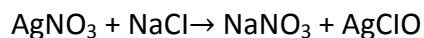


Some Basic Concepts of Chemistry

Sol 1.

Write the balance chemical equation and use mole concept for limiting reagent.



170 g 58.5 g 143.5 g

From the given data, we find AgNO_3 is limiting reagent as NaCl is in excess.

\therefore 170.0 g of AgNO_3 precipitates $\text{AgCl} = 143.5$ g

\therefore 5.77 g of AgNO_3 precipitates AgCl

$$= 143.5/170 * 5.77 = 4.87 \text{ g}$$

Sol 2.

(i) Find volume of H_2 at N.T.P.

(ii) Total amount of H_2 liberated = H_2 liberated by Mg & HCl + H_2 liberated by Al & HCl .

Conversion of volume of H_2 to N.T.P

Given condition N.T.P condition

$$P_1 = 0.92 \text{ atm.} \quad P_2 = 1 \text{ atm.}$$

$$V_1 = 1.20 \text{ litres} \quad V_2 = ?$$

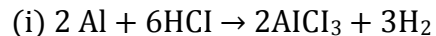
$$T_1 = 0 + 273 = 273 \text{ K} \quad T_2 = 273 \text{ K}$$

Applying ideal gas equation, $P_1V_1/T_1 = P_2V_2/T_2$

$$0.92 * 1.20 / 273 = 1 * V_2 / 273, V_2 = 0.92 * 1.20 * 273 / 273 * 1 \text{ litres} = 1.104$$

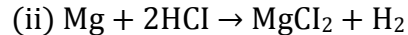
Litres = 1104 ml

The relevant chemical equation are



$$2 * 27 \quad 3 * 22400$$

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



$$24 \text{ g} \quad 22400 \text{ ml at NTP}$$

$$\text{Wt. of alloy} = 1 \text{ g}$$

$$\text{Let the wt. of aluminium in alloy} = x \text{ g}$$

$$\therefore \text{Wt. of magnesium in alloy} = (1 - x) \text{ g}$$

According to equation (i)

$$54 \text{ g Al} = 67200 \text{ ml of H}_2 \text{ at N.T.P}$$

$$\therefore x \text{ g of Al} = 67200/54 * x = 1244.4 x \text{ ml of H}_2 \text{ at N.T.P}$$

Similarly, from equation (ii)

$$24 \text{ g of Mg} = 222400 \text{ ml of H}_2 \text{ at N.T.P}$$

$$(1 - x) \text{ g of Mg} = 222400/24 * (1 - x) = 933.3 (1 - x) \text{ ml of H}_2$$

Hence total vol. of H₂ collected at N.T.P

$$= 12244.4 x + 933.3 (1 - x) \text{ ml}$$

But total vol. of H₂ as calculated above = 1104 ml

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

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

$$311.1 x = 170.7, x = 0.5487$$

Hence 1 g of alloy contains Al = 0.5487 g

$$\therefore \text{percentage of Al in alloy} = 0.5487 * 100/1 = 54.87\%$$

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

Sol 3.

$$3 (54.9 + 32) \quad (3 * 54.9 + 64)$$

$$= 260.7 \text{ g} \quad = 228.7 \text{ g}$$

Let the amount of pyrolusite ignited = 100.00 g

$$\therefore \text{Wt. of MnO}_2 = 80 \text{ g (80\% of 100 g = 80 g)}$$

Wt. of SiO₂ and other inert substances = 15 g

$$\text{Wt. of water } 100 - (80 + 15) = 5 \text{ g}$$

According to equation,

$$260.7 \text{ g of MnO}_2 \text{ gives } = 228.7 \text{ g of Mn}_3\text{O}_4$$

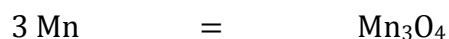
$$\therefore 80 \text{ g of MnO}_2 \text{ gives } = 228.7/260.7 * 80 = 70.2 \text{ g of Mn}_3\text{O}_4$$

NOTE :

During ignition, H₂O present in pyrolusite is removed while silica and other inert substances remain as such.

$$\therefore \text{Total wt. of the residue} = 70.2 + 15 = 85.2 \text{ g}$$

Calculation of % of Mn in ignited Mn₃O₄



$$3 * 54.9 = 164.7 \text{ g} \quad 3 * 54.9 + 64 = 228.7 \text{ g}$$

Since, 228.7 g of Mn₃O₄ contains 164.7 g of Mn

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

Weight of residue = 85.2 g

Hence, percentage of Mn in the ignited sample

$$= 50.55/85.2 * 100 = 59.33\%$$

Sol 4.

TIPS/Formulae:

- (i) Find the volume of CO₂ at NTP
- (ii) Find molecular wt. of metal carbonate
- (iii) Find the wt. of metal
- (iv) Calculate equivalent weight of metal

Given $P_1 = 700$ mm, $P_2 = 760$ mm, $V_1 = 1336$ ml, $V_2 = ?$ $T_1 = 300$ K, $T_2 = 273$ K

$$P_1V_1/T_1 = P_2V_2/T_2, \text{ or } V_2 = P_1V_1T_2/P_2T_1 = 700 * 1336 * 273/760 * 300$$

$$= 1119.378 \text{ ml} = 1.12 \text{ L at NTP}$$

$$\therefore 1.12 \text{ L of CO}_2 \text{ is given by carbonate} = 4.215 \text{ g}$$

$$\text{Molecular weight of metal carbonate} = 4.215/1.12 * 22.4 = 84.3$$

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

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

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

Sol 5.

(a) Equivalents of KMnO₄ = Equivalents of FeSO₄ · 7H₂O

$$5.4 \text{ ml } 0.1 \text{ N KMnO}_4 = 5.4 * 0.1/1000 = 5.4 * 10^{-4} \text{ equivalents}$$

$$\text{Amount of FeSO}_4 = 5.4 * 10^{-4} * \text{Mol. wt. of FeSO}_4 \cdot 7\text{H}_2\text{O} = 5.4 * 10^{-4} * 278 = 0.150 \text{ g}$$

$$\text{Total weight of mixture} = 5.5 \text{ g}$$

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

$$\text{Hence Moles of ferric sulphate} = \text{Mass}/\text{M. wt.} = 5.35/562 = 9.5 * 10^{-3} \text{ gram-mole}$$

(b) Using the relation, Mol. wt. = 2 * vapour density, we get

$$\text{Mol. wt.} = 2 * 38.3 = 76.6$$

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

Let weight of NO_2 in mixture = x g

Then weight of N_2O_4 in mixture = $100 - x$

No. of moles of NO_2 = Mass/Mol. wt. = $x/46$ (i)

No. of moles of N_2O_4 = Mass/Mol. wt. = $100 - x/92$ (ii)

According to problem

$$1.30 = x/46 + 100 - x/92$$

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

\therefore weight of $\text{NO}_2 = 20.1$ g

Moles of $\text{NO}_2 = \text{Mass/M. wt.} = 20.1/46 = 0.437$ moles.

Sol 6.

Volume of oxygen taken = 30 ml

Volume of unused oxygen = 15 ml

Volume of O_2 used = Volume of O_2 added – Volume of O_2 left

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

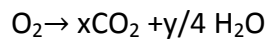
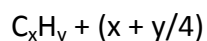
Volume of CO_2 produced

= Volume of gaseous mixture after explosion – Volume of unused oxygen

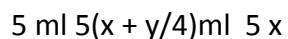
Or Volume of CO_2 produced = $25 - 15 = 10$ ml

Volume of hydrocarbon = 5 m;

General equation for combustion of a hydrocarbon is as follows –



(Hydrocarbon)



\therefore Volume of CO_2 produced = $5x$, Since Volume of $\text{CO}_2 = 10$ ml

$$\therefore 5x = \Rightarrow x = 2, \text{ Volume of } O_2 \text{ used} = \text{ml}$$

$$\therefore 5(x + y/4) = 15 \Rightarrow x + y/4 = 3$$

$$\Rightarrow 2 + y/4 = 3 \quad (\because x = 2)$$

$$\Rightarrow 8 + y = 12 \therefore y = 4$$

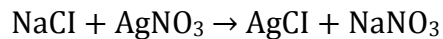
Sol 7.

(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

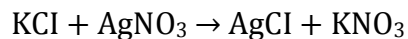


58.5 g 143.5 g

\therefore 58.5 g NaCl gives AgCl = 143.5 g

\therefore x g NaCl gives AgCl = $143.5/58.5 * x$ g

Again



74.5 g 143.5 g

\therefore 74.5 g KCl gives AgCl = 143.5 g

\therefore $(0.118 - x)$ g KCl gives AgCl = $(143.5/74.5 * 0.118 - x)$ g

Total weight of AgCl = 0.2451g

$\therefore (143.5/58.5 * x) + [143.5/74.5 * (0.118 - x)] = 0.2451$

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

\therefore Amount of NaCl in mixture = 0.0338g

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

Since $2\text{NaCl} \equiv \text{Na}_2\text{O}$

$2 * 58.5 \qquad \qquad \qquad 62$

$$= 117.0$$

\therefore 117g NaCl is equivalent = 62.0g Na₂O

\therefore 0.033g NaCl is equivalent to = $62.0/117 * 0.0338$ g Na₂O

$$= 0.0179\text{g}$$

% of Na₂O in 0.5g of feldspar = $0.0179/0.500 * 100 = 3.58\%$



$$2 * 774.5 = 149 \quad 94$$

\therefore 149g of KCl is equivalent to = 94g K₂O

\therefore 0.0842g of KCl is equivalent to $94/149 * 0.0842$

$$= 0.0531 \text{ g K}_2\text{O}$$

\therefore % of K₂O in 0.5 of feldspar = $0.0531/0.5 * 10.62\%$

% of Na₂O in feldspar = 3.58%

% of K₂O in feldspar = 10.62%

Sol 8.

According to problem, three atoms of M combine with 2 atoms of N

Equivalent wt of N = $14/3$ (\therefore valency of N in compound is 3)

\therefore 288 g n combines with = 72g metal

\therefore $14/3$ N combines with = $72/28 * 14/3 = 12$

\therefore Eq. wt. of metal = 12

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

Sol 9.

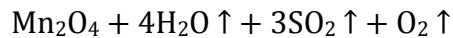
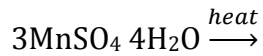
In the given reaction two hydrogen atoms of H₃PO₄ are replaced so the basicity of H₂PO₄ is 2.

Using the relation, eq. wt. of acid = Mol. wt. of acid/Basicity

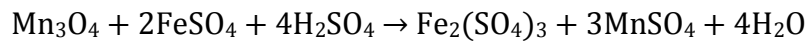
We get, eq. wt. of H₃PO₄ = Molwt/2 = $(1 * 3) + 31 + (4 * 16)/2 = 98/2 = \text{g}$

Sol 10.

Following reaction take place-



(residue)



Milliequivalents of FeSO_4 in 30 ml of 0.1N $\text{FeSO}_4 = 30 * 0.1 = 3 \text{ m.eq.}$

According to problem step (iv)

25 ml of KMnO_4 reacts with = 3 m eq of FeSO_4

Thus in step (iii) of the problems,

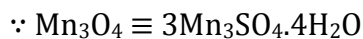
50 ml of KMnO_4 reacts with = $3/25 * 50 \text{ m. eq. of FeSO}_4 = 6 \text{ meq of FeSO}_4$

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

FeSO_4 which reacted with $\text{Mn}_3\text{O}_4 = (10-6) = 4 \text{ m eq.}$

Milli eq of $\text{FeSO}_4 =$ Millie q. of Mn_3O_4

(\because Millie q of oxidizing agent and reducing agent are equal)



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

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

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

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

Sol 11.

(a) $\text{CaCl}_2 \equiv \text{CaCO}_3 \equiv \text{MgCl}_2$

M. wt. 111 100 95

From this it is evident, the

111 mg CaCl_2 will give $\text{CaCO}_3 = 100\text{mg}$

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

95 mg MgCl_2 gives $\text{CaCO}_3 = 100\text{mg}$

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

\therefore Total CaCO_3 formed by 1 mg CaCl_2 and 1 mg $\text{MgCl}_2 = 0.90 + 1.05 = 1.95 \text{ mg}$

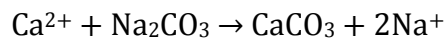
\therefore Amount of CaCO_3 present per litre of water = 1.95mg

\therefore wt of 1 ml of water = 1g = 10^3 mg

\therefore wt of 100 ml of water = $10^3 * 10^3 = 10^6\text{mg}$

\therefore Total hardness of water in terms of parts of CaCO_3 per 10^6 parts of water by weight = 1.95 parts.

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



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

1 milliequivalent of Na_2CO_3 is required to soften 1 litre of hard water.



$$2 * 24 = 48\text{g} \quad 32\text{g} \quad 2(24 + 16) = 80\text{g}$$

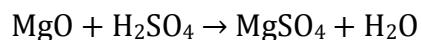
$\therefore 32\text{g}$ of O_2 reacts with = 48g Mg

$\therefore 0.5\text{g}$ of O_2 reacts with = $48/32 * 0.5 = 0.75\text{g}$

Weight of unreacted Mg = $1.00 - 0.75 = 0.25\text{g}$

Thus Mg is left in excess.

Weight of MgO formed = $80/48 * 0.75 = 1.25\text{g}$



(40g)

According to reaction

$\therefore 40\text{g}$ MgO is dissolved it gives 1000 ml of 1 N. H_2SO_4

∴ 40 g MgO is dissolved it gives 2000 ml 0.5 N H₂SO₄

∴ 1.25 MgO is dissolved it gives

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

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

Sol 12.

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

$PV = nRT$ (Ideal gas equation)

$$\text{Or } n = PV/RT = 1 * 1 / 0.082 * 400 = 0.0304$$

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

Element	Wt. of element	Relation no. of atom	Ratio of of atoms	Whole no. of atoms
C	10.5	$10.5/12$ $= 0.875$	$0.875/0.875$ $= 1$	$1 * 7 = 7$
H	1.0	$1.0/1 = 1$	$1/0.875 = 1.14$	$1.14 * 7 = 8$

∴ Empirical formula = C₇H₈

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

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

molecular formula = $n * \text{empirical formula}$

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

Sol 13.

(i) No. of C atom in 14g of ¹⁴C = $6.02 * 10^{23}$

∴ No. of C atom in 7 mg (7/1000g) of ¹⁴C

$$= 6.02 * 10^{23} * 7 / 14 * 1000 = 3.01 * 10^{20}$$

No. of neutrons in 1 carbon atom = 7

∴ Total no. of neutrons in 7 mg of ¹⁴C = $3.01 * 10^{20} * 7$

$$= 21.07 * 10^{20}$$

Wt of neutron = wt of 1 hydrogen atom

$$= 1/6.02 * 10^{23} \text{ g}$$

∴ Wt of $3.01 * 10^{20} * 7$ neutrons

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

Sol 14.

Weight of AgCl formed = 2.567 g

Amount of AgCl formed due to MCl = 1.341 g

(∵ NaCl does not decompose on heating to 300°C)

∴ Weight of AgCl formed due to NaCl

$$= 2.567 - 1.341 = 1.226 \text{ g}$$

NaCl ≡ AgCl ≡ MCl

58.5 143.5

{NaCl + AgNO₃ → AgCl + NaNO₃ MCl + AgNO₃ → AgCl + MNO₃}

∴ 143.5g of AgCl is obtained from NaCl = 58.5g

∴ 1.226 g of AgCl is obtained from NaCl

$$= 58/143.5 * 1.226 = 0.4997 \text{ g}$$

∴ Wt of MCl 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

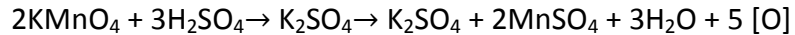
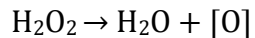
$$= 0.5003/1.341 * 143.5 * 53.53 \text{ g}$$

∴ Molecular weight of MCl = 53.53

Sol 15.

(i) Find equivalent wt. of H₂O₂ and KMnO₄.

(ii) X mol of N.KMnO₄ = X ml of N.H₂O₂



$$\therefore \text{Equivalent wt. of H}_2\text{O}_2 = \text{Mol. wt.}/2 = 34/2 = 17$$

$$\text{Equivalent wt. of KMnO}_4 = \text{Mol. wt.}/5 = 158/5 = 31.6$$

Let normality of KMnO_4 solution = N , then

$$X \text{ ml. of } N.\text{KMnO}_4 = X \text{ ml. of } N.\text{H}_2\text{O}_2$$

$$w = X * N * 17 / 1000 \text{ g of H}_2\text{O}_2$$

According to problem $X * 17 * N / 1000$ g of H_2O_2 is present in

$$1 \text{ g of solution, Also given } \dots\dots(1)$$

$$100 \text{ g of H}_2\text{O}_2 \text{ contain } X \text{ g of H}_2\text{O}_2$$

$$1 \text{ g of H}_2\text{O}_2 \text{ contain } = X/100 * 1 \text{ g of H}_2\text{O}_2 \dots\dots(2)$$

Comparing the two relation (1) and (2)

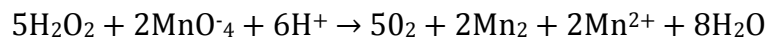
$$17 * N * X / 1000 = X / 100$$

$$\text{Or } N = X / 100 * 1000 / 17 \quad X = 1000 / 100 * 17 = 10 / 17$$

Hence, normality of KMnO_4 solution is $10/17$ *N or 0.588 N*

ALTERNATIVE SOLUTION:

The complete oxidation under acidic conditions can be represented as follows:



Since 34 g of $\text{H}_2\text{O}_2 = 2000$ ml of 1N . H_2O_2

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

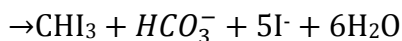
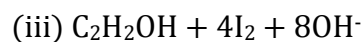
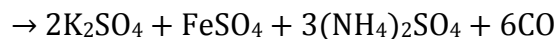
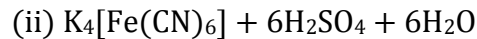
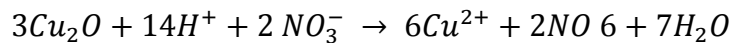
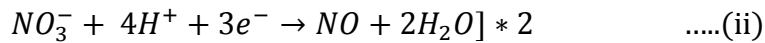
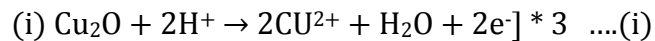
$$\therefore 34 \text{ g of H}_2\text{O}_2 = 2000 \text{ ml of } 1 \text{ N KMnO}_4 [\because N_1 V_1 = N_2 V_2]$$

$$\text{Or } X/100 \text{ g of H}_2\text{O}_2 = 2000 * X/100 * 34 \text{ ml of } 1 \text{ N KMnO}_4$$

Therefore the unknown normality = $2000 * X/34 * 100 * X = 10/17$ or 0.588 N

Sol 16.

Balance the reactions by ion electron method.

**Sol 17.**

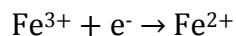
TIPS/Formulae:

(i) Write balanced reaction for changes.

(ii) M. eq. of Fe^{2+} formed = M. eq. KMnO_4 used

$$= V \text{ in ml} * \text{molarity} * x$$

The given redox changes are



[In NH_2OH , O.S. of N = - 1 and in N_2O , O.S. of N = + 1. For two molecules of NH_2OH , electron involved = 4]



$$= \text{Meq. Of } \text{KMnO}_4 \text{ used} = \text{Vol. in ml} * \text{Molarity} * x$$

$$= 12 * 0.02 * 5 = 12 \quad [50/10 = 5]$$

\therefore Meq. of Fe^{2+} formed by NH_2OH in 1000 ml of dil. Solution

$$= 1.2 * 1000/50 = 24$$

Meq. of original solution

= Meq. of NH_2OH in 1000 ml. of dilute solution

$$W_{\text{NH base } 2\text{OH}}/33/2 * 1000 = 24 \text{ [Eq. wt} = 33/2\text{]}$$

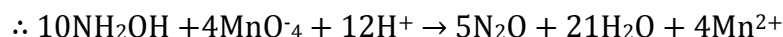
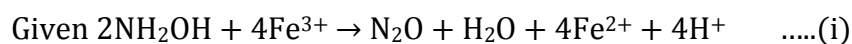
$$W_{\text{NH base } 2\text{OH}} = 24 * 16.5/1000 = 0.396 \text{ g}$$

\therefore Wt. of NH_2OH in 10 ml. of original solution = 0.396 g

\therefore Wt. of NH_2OH in 1 litre of original solution

$$= 0.396 * 100 = 39.6 \text{ g l}^{-1}$$

ALTERNATIVE SOLUTION:



[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 400 ml of 1M MnO_4^- would react with $\text{NH}_2\text{OH} = 330\text{g}$

\therefore 12 ml of 0.02 M KMnO_4 would react with $\text{NH}_2\text{OH} = 330 * 12 * 0.02/400 \text{ g}$

\therefore Amount of NH_2OH present in 1000 ml of dilute solution $330 * 12 * 0.02 * 1000/400 * 50 \text{ g}$

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

\therefore A amount of hydroxyl amine in one litre of original solution $= 330 * 0.02 * 12 * 1000/4000 * 50 * 1000/10 \text{ g} = 39.6 \text{ g}$

Sol 18.

TIPS/Formulae:

(i) Mole fraction = Moles of substance/Total moles

(ii) 1 mole of $\text{Na}_2\text{S}_2\text{O}_3$ gives 2 moles of Na^+ and 1 mole $\text{S}_2\text{O}_3^{2-}$

Molecular wt. of sodium thiosulphate solution ($\text{Na}_2\text{S}_2\text{O}_3$) = $23 * 2 + 32 * 2 + 16 * 3 = 158$

(i) The percentage by weight of $\text{Na}_2\text{S}_2\text{O}_3$

$$= \text{wt of } \text{Na}_2\text{S}_2\text{O}_3/\text{wt of solution} * 100 = 3 * 158 * 100/100 * 1.25 = 37.92$$

[wt. of $\text{Na}_2\text{S}_2\text{O}_3 = \text{Molarity} * \text{Molwt}$]

(ii) Mass of 1 litre solution = $1.25 * 1000 \text{ g} = 1250 \text{ g}$

[∵ density = 1.25g/l]

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

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

Moles of water = $1250 - 158 * 3 / 18 = 43.1$

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

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

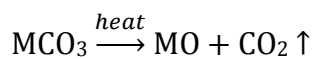
Molality of $\text{Na}_2\text{S}_2\text{O}_3 = 3 * 1000 / 776 = 3.87$

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

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

Sol 19.

One mole of metal carbonate, on heating, decomposes to form 1 mole of oxide with the evolution of 1 mole of CO_2 gas.



1 mole 1 mole 1 mole

Thus on heating, metal carbonate will lose weight corresponding to the weight of carbon dioxide.

From the given data, loss in wt. = $4.08 - 3.64 \text{ g} = 0.44 \text{ g}$

Conversion of 0.44 g of CO_2 into mole of CO_2

44 of $\text{CO}_2 = 1$ mole of CO_2

∴ $0.44 \text{ g of } \text{CO}_2 = 0.44 / 44 = 0.01$ mole

From the equation, it is also obvious that

1 mole of $\text{CO}_2 \equiv 1$ mole of $\text{MCO}_3 =$

$\therefore 0.01 \text{ mole of CO}_2 \equiv 0.01 \text{ mole of MCO}_3$

Thus the given mixture has 0.01 mole of MCO_3 which will yield 0.01 mole of MO.

From the problem, we also know that

16 ml of 2.5 N NaOH \equiv 16 ml of 2.5 n HCl [$N_1V_1 = N_2V_2$]

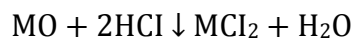
$= 16 * 2.5 \text{ ml of N HCl} = 40 \text{ ml of N HCl}$

$\therefore \text{volume of N HCl remain unused} = 40 \text{ ml}$

Total volume of N HCl added = 100 ml

$\therefore \text{Volume of N HCl used} = 100 - 40 = 60 \text{ ml}$

Hence 60 ml of N HCl is used in neutralizing BaO and MO according to the following reaction.



Thus 0.01 mole of MO \equiv 0.02 mole of HCl

$= 0.02 * 1000 \text{ ml of N HCl}$

$= 20 \text{ ml of N HCl}$

$\therefore \text{Vol. of N HCl used} = 60 - 20 = 40 \text{ ml}$

Eq. wt. of BaO = $138 + 16/2 = 77$

40 ml of N HCl = 40 ml of N BaO = $40 * 77/1000 = 3.08 \text{ g of BaO}$

Total wt. of the oxides = 3.64 g

$\therefore \text{Wt. of MO in the mixture} = 3.64 - 3.08 = 0.56 \text{ g}$

But we know that 0.01 mole of MO is present in the residue it means

0.01 mole of MO = 0.56 g of MO

1 mole of MO = 56 g of MO

Suppose the atomic weight of M = a

Then the mol. wt. of MO = a + 16

We also know that 1 mole of MO (i.e. Mol. wt. of Mo) = 56

$$\therefore a + 16 = 56 \text{ or } a = 56 - 16 = 40$$

$$\therefore \text{The atomic weight of metal M} = 40$$

Hence, the metal M must be calcium

ALTERNATIVE SOLUTION:

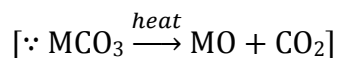
Weight of MCO_3 and $\text{BaO} = 4.08 \text{ g}$ (given)

Weight of residue = 3.64 g (given)

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

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

Number of moles of $\text{MCO}_3 \equiv 0.01 \text{ mole}$

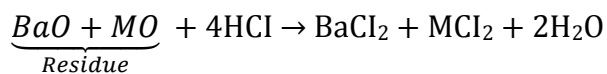


Volume of 1N HCl in which residue is dissolved = 100 ml

$$\text{Volume of 1N HCl used for dissolution} = (100 - 2.5 * 16) \text{ ml} = 60 \text{ ml}$$

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

The chemical equation for dissolution can be written as



$$[\text{Number of moles of BaO and MO} = 1 + 1 = 2]$$

$$\text{Number of moles of BaO} + \text{Number of moles of MO} = 0.06/2 = 0.03$$

$$\text{Number of moles of BaO} = (0.03 - 0.01)$$

$$= 0.02 \text{ moles}$$

$$\text{Molecular weight of BaO} = 138 + 16 = 154$$

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

$$= 3.08 \text{ g}$$

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

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

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

$$\text{Hence atomic weight of unknown M} = (100 - 60) = 40$$

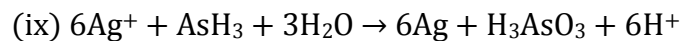
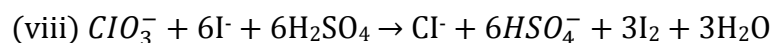
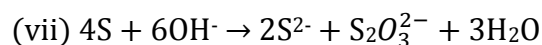
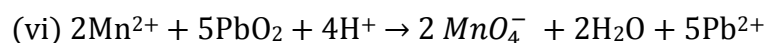
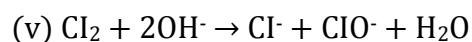
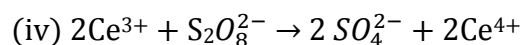
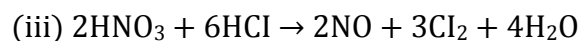
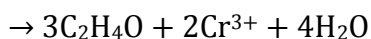
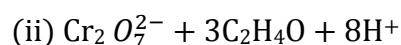
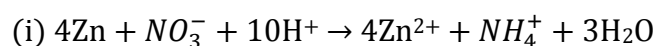
The atomic weight of metal is 40 so the metal M is Ca

Sol 20.

TIPS/Formulae:

Balanced the atoms as well as charges by ion electron/ oxidation number method.

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



Sol 21.

TIPS/Formulae:

Equivalents of A oxidized = Equivalents of A reduced. Since in acidic medium, A^{n+} is oxidized to 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 2.68×10^{-3} moles of A^{n+}

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

Thus the number of electrons added to reduce 1.61×10^{-3}

Moles of Mn^{+4} to Mn^{2+} , i.e.

$$(+7) \text{ to } (+2) = 1.61 \times 10^{-3} \times 5$$

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

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

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

Sol 22.

TIPS/Formulae:

(i) Find normality of acid mixture and $Na_2CO_3 \cdot 10H_2O$. Equate them to find volume of H_2SO_4 .

(ii) Meq. Of $H_2SO_4 = V \times N = V \times N / 1000$ eq.

(iii) Equivalent of $SO_4^{2-} = \text{Equivalents of } H_2SO_4 \times \text{Eq. wt. of } SO_4^{2-}$

$$N \times V(\text{ml.}) = \text{meq}$$

Acid mixture contains 5 ml of 8N, HNO_3 , 4.8 ml of 5N, HCl

and say, 'V' ml of 17 M \equiv 34 N, H_2SO_4 .

$$[1MH_2SO_4 = 2N.H_2SO_4]$$

N of the acid mixture = meq. (total) of acid/mls. Of solution

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

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

\therefore Eq. of wt. of $Na_2CO_3 \cdot 10H_2O = \text{Mol. wt.} / 2$

$$= 106 + 180 / 2 = 143$$

N of $Na_2CO_3 = \text{Meq. Of } Na_2CO_3 / \text{Volume of solution (ml)}$

$$= 1 / 143 \times 1000 / 100 = 1 / 143 \times 1000 / 100 = 0.069N$$

$$N_1V_1 = N_2V_2$$

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

(acid) (sod. carbonate)

$$\therefore N_{\text{mixture}} = 42.9 * 0.069 / 30 = 0.0986 \text{ N}$$

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

$$64 + 34 V = 0.0986 * 2000, 64 + 34 V = 197.2$$

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

$$\text{Hence meq. Of } H_2SO_4 = V * N \text{ of } H_2SO_4$$

$$= 3.9 * 34 = 132 \text{ meq.}$$

$$= 0.1326 \text{ eq. of } H_2SO_4$$

$$= 0.1326 \text{ eq. of } SO_4^{2-}$$

$$= 0.1326 * 48 \text{ g of } SO_4^{2-}$$

$$(\because \text{ Eq. wt. of } SO_4^{2-} = 32 + 64 / 2 = 48)$$

$$= 6.3648 \text{ g of } SO_4^{2-} \text{ are in 3.9 ml of 17M } H_2SO_4$$

Sol 23.

$HI < I_2 < ICl < HIO_4$; O.N. of I in $I_2 = 0$, $HI = -1$, $ICl = +1$, $HIO_4 = +7$.

Sol 24.

(i) From the given half-cell reaction,

$$\text{Here Eq. wt. of } NaBrO_3 = \text{Mol. wt.} / 6 = 151 / 6 = 25.17$$

[\because number of electron involved = 6]

Now we know that

$$\text{Meq.} = \text{Normally} * \text{Vol. in ml.} = 85.5 * 0.672 = 57.456$$

$$\text{Aslo Meq.} = W_{NaBrO_3} / \text{Eq. wt.}_{NaBrO_3} * 1000$$

$$= W_{NaBrO_3} / 25.17 * 1000$$

$$W_{NaBrO_3} / 25.17 * 1000 = 57.456 \text{ g}$$

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

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

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

(ii) From the given-cell reaction,

$$\text{Eq. wt. of } \text{NaBrO}_3 = \text{Mol. wt.}/5 = 151/5 = 30.2$$

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

Thus, the amount of NaBrO_3 required for preparing 1000 ml. of 1 N $\text{NaBrO}_3 = 30.2 \text{ g}$

\therefore The amount of NaBrO_3 required for preparing 85.5 ml of 0.672 N NaBrO_3 .

$$= 30.2 * 0.672 * 85.5/1000 = 1.7532 \text{ g}$$

$$\text{Hence, Molality} = 0.672/5 = 0.1344 \text{ M}$$

Sol 25.

(i) Molal concentration = Molality

$$= \text{Moles of solute/Mass of solvent in Kg}$$

(ii) Mole fraction of sugar

$$= \text{Moles of sugar/Total moles in solution}$$

(i) Weight of sugar syrup = 214.2 g

Weight of sugar in the syrup = 34.2 g

$$\therefore \text{Weight of water in the syrup} = 214.2 - 34.2 = 180.0 \text{ g}$$

Mol. wt. of sugar, $\text{C}_{12}\text{H}_{22}\text{O}_{11} = 342$

$$\therefore \text{Molal concentration} = 34.2 * 1000 / 342 * 180 = 0.56$$

(ii) Mol. wt. water, $\text{H}_2\text{O} = 18$

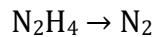
$$\therefore \text{Mole fraction of sugar} = 34.2/342 / 180/18 + 34.2/342$$

$$= 0.1/10 + 0.1 = 0.1/10.1 = 0.0099$$

Sol 26.

No. of equivalent of KMnO_4

= No. of equivalents of hydrazine sulphate.



Change in oxidation state for each $\text{N}_2\text{H}_4 = 2 * 2 = 4$

Equivalent weight of $\text{N}_2\text{H}_6\text{SO}_4 = 130/4 = 32.5$

Normality of $\text{KMnO}_4 = 5 * 450$ (\because valence factor = 5)

Number of equivalents of $\text{KMnO}_4 = 20 * 5/50 * 1000 = 1/500$ and if weight of hydrazine sulphate be x gm then equivalents of hydrazine sulphate = $x/32.5$

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

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

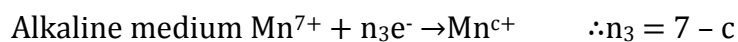
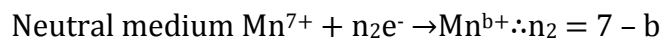
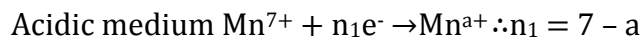
\therefore Wt. of $\text{N}_2\text{H}_6\text{SO}_4$ in 1000 ml solution = 6.5 g

Sol 27.**TIPS/Formulae:**

No. of equivalents of KMnO_4 in neutral medium

= No. of equivalents of reducing agent.

Assuming that KMnO_4 shows the following changes during its oxidising nature.



Let V ml. of reducing agent be used for KMnO_4 in different medium.

\therefore Meq. of reducing agent

= Meq. of KMnO_4 in acid medium

Meq. of KMnO_4 in neutral medium

= Meq. of KMnO_4 in alkaline medium

$$= 1 * n_1 * 20 = 1 * n_2 * 33.4 = 1 * n_3 * 100$$

$$= n_1 = 1.667 n_2 = 5 n_3$$

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 $\text{K}_2\text{Cr}_2\text{O}_7$

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

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

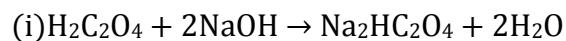
Sol 28.

TIPS/Formulae:

No. of equivalent of KMnO_4

= No. of equivalent of reducing agents.

Case I. Reaction of NaOH with $\text{H}_2\text{C}_2\text{O}_4$ and NaHC_2O_4 .



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

\therefore Combined normality of the mixture titrated with NaOH

$$= 0.3/10 = 0.03$$

Case II. Reaction of C_2O_4^- ion and KMnO_4

(iii) $5\text{C}_2\text{O}_4^{2-} + \text{MnO}_4^- + 16\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 10\text{CO}_2 + 8\text{H}_2\text{O}$ KMnO_4 will react in same manner with both NaHC_2O_4 and $\text{H}_2\text{C}_2\text{O}_4$ as it can be seen from the above reaction.

Number of milliequivalents of $\text{KMnO}_4 = 4.0 * 0.1 = 0.4$

\therefore Combined normality of the mixture titrated with KMnO_4

$= 0.4/10 = 0.04$

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 $\text{NaHC}_2\text{O}_4 = 112$

Amount of NaHC_2O_4 in one litre of solution formed

$= 0.01 * 112 = 1.12\text{ g}$ and amount of $\text{H}_2\text{C}_2\text{O}_4$

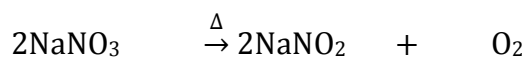
$= 2.02 - \text{Wt. of NaHC}_2\text{O}_4 = 2.02 - 1.12 = 0.90\text{ g}$

Sol 29.

Let the amount of NaNO_3 in the mixture = $x\text{ g}$

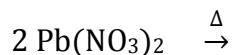
\therefore The amount of $\text{Pb}(\text{NO}_3)_2$ in the mixture = $(5 - x)\text{ g}$

Heating effect of sodium nitrate and lead nitrate



$2(23+14+48) = 170\text{ g}$

$2 * 16 = 32\text{ g}$



$2(207 + 28 + 96) = 662\text{ g}$



$\frac{4(14 + 32) = 184\text{ g} \quad 2 * 16 = 32\text{ g}}{216\text{ g}}$

Now since, 170 g of NaNO_3 gives $= 32\text{ g}$ of O_2

$\therefore x\text{ g}$ of NaNO_3 gives $= 32/170 * x\text{ g}$ of O_2

Similarly, 662 g of $\text{Pb}(\text{NO}_3)_2$ gives = 216 g of gases

$(5 - x)$ g of $\text{Pb}(\text{NO}_3)_2$ gives = $216/662 * (5 - x)$ g of gases

$(\text{NO}_2 + \text{O}_2)$

Actual loss, on heating, is 28% of 5 g of mixture

$$= 5 * 28/100 = 1.4 \text{ g}$$

$$\therefore 32x/170 + 216/662 * (5 - x) = 1.4$$

$$32x * 662 + 216(5 - x) * 170 = 1.4 * 170 * 662$$

$$21184x + 183600 - 36720x = 157556$$

$$-15536x = -26044$$

$$X = 1.676 \text{ g}$$

$$\text{Wt. of } \text{NaNO}_3 = 1.676 \text{ g}$$

$$\text{And Wt. of } \text{Pb}(\text{NO}_3)_2 = 5 - 1.676 \text{ g} = 3.324 \text{ g}$$

Sol 30.

Molality = Mass of solute/M. wt. of solute/Mass of solvent in kg

Mass of H_2SO_4 in 100 ml of 93% H_2SO_4 solution = 93 g

\therefore Mass of H_2SO_4 in 1000 ml of the H_2SO_4 solution = 930 g

Mass of 1000 ml H_2SO_4 solution = $1000 * 1.84 = 1840$ g

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

$$= 910 \text{ g} = 0.910 \text{ kg}$$

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

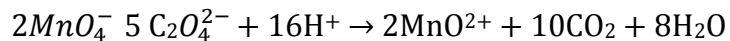
\therefore Moles of H_2SO_4 in 1 kg of water

$$= 930/98 * 1/0.910 = 10.43 \text{ mol}$$

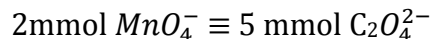
\therefore Molality of solution = 10.43m

Sol 31.

In the given problem, a solution containing Cu^{2+} and $\text{C}_2\text{O}_4^{2-}$ is titrated first with KMnO_4 and then with $\text{Na}_2\text{S}_2\text{O}_3$ in presence of KI . In titration with KMnO_4 , it is the $\text{C}_2\text{O}_4^{2-}$ ions that react with the C_2MnO_4^- ions. The concerned balanced below may be written as given below.



Thus according to the above reaction



However,

No. of mmol of MnO_4^- used in titration = Vol. in ml * M

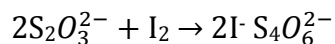
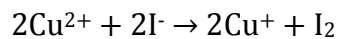
$$= 22.6 * 0.02 = 0.452\text{ mmol MnO}_4^-$$

since $2\text{ mmol MnO}_4^- \equiv \text{mmol C}_2\text{O}_4^{2-}$

$$0.452\text{ mmol MnO}_4^- \equiv 5/2 * 0.452 = 1.130\text{ mmol C}_2\text{O}_4^{2-}$$

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

Here 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

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

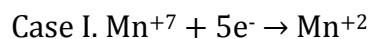
Now since $2\text{ mmol S}_2\text{O}_3^{2-} \equiv 2\text{ mmol Cu}^{2+}$ [From above equation]

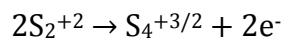
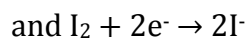
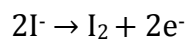
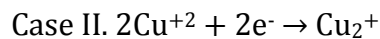
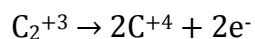
$$0.565\text{ mmol S}_2\text{O}_3^{2-} = 2/2 * 0.565\text{ mmol Cu}^{2+}$$

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

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

Balanced equations in two cases





Sol 32.

Mass of Fe_2O_3 in the sample = $55.2/100 * 1 = 0.552$ g

Number of moles of $Fe_2O_3 = 0.552/159.8 = 3.454 * 10^{-3}$

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

$$= 6.9 * 10^{-3} \text{ mol} = 6.90 \text{ mmol}$$

Since its only 1 electron is exchanged in the conversion of Fe^{3+} to Fe^{2+} , the molecular mass is the same as equivalent mass.

\therefore Amount of Fe^{2+} ion in 100 ml. of sol. = 6.90 meq

Volume of oxidant used for 100 ml of Fe^{2+} sol

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

Amount of oxidation used = $68 * 0.0167$ mmol

$$= 1.1356 \text{ mmol}$$

Let the number of electrons taken by the oxidant = n

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

$$\text{Thus } 1.1356 * n = 6.90$$

$$N = 6.90/1.1356 = 6$$

Sol 33.

1.5 g of sample require = 150 ml. of M/10 HCl

\therefore 2 g of sample require = $150/1.5$ ml of M/10 HCl

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

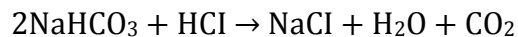
On heating, the sample, only NaHCO_3 undergoes decomposition as given below.



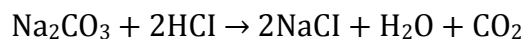
2 moles 1 mole 1 mole

2 eq

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



1 eq 1 eq



1 mole 1 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 CO_2 formed, i.e.

$$N = PV/RT = 750/760 * 123.9/0.082 = 0.005$$

Moles of NaHCO_3 in the sample (2 g) $= 2 * 0.005 = 0.01$

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

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

% of $\text{NaHCO}_3 = 0.84 * 100/2 = 42\%$

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

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

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

\therefore % of Na_2SO_3 in the mixture $= 100 - (42 + 26.5) = 31.5\%$

Sol 34.

Reaction involved titration is



1 mole 2 mole

20 ml. of stock KI solution \equiv 30 ml. of M/10 KIO_3 solution

Molarity of KI solution = $30 * 1 * 2 / 20 * 10 = 3/10$

Millimoles in 50 ml. of KI solution = $50 * 3/10 = 15$

Millimoles of KI left unreacted with AgNO_3 solution

= $2 * 50 * 1/10 = 10$

\therefore millimoles of KI reacted with $\text{AgNO}_3 = 15 - 10 = 5$

Millimoles of AgNO_3 present in AgNO_3 solution = 5

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

% AgNO_3 in the sample = $0.850/1 * 100 = 85\%$

Sol 35.

Calculation of number of moles in 45 ml. of 0.025 M $\text{Pb}(\text{NO}_3)_2$

Moles of $\text{Pb}(\text{NO}_3)_2 = 0.025 * 45/1000 = 0.01125$

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

Moles of $\text{NO}_3^- = 0.01125 * 2 = 0.02250$ [1 mole $\text{Pb}(\text{NO}_3)_2 \equiv 2$ moles of NO_3^-]

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

Moles of chromic sulphate ($\text{Cr}_2(\text{SO}_4)_3 = 0.1 * 25/1000 = 0.0025$ moles.

Moles of $\text{SO}_4^{2-} = 0.0025 * 3 = 0.0075$ [1 Mole of chromic sulphate $\equiv 3$ moles of SO_4^{2-}]

Moles of PbSO_4 formed = 0.0075 [SO_4^{2-} is totally consumed]

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

Moles of NO_3^- left = 0.02250 [NO_3^- remain untreated]

Moles of chromium ions = $0.0025 * 2 = 0.005$

Total volume of the solution = $45 + 25 = 70 \text{ ml.}$

\therefore Molar concentration of the species left

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

$$(ii) NO_3^- = 0.0225/70 * 1000 = 0.3214 \text{ M}$$

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

Sol 36.

In pure iron oxide (FeO), iron and oxygen are present in the ratio 1 : 1.

However, here number of Fe^{2+} present = 0.93

Or No. of Fe^{2+} ions missing = 0.07

Since each Fe^{2+} ion has 2 positive charge, the total number of charge due to missing (0.07)

$$Fe^{2+} \text{ ions} = 0.07 * 2 = 0.14$$

To maintain electrical neutrality, 0.14 positive charge is compensated by the presence of Fe^{3+} ions. Now since, replacement of one Fe^{2+} ion by one Fe^{3+} ion increases one positive charge, 0.14 positive charge must be compensated by the presence of 0.14 Fe^{3+} ions.

In short, 0.93 Fe^{2+} ions have 0.14 Fe^{3+} ions

$$100 \text{ } Fe^{2+} \text{ ions have} = 0.14/0.93 * 100 = 15.05\%$$

Sol 37.

(i) Density = Mass/Volume

(ii) Molarity = Moles of solute/Volume of solution in L

(iii) Molality = Moles of solute/Mass of solvent in kg

(iv) Mole fraction of solute = Moles of solute/Total moles

The formula of Glauber's salt is $Na_2SO_4 \cdot 10H_2O$

Molecular mass of $Na_2SO_4 \cdot 10H_2O$

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

Weight of the Glauber's salt taken = 80.575 gm

Out of 80.575 g of salt, weight of anhydrous Na_2SO_4

$$= 142.1/322.3 * 80.575 = 35.525 \text{ g}$$

Number of moles of Na_2SO_4 per dm^3 of the solution

$$= 35.525/142.1 = 0.25$$

Molarity of the solution = 0.25 M

Density of solution = 1077.2 kgm⁻³

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

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

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

Weight of water = 1077.2 - 35.252 = 1041.67 g

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

Number of moles of water in the solution = $1041.67 / 18 = 57.87$

Mole fraction of Na₂SO₄

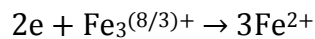
$$= \text{No. of moles of Na}_2\text{SO}_4 / \text{Total number of moles} = 0.25 / 0.25 + 57.87$$

$$= 0.0043 = 4.3 * 10^{-3}$$

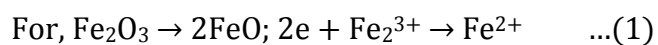
Sol 38.

Find the milliequivalents and equate them as per data given in question.

For Fe₃O₄ → 3FeO



Thus, valence factor for Fe₃O₄ is 2 and for FeO is 2/3



Thus valence factor for Fe₃O₃ is 2 and for FeO is 1.

Let Meq. of Fe₃O₄ and Fe₃O₃ be *a* and *b* respectively.

∴ Meq. of Fe₃O₄ + Meq. Fe₃O₃ = Meq. of I₂ liberated

= Meq. of hypo used

$$A + b = 11 * 0.5 * 100 / 20 = 27.5$$

Now, the Fe²⁺ ions are again oxidized to Fe³⁺ by KMnO₄.

Note that in the change



Thus,

$$\begin{aligned} & \text{Meq. of Fe}^{2+} \text{ (from Fe}_3\text{O}_4\text{)} + \text{Meq. of Fe}^{2+} \text{ (from Fe}_3\text{O}_3\text{)} \\ & = \text{Meq. of KMnO}_4 \text{ used} \quad \dots(2) \end{aligned}$$

If valence factor for Fe^{2+} is $2/3$ from Eq. (1),

$$\text{Then Meq. of Fe}^{2+} \text{ (from Fe}_3\text{O}_4\text{)} = a$$

If valence factor for Fe^{2+} is 1

$$\text{Then Meq. of Fe}^{2+} \text{ (from Fe}_3\text{O}_4\text{)} = 3a/2 \quad \dots(3)$$

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

$$\therefore 3a/2 + b = 0.25 * 5 * 12.8 * 100/50 = 32$$

$$\text{or } 3a + 2b = 64$$

from Eqs (3) and (4)

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

$$\therefore W_{\text{Fe base 30 base 3}} = 9 * 232/2 * 1000 = 1.044 \text{ g}$$

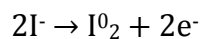
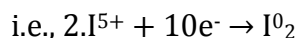
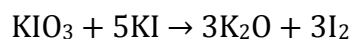
$$\text{And } W_{\text{Fe base 30 base 3}} = 18.5 * 160/2 * 1000 = 1.48 \text{ g}$$

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

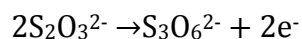
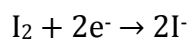
$$\text{and } \% \text{ of Fe}_3\text{O}_3 = 1.48 * 100/3 = 49.33$$

Sol 39.

Write the reactions taking place, balance them and equate moles of I_2 and $\text{Na}_2\text{S}_2\text{O}_3$.



Now liberated I_2 reacts with $\text{Na}_2\text{S}_2\text{O}_3$



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

Thus, m mole of I_2 liberated

$$= m \text{ mole of } Na_2S_4O_6 \text{ used} * 1/2 = 45 * M * 1/2$$

[M is molarity of thiosulphate]

$$\text{Also } m \text{ mole of } KIO_3 = 0.1/214 * 1000$$

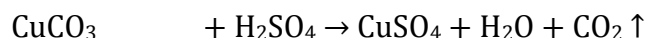
Now m mole ratio of $KIO_3 : I_2 = 1 : 3$

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

$$\therefore M = 0.1 * 100 * 3 * 2 / 214 * 45 = 0.062$$

Sol 40.

Use molarity equation to find volume of H_2SO_4 solution.



$$63.5 + 12 + 48.98 \quad 98g$$

$$= 123.5 g$$

\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.

$$= 98 * 0.5 / 123.5 g = 0.39676 g H_2SO_4$$

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

Weight of solute in grams

$$= \text{Mol. wt.} * \text{Molarity} * \text{Volume in mL} / 1000$$

$$0.39676 = 98 * 0.5 * V / 1000$$

$$\text{or } V = 0.39676 * 1000 / 98 * 0.5 \text{ ml}$$

$$\text{volume of } H_2SO_4 \text{ solution} = 8.097 \text{ ml}$$

Sol 41.

(i) Volume of virus = $\pi r^2 \ell$ (Volume of cylinder)

(ii) Mass of single virus = Volume / Sp. Volume

(iii) Molecular mass of virus

$$= \text{Mass of single virus} * 6.02 * 10^{23}$$

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

$$= \frac{22}{7} * \frac{150}{2} * \frac{150}{2} * 10^{-16} * 500 * 10^{-8}$$

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

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

$$= 1.178 * 10^{-19} \text{ g}$$

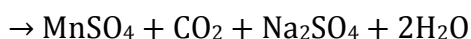
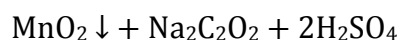
$$\therefore \text{Mol. wt. of virus} = 1.178 * 10^{-19} * 6.02 * 10^{23}$$

$$= 7.09 * 10^7$$

Sol 42.

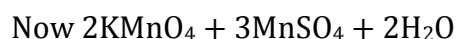
Write the balanced chemical reaction for change and apply mole concept.

The given reactions are :



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

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



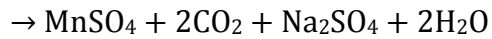
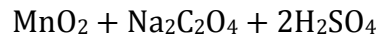
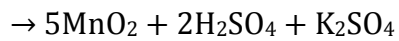
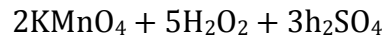
Since eq. wt. of MnO₂ is derived from KMnO₄ and MnSO₄ both, thus it is better to proceed by mole concept

$$\text{mM of KMnO}_4 \equiv \text{mM of MnI}_2 * \left(\frac{2}{5}\right) = \frac{4}{5}$$



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

$$\therefore M_{\text{H}_2\text{O}_2} * 20 = 2 \text{ or } M_{\text{H}_2\text{O}_2} = 0.1$$



Sol 43.

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

$\equiv 1000/18 = 55.55$ moles of water

So, molarity of water = 55.55M