g}$  and amount of  $H_2C_2O_4 = 2.02 - \text{Wt. of } NaHC_2O_4 = 2.02 - 1.12 = 0.90\text{ g}$

**Q. 28. A solid mixture (5.0 g) consisting of lead nitrate and sodium nitrate was heated below  $600^\circ\text{C}$  until the weight of the residue was constant. If the loss in weight is 28.0 per cent, find the amount of lead nitrate and sodium nitrate in the mixture. (1990 - 4 Marks)**

**Ans. Sol. TIPS/Formulae :** Let the amount of  $NaNO_3$  in the mixture  $= x\text{ g}$

$\therefore$  The amount of  $Pb(NO_3)_2$  in the mixture  $= (5 - x)\text{ g}$

Heating effect of sodium nitrate and lead nitrate



Now since, 170 g of  $NaNO_3$  gives = 32 g of  $O_2$

$$\therefore x\text{ g of } NaNO_3 \text{ gives} = \frac{32}{170} \times x\text{ g of } O_2$$

## Subjective questions of Some Basic Concepts of Chemistry (Part -3)

**Q. 29.** Calculate the molality of 1 litre solution of 93%  $\text{H}_2\text{SO}_4$  (weight/volume). The density of the solution is 1.84 g/ml. (1990 - 1 Marks)

**Ans. Sol.** TIPS/Formulae :

$$\text{Molality} = \frac{\text{Mass of solute/M. wt. of solute}}{\text{Mass of solvent in kg}}$$

Mass of  $\text{H}_2\text{SO}_4$  in 100 ml of 93%  $\text{H}_2\text{SO}_4$  solution = 93 g

$\therefore$  Mass of  $\text{H}_2\text{SO}_4$  in 1000 ml of the  $\text{H}_2\text{SO}_4$  solution = 930 g

Mass of 1000 ml  $\text{H}_2\text{SO}_4$  solution =  $1000 \times 1.84 = 1840$  g

Mass of water in 1000 ml of solution =  $1840 - 930$  g = 910 g = 0.910 kg

$$\text{Moles of } \text{H}_2\text{SO}_4 = \frac{\text{Wt. of } \text{H}_2\text{SO}_4}{\text{Mol. wt. of } \text{H}_2\text{SO}_4} = \frac{930}{98}$$

$\therefore$  Moles of  $\text{H}_2\text{SO}_4$  in 1 kg of water

$$= \frac{930}{98} \times \frac{1}{0.910} = 10.43 \text{ mol}$$

$\therefore$  Molality of solution = 10.43m

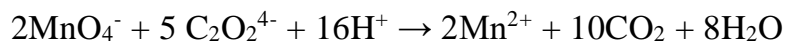
**Q. 30.** A solution of 0.2 g of a compound containing  $\text{Cu}^{2+}$  and  $\text{C}_2^{2-} \text{O}_4$  ions on titration with 0.02 M  $\text{KMnO}_4$  in presence of  $\text{H}_2\text{SO}_4$  consumes 22.6 ml. of the oxidant. The resultant solution is neutralized with  $\text{Na}_2\text{CO}_3$ , acidified with dil. acetic acid and treated with excess KI. The liberated iodine requires 11.3 ml of 0.05 M  $\text{Na}_2\text{S}_2\text{O}_3$  solution for complete reduction.

Find out the molar ratio of  $\text{Cu}^{2+}$  to  $\text{C}_2\text{O}_4^{2-}$  in the compound.

Write down the balanced redox reactions involved in the above titrations. (1991 - 5 Marks)



**Ans. Sol.** In the given problem, a solution containing  $\text{Cu}^{2+}$  and  $\text{C}_2\text{O}_4^{2-}$  is titrated first with  $\text{KMnO}_4$  and then with  $\text{Na}_2\text{S}_2\text{O}_3$  in presence of  $\text{KI}$ . In titration with  $\text{KMnO}_4$ , it is the  $\text{C}_2\text{O}_4^{2-}$  ions that react with the  $\text{MnO}_4^-$  ions. The concerned balanced equation may be written as given below.



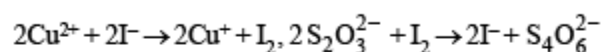
Thus according to the above reaction  $2 \text{ mmol MnO}_4^- \equiv 5 \text{ mmol C}_2\text{O}_4^{2-}$

However, No. of mmol of  $\text{MnO}_4^-$  used in titration = Vol. in ml  $\times$  M =  $22.6 \times 0.02 = 0.452 \text{ mmol MnO}_4^-$

Since  $2 \text{ mmol MnO}_4^- \equiv 5 \text{ mmol C}_2\text{O}_4^{2-}$   $0.452 \text{ mmol MnO}_4^- \equiv \frac{5}{2} \times 0.452 = 1.130 \text{ mmol C}_2\text{O}_4^{2-}$

**Titration with  $\text{Na}_2\text{S}_2\text{O}_3$  in the presence of  $\text{KI}$ .**

Here  $\text{Cu}^{2+}$  react and the reactions involved during titration are



Thus  $2\text{Cu}^{2+} \equiv \text{I}_2 \equiv 2\text{S}_2\text{O}_3^{2-}$

No. of mmol of  $\text{S}_2\text{O}_3^{2-}$  used in titration

$$= 0.05 \times 11.3 = 0.565 \text{ mmol S}_2\text{O}_3^{2-}$$

Now since  $2 \text{ mmol S}_2\text{O}_3^{2-} \equiv 2 \text{ mmol Cu}^{2+}$  [From above equation]

$$0.565 \text{ mmol S}_2\text{O}_3^{2-} = \frac{2}{2} \times 0.565 \text{ mmol Cu}^{2+}$$

$$= 0.565 \text{ mmol Cu}^{2+}$$

$$\therefore \text{Molar ratio of Cu}^{2+} \text{ to } \text{C}_2\text{O}_4^{2-} = \frac{0.565 \text{ mmol}}{1.130 \text{ mmol}} = 1 : 2$$

Balanced equations in two cases **Case I.**  $\text{Mn}^{7+} + 5\text{e}^- \rightarrow \text{Mn}^{2+}$



**Case II.**  $2\text{Cu}^{+2} + 2\text{e}^- \rightarrow \text{Cu}^{2+}$   $2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$  and  $\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$   $2\text{S}_2\text{O}_3^{2-} \rightarrow \text{S}_4^{+3/2} + 2\text{e}^-$





**Q. 31.** A 1.0 g sample of  $\text{Fe}_2\text{O}_3$  solid of 55.2% purity is dissolved in acid and reduced by heating the solution with zinc dust. The resultant solution is cooled and made upto 100.0 ml. An aliquot of 25.0 ml of this solution requires 17.0 ml of 0.0167 M solution of an oxidant for titration. Calculate the number of electrons taken up by the oxidant in the reaction of the above titration. (1991 - 4 Marks)

**Ans. Sol.** Mass of  $\text{Fe}_2\text{O}_3$  in the sample =  $\frac{55.2}{100} \times 1 = 0.552 \text{ g}$

Number of moles of  $\text{Fe}_2\text{O}_3 = \frac{0.552}{159.8} = 3.454 \times 10^{-3}$

Number of moles of  $\text{Fe}^{3+}$  ions =  $2 \times 3.454 \times 10^{-3}$

=  $6.9 \times 10^{-3} \text{ mol} = 6.90 \text{ mmol}$  Since its only 1 electron is exchanged in the conversion of  $\text{Fe}^{3+}$  to  $\text{Fe}^{2+}$ , the molecular mass is the same as equivalent mass.

$\therefore$  Amount of  $\text{Fe}^{2+}$  ion in 100 ml. of sol. = 6.90 meq Volume of oxidant used for 100 ml of  $\text{Fe}^{2+}$  sol.

=  $17 \times 4 = 68 \text{ ml.}$

Amount of oxidant used =  $68 \times 0.0167 \text{ mmol}$

= 1.1356 mmol Let the number of electrons taken by the oxidant = n

$\therefore$  No. of meq. of oxidant used =  $1.1356 \times n$

Thus  $1.1356 \times n = 6.90 \Rightarrow n = \frac{6.90}{1.1356} = 6$

**Q. 32.** A 2.0 g sample of a mixture containing sodium carbonate, sodium bicarbonate and sodium sulphate is gently heated till the evolution of  $\text{CO}_2$  ceases. The volume of  $\text{CO}_2$  at 750 mm Hg pressure and at 298 K is measured to be 123.9 ml. A 1.5g of the same sample requires 150 ml. of (M/10) HCl for complete neutralisation. Calculate the % composition of the components of the mixture. (1992 - 5 Marks)

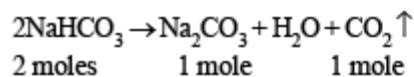
**Ans. Sol.** 1.5 g of sample require = 150 ml. of  $\frac{\text{M}}{10}$  HCl

$\therefore$  2 g of sample require =  $\frac{150 \times 2}{1.5} \text{ ml. of } \frac{\text{M}}{10} \text{ HCl}$

= 200 ml. of  $\frac{\text{M}}{10}$  HCl

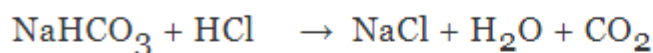


On heating, the sample, only  $\text{NaHCO}_3$  undergoes decomposition as given below.

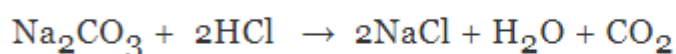


2 equ.

Neutralisation of the sample with HCl takes place as given below.



1 eq.            1 eq.



1 mole            2 mole

2 eq.            2 eq.

Hence, 2 g sample  $\equiv$  200 ml. of M/10 HCl

= 200 ml. of N/10 HCl = 20 meq = 0.020 eq

Number of moles of  $\text{CO}_2$  formed, i.e.

$$n = \frac{PV}{RT} = \frac{750}{760} \times \frac{123.9}{1000} \times \frac{1}{0.082 \times 298} = 0.005$$

Moles of  $\text{NaHCO}_3$  in the sample (2 g) =  $2 \times 0.005 = 0.01$

Equivalent of  $\text{NaHCO}_3 = 0.01$

Wt. of  $\text{NaHCO}_3 = 0.01 \times 84 = 0.84 \text{ g}$

$$\% \text{ of } \text{NaHCO}_3 = \frac{0.84 \times 100}{2} = 42\%$$

Equivalent of  $\text{Na}_2\text{CO}_3 = 0.02 - 0.01 = 0.01$

Wt. of  $\text{Na}_2\text{CO}_3 = 0.01 \times 53 = 0.53 \text{ g}$

$$\therefore \% \text{ of } \text{Na}_2\text{CO}_3 = \frac{0.53 \times 100}{2} = 26.5\%$$

$\therefore$  % of  $\text{Na}_2\text{SO}_4$  in the mixture =  $100 - (42 + 26.5) = 31.5\%$

**Q. 33. One gram of commercial  $\text{AgNO}_3$  is dissolved in 50 ml. of water. It is treated with 50 ml. of a KI solution. The silver iodide thus precipitated is filtered off.**

**Excess of KI in the filtrate is titrated with (M/10)  $\text{KIO}_3$  solution in presence of 6M**

HCl till all  $I^-$  ions are converted into ICl. It requires 50 ml. of (M/10)  $KIO_3$  solution. 20 ml. of the same stock solution of KI requires 30 ml. of (M/10)  $KIO_3$  under similar conditions.

Calculate the percentage of  $AgNO_3$  in the sample.

(Reaction :  $KIO_3 + 2KI + 6HCl \rightarrow 3ICl + 3KCl + 3H_2O$ ) (1992 - 4 Marks)

**Ans. Sol.** Reaction involved titration is 1 mole 2 moles  $KIO_3 + 2KI + 6HCl \rightarrow 3ICl + 3KCl + 3H_2O$

20 ml. of stock KI solution  $\equiv$  30 ml. of  $\frac{M}{10}$   $KIO_3$  solution

$$\text{Molarity of KI solution} = \frac{30 \times 1 \times 2}{20 \times 10} = \frac{3}{10}$$

$$\text{Millimoles in 50 ml. of KI solution} = 50 \times \frac{3}{10} = 15$$

Millimoles of KI left unreacted with  $AgNO_3$  solution

$$= 2 \times 50 \times \frac{1}{10} = 10$$

$$\therefore \text{Millimoles of KI reacted with } AgNO_3 = 15 - 10 = 5$$

Millimoles of  $AgNO_3$  present in  $AgNO_3$  solution = 5

Molecular weight of  $AgNO_3 = 170$

$$\therefore \text{Wt. of } AgNO_3 \text{ in the solution} = 5 \times 10^{-3} \times 170 = 0.850 \text{ g}$$

$$\% AgNO_3 \text{ in the sample} = \frac{0.850}{1} \times 100 = 85\%$$

**Q. 34.** Upon mixing 45.0 ml. of 0.25 M lead nitrate solution with 25.0 ml of 0.10 M chromic sulphate solution, precipitation of lead sulphate takes place. How many moles of lead sulphate are formed? Also, calculate the molar concentrations of the species left behind in the final solution. Assume that lead sulphate is completely insoluble. (1993 - 3 Marks)

**Ans. Sol.** Calculation of number of moles in 45 ml. of 0.025 M  $Pb(NO_3)_2$



$$\text{Moles of Pb(NO}_3)_2 = 0.25 \times \frac{45}{1000} = 0.01125$$

∴ Initial moles of  $\text{Pb}^{2+} = 0.01125$

Moles of  $\text{NO}_3^- = 0.01125 \times 2 = 0.02250$  [1 mole  $\text{Pb(NO}_3)_2 \equiv 2$  moles of  $\text{NO}_3^-$ ]

Calculation of number of moles in 25 ml. of 0.1 M chromic sulphate

Moles of chromic sulphate ( $\text{Cr}_2(\text{SO}_4)_3$ )

$$= 0.1 \times \frac{25}{1000} = 0.0025 \text{ moles}$$

Moles of  $\text{SO}_4^{2-} = 0.0025 \times 3 = 0.0075$  [1 Mole of chromic sulphate  $\equiv 3$  moles of  $\text{SO}_4^{2-}$ ]

Moles of  $\text{PbSO}_4$  formed = 0.0075 [ $\text{SO}_4^{2-}$  is totally consumed]

Moles of  $\text{Pb}^{2+}$  left =  $0.01125 - 0.0075 = 0.00375$

Moles of  $\text{NO}_3^-$  left = 0.02250 [ $\text{NO}_3^-$  remain unreacted]

Moles of chromium ions =  $0.0025 \times 2 = 0.005$

Total volume of the solution =  $45 + 25 = 70$  ml.

∴ Molar concentration of the species left

$$\text{(i) } \text{Pb}^{2+} = \frac{0.00375}{70} \times 1000 = 0.05357 \text{ M}$$

$$\text{(ii) } \text{NO}_3^- = \frac{0.0225}{70} \times 1000 = 0.3214 \text{ M}$$

$$\text{(iii) } \text{Cr}^{3+} = \frac{0.005}{70} \times 1000 = 0.0714 \text{ M}$$

**Q. 35. The composition of a sample of Wustite is  $\text{Fe}_{0.93}\text{O}_{1.00}$ .**

**What percentage of the iron is present in the form of Fe (III)? (1994 - 2 Marks)**

**Ans. Sol.** In pure iron oxide ( $\text{FeO}$ ), iron and oxygen are present in the ratio 1 : 1. However, here number of  $\text{Fe}^{2+}$  present = 0.93 or No. of  $\text{Fe}^{2+}$  ions missing = 0.07 Since each  $\text{Fe}^{2+}$  ion has 2 positive charge, the total number of charge due to missing (0.07)  $\text{Fe}^{2+}$  ions =  $0.07 \times 2 = 0.14$  To maintain electrical neutrality, 0.14 positive charge is compensated by the presence of  $\text{Fe}^{3+}$  ions. Now since, replacement of one  $\text{Fe}^{2+}$  ion by

one  $\text{Fe}^{3+}$  ion increases one positive charge, 0.14 positive charge must be compensated by the presence of 0.14  $\text{Fe}^{3+}$  ions.

In short, 0.93  $\text{Fe}^{2+}$  ions have 0.14  $\text{Fe}^{3+}$  ions 100  $\text{Fe}^{2+}$  ions have =  $\frac{0.14}{0.93} \times 100 = 15.05\%$

**Q. 36.**  $8.0575 \times 10^{-2}$  kg of Glauber's salt is dissolved in water to obtain  $1 \text{ dm}^3$  of a solution of density  $1077.2 \text{ kg m}^{-3}$ . Calculate the molarity, molality and mole fraction of  $\text{Na}_2\text{SO}_4$  in the solution. (1994 - 3 Marks)

**Ans. Sol.** The formula of Glauber's salt is  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$  Molecular mass of  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$

$$= [2 \times 23 + 32.1 + 4 \times 16] + 10 (1.01 \times 2 + 16) = 322.3 \text{ g mol}^{-1}$$

Weight of the Glauber's salt taken = 80.575 gm Out of 80.575 g of salt, weight of anhydrous  $\text{Na}_2\text{SO}_4$

$$= \frac{142.1}{322.3} \times 80.575 = 35.525 \text{ g}$$

$$\text{Number of moles of } \text{Na}_2\text{SO}_4 \text{ per } \text{dm}^3 \text{ of the solution} = \frac{35.525}{142.1} = 0.25$$

$$\text{Molarity of the solution} = 0.25 \text{ M}$$

$$\text{Density of solution} = 1077.2 \text{ kgm}^{-3}$$

$$= \frac{1077.2 \times 10^3}{10^6} \text{ gm cm}^{-3} = 1.0772 \text{ g cm}^{-3}$$

$$\text{Total weight of sol} = V \times d = 1 \text{ dm}^3 \times d$$

$$= 1000 \text{ cm}^3 \times 1.0772 \text{ gcm}^{-3} = 1077.2 \text{ g}$$

$$\text{Weight of water} = 1077.2 - 35.525 = 1041.67 \text{ g}$$

$$\text{Molality of sol.} = \frac{0.25}{1041.67 \text{ g}} \times 1000 \text{ g} = 0.2399 = 0.24 \text{ m}$$

$$\text{Number of moles of water in the solution} = \frac{1041.67}{18} = 57.87$$

Mole fraction of  $\text{Na}_2\text{SO}_4$



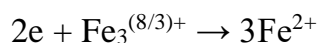
$$= \frac{\text{No. of moles of Na}_2\text{SO}_4}{\text{Total number of moles}} = \frac{0.25}{0.25 + 57.87}$$

$$= 0.0043 = 4.3 \times 10^{-3}$$

**Q. 37.** A 3.00 g sample containing  $\text{Fe}_3\text{O}_4$ ,  $\text{Fe}_2\text{O}_3$  and an inert impure substance, is treated with excess of KI solution in presence of dilute  $\text{H}_2\text{SO}_4$ . The entire iron is converted into  $\text{Fe}^{2+}$  along with the liberation of iodine. The resulting solution is diluted to 100 ml. A 20 ml of the diluted solution requires 11.0 ml of 0.5 M  $\text{Na}_2\text{S}_2\text{O}_3$  solution to reduce the iodine present. A 50 ml of the diluted solution, after complete extraction of the iodine requires 12.80 ml of 0.25 M  $\text{KMnO}_4$  solution in dilute  $\text{H}_2\text{SO}_4$  medium for the oxidation of  $\text{Fe}^{2+}$ . Calculate the percentages of  $\text{Fe}_2\text{O}_3$  and  $\text{Fe}_3\text{O}_4$  in the original sample. (1996 - 5 Marks)

**Ans. Sol.** TIPS/Formulae : Find the milliequivalents and equate them as per data given in question.

For  $\text{Fe}_3\text{O}_4 \rightarrow 3\text{FeO}$



Thus, valence factor for  $\text{Fe}_3\text{O}_4$  is 2 and for  $\text{FeO}$  is 2/3.

For,  $\text{Fe}_2\text{O}_3 \rightarrow 2\text{FeO}$ ;  $2e + \text{Fe}_2^{3+} \rightarrow 2\text{Fe}^{2+}$  ... (1)

Thus valence factor for  $\text{Fe}_2\text{O}_3$  is 2 and for  $\text{FeO}$  is 1.

Let Meq. of  $\text{Fe}_3\text{O}_4$  and  $\text{Fe}_2\text{O}_3$  be a and b respectively.

$\therefore$  Meq. of  $\text{Fe}_3\text{O}_4$  + Meq.  $\text{Fe}_2\text{O}_3$  = Meq. of  $\text{I}_2$  liberated

= Meq. of hypo used

$$a + b = \frac{11 \times 0.5 \times 100}{20} = 27.5$$

Now, the  $\text{Fe}^{2+}$  ions are again oxidised to  $\text{Fe}^{3+}$  by  $\text{KMnO}_4$ .

Note that in the change  $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + e^-$ ; valence factor of  $\text{Fe}^{2+}$  is 1.

Thus, Meq. of  $\text{Fe}^{2+}$  (from  $\text{Fe}_3\text{O}_4$ ) + Meq. of  $\text{Fe}^{2+}$  (from  $\text{Fe}_2\text{O}_3$ ) = Meq. of  $\text{KMnO}_4$  used .... (2)

If valence factor for  $\text{Fe}^{2+}$  is  $2/3$  from Eq. (1), then Meq. of  $\text{Fe}^{2+}$  (from  $\text{Fe}_3\text{O}_4$ ) = a

If valence factor for  $\text{Fe}^{2+}$  is 1 then Meq. of  $\text{Fe}^{2+}$  (from  $\text{Fe}_3\text{O}_4$ ) =  $3a/2$  ... (3)

Similarly, from Eq. (2), Meq. of  $\text{Fe}^{2+}$  from ( $\text{Fe}_2\text{O}_3$ ) = b.

$$\therefore 3a/2 + b = 0.25 \times 5 \times 12.8 \times 100/50 = 32 \text{ or } 3a + 2b = 64 \dots(4)$$

From Eqs. (3) and (4)

Meq. of  $\text{Fe}_3\text{O}_4$  = a = 9 & Meq. of  $\text{Fe}_2\text{O}_3$  = b = 18.5

$$\therefore W_{\text{Fe}_3\text{O}_4} = \frac{9 \times 232}{2 \times 1000} = 1.044\text{g}$$

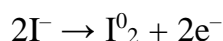
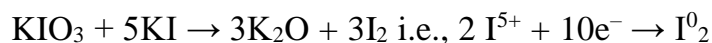
$$\text{and } W_{\text{Fe}_2\text{O}_3} = \frac{18.5 \times 160}{2 \times 1000} = 1.48\text{g}$$

$$\therefore \% \text{ of } \text{Fe}_3\text{O}_4 = \frac{1.044 \times 100}{3} = 34.8$$

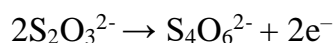
$$\text{and } \% \text{ of } \text{Fe}_2\text{O}_3 = \frac{1.48 \times 100}{3} = 49.33$$

**Q. 38. An aqueous solution containing 0.10 g  $\text{KIO}_3$  (formula weight = 214.0) was treated with an excess of KI solution. The solution was acidified with HCl. The liberated  $\text{I}_2$  consumed 45.0 mL of thiosulphate solution to decolourise the blue starch-iodine complex. Calculate the molarity of the sodium thiosulphate solution. (1998 - 5 Marks)**

**Ans. Sol.** TIPS/Formulae : Write the reactions taking place, balance them and equate moles of  $\text{I}_2$  and  $\text{Na}_2\text{S}_2\text{O}_3$ .



Now liberated  $\text{I}_2$  reacts with  $\text{Na}_2\text{S}_2\text{O}_3$   $\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$



$\therefore$  millimole ratio of  $\text{I}_2 : \text{S}_2\text{O}_3 = 1 : 2$



Thus, m mole of  $I_2$  liberated = m mole of  $Na_2S_4O_6$  used  $\times \frac{1}{2} = 45 \times M \times \frac{1}{2}$

[M is molarity of thiosulphate]

Also m mole of  $KIO_3 = \frac{0.1}{214} \times 1000$

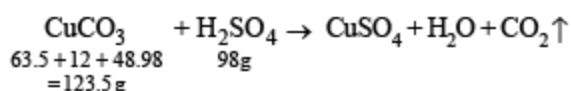
Now m mole ratio of  $KIO_3 : I_2 = 1 : 3$

Thus,  $\frac{(0.1/214) \times 1000}{(45M/2)} = \frac{1}{3}$

$\therefore M = \frac{0.1 \times 1000 \times 3 \times 2}{214 \times 45} = 0.062$

**Q. 39. How many millilitres of 0.5 M  $H_2SO_4$  are needed to dissolve 0.5 g of copper(II) carbonate? (1999 - 3 Marks)**

**Ans. Sol. TIPS/Formulae :** Use molarity equation to find volume of  $H_2SO_4$  solution.



$\therefore$  For 123.5 gms of Cu(II) carbonate 98 g of  $H_2SO_4$  are required.

For 0.5 gms of Cu(II) carbonate weight of  $H_2SO_4$  reqd.

$$= \frac{98 \times 0.5}{123.5} \text{ g} = 0.39676 \text{ g } H_2SO_4$$

Weight of required  $H_2SO_4 = 0.39676 \text{ g}$

Weight of solute in grams

$$= \frac{\text{Mol.wt.} \times \text{Molarity} \times \text{Volume in mL}}{1000}$$

$$0.39676 = \frac{98 \times 0.5 \times V}{1000} \text{ or } V = \frac{0.39676 \times 1000}{98 \times 0.5} \text{ ml}$$

Volume of  $H_2SO_4$  solution = 8.097 ml

**Q. 40. A plant virus is found to consist of uniform cylindrical particles of 150 Å in diameter and 5000 Å long. The specific volume of the virus is 0.75 cm<sup>3</sup>/g. If**



the virus is considered to be a single particle, find its molar mass.(1999 - 3 Marks)

**Ans. Sol.** TIPS/Formulae : (i) Volume of virus =  $\pi r^2 \ell$  (Volume of cylinder)

$$(ii) \text{ Mass of single virus} = \frac{\text{Volume}}{\text{Sp. volume}}$$

$$(iii) \text{ Molecular mass of virus} = \text{Mass of single virus} \times 6.02 \times 10^{23}$$

$$\text{Volume of virus} = \pi r^2 \ell$$

$$= \frac{22}{7} \times \frac{150}{2} \times \frac{150}{2} \times 10^{-16} \times 5000 \times 10^{-8}$$

$$= 0.884 \times 10^{-16} \text{ cm}^3$$

$$\text{Weight of one virus} = \frac{0.884 \times 10^{-16}}{0.75} \text{ g}$$

$$= 1.178 \times 10^{-16} \text{ g}$$

$$\therefore \text{ Mol. wt. of virus} = 1.178 \times 10^{-16} \times 6.02 \times 10^{23} = 7.09 \times 10^7$$

**Q. 41.** Hydrogen peroxide solution (20 ml) reacts quantitatively with a solution of  $\text{KMnO}_4$  (20 ml) acidified with dilute  $\text{H}_2\text{SO}_4$ .

The same volume of the  $\text{KMnO}_4$  solution is just decolourised by 10 ml of  $\text{MnSO}_4$  in neutral medium simultaneously forming a dark brown precipitate of hydrated  $\text{MnO}_2$ . The brown precipitate is dissolved in 10 ml of 0.2 M sodium oxalate under boiling condition in the presence of dilute  $\text{H}_2\text{SO}_4$ . Write the balanced equations involved in the reactions and calculate the molarity of  $\text{H}_2\text{O}_2$ . (2001 - 5 Marks)

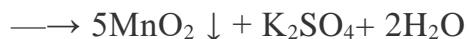
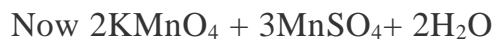
**Ans. Sol.** TIPS/Formulae : Write the balanced chemical reaction for change and apply mole concept.

The given reactions are  $\text{MnO}_2 \downarrow + \text{Na}_2\text{C}_2\text{O}_4 + 2\text{H}_2\text{SO}_4$

$\longrightarrow \text{MnSO}_4 + \text{CO}_2 + \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$

$$\therefore \text{ Meq. of } \text{MnO}_2 \equiv \text{ Meq of } \text{Na}_2\text{C}_2\text{O}_4 = 10 \times 0.2 \times 2 = 4$$

$$\therefore \text{ mM of } \text{MnO}_2 = \frac{4}{2} = 2 \quad \left[ \begin{array}{l} \text{Mn}^{4+} + 2\text{e} \rightarrow \text{Mn}^{2+} \\ \therefore \text{Valance factor of } \text{MnO}_2 = 2 \end{array} \right]$$



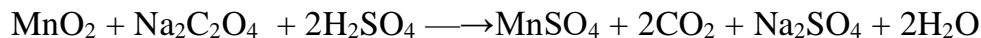
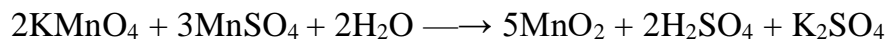
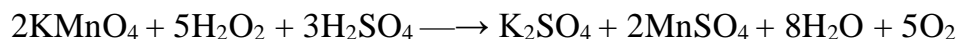
Since eq. wt. of  $\text{MnO}_2$  is derived from  $\text{KMnO}_4$  and  $\text{MnSO}_4$  both, thus it is better to proceed by mole concept

$$\text{mM of } \text{KMnO}_4 \equiv \text{mM of } \text{MnO}_2 \times (2/5) = 4/5$$



$$\therefore \text{mM of } \text{H}_2\text{O}_2 = \text{mM of } \text{KMnO}_4 \times \frac{5}{2} = \frac{4}{5} \times \frac{5}{2} = 2$$

$$\therefore \text{M}_{\text{H}_2\text{O}_2} \times 20 = 2 \text{ or } \text{M}_{\text{H}_2\text{O}_2} = 0.1$$



**Q. 42. Calculate the molarity of water if its density is  $1000 \text{ kg/m}^3$ . (2003 - 2 Marks)**

**Ans. Sol.** 1 litre water = 1 kg i.e. 1000 g water ( $\because d = 1000 \text{ kg/m}^3$ )

$$\equiv \frac{1000}{18} = 55.55 \text{ moles of water}$$

So, molarity of water = 55.55M

