# TOPIC : REDOX REACTION (EQUIVALNET CONCEPT) EXERCISE # 1

#### SECTION (A)

1.	$2(+1) + 2 x = 0 \therefore x = -1$					
2.	Oxidation state of F may be zero and $-1$ . Oxidation state of CI may be $-1$ to $+7$ Oxidation state of He is zero Oxidation state of Na may be 0 and $+1$					
3.2	x + 4 (-2) = 0 x = +8					
4.22	x + 4 (+1) = +1 x = -3					
5.	$I_{3}^{-}$ $3(x) = -1$ $x = -\frac{1}{3}$					
6.24	$[Co(NO_2)_6]^{3-}$ x + 6 (-1) = - 3 x = +3					
7.	Phosphorous acid $HO - P - OH$ oxidation state of P = + 3 H H O					
	Orthophosphoric acid $HO - P - OH$ oxidation state of P = + 5					
	Hypophosphorous acid $H - P - OH$ oxidation state of P = + 1 H					
	Metaphosphoric acid $  I  $ OH HO $/P O$ oxidation state of P = + 5 HO $// O$ OH					
8.	2(+2) + 2x + 7 (-2) = 0 : $x = +5$					
9.	NH_OH					

- $NH_{2}OH$ x + 3(+1) + 1(-2) = 0 x = -1
- **10.** The oxidation states show a change only in reaction

#### SECTION (B)

**1.** A reducing agent itself oxidises & reduces others. So an element shows increase in oxidation number in it & loss of electron occurs.

- 2.a Br<sub>2</sub> undergoes disproportionation, i.e. it undergoes both oxidation & reduction.
- **3.** Soxidises from +2 to + 2.5 and I reduces from 0 to 1. Hence  $3^{rd}$  reaction is a redox reaction.
- **4.** Increase in oxidation number is oxidation & decrease in oxidation number is reduction. Here,

 $Zn \rightarrow Zn^{2+}$  (oxidation)  $Cu^{2+} \rightarrow Cu$  (reduction)

**5.** In 4<sup>th</sup> reaction, N undergoes oxidation while Cr undergoes reduction.

6.  $X^- + XO_3^- + H^+ \longrightarrow X_2 + H_2O$ V.f. = 1 V.f. = 5  $\therefore$  Molar ratio = 5 : 1

- 7. In  $H_2O_2$  oxidation state of oxygen is -1. It can undergo both oxidation as well as reduction  $O^- - e^- \rightarrow O$  (Oxidation)  $O^- + e^- \rightarrow O^{2-}$  (reduction) Hence it can act both as oxidizing as well as reducing agent.
- 8. SO<sub>2</sub> Here S is in +4 state

It can be oxidized to +6 state as also reduced to -2 state.

$$9.2$$

$$Mn_{2}O_{3} \xleftarrow{+4e^{-}} [KMnO_{4}] \xrightarrow{+7} [MnO_{4}] \xrightarrow{+e^{-}} [MnO_{4}]^{2-}$$

$$Mn^{2+} MnO_{2}$$

#### SECTION (C)

1.2 
$$N_2 \longrightarrow NH_3$$
  

$$0 -3$$
  

$$|\_\_\_\_]$$
  

$$n = 3 \times 2 = 6$$
  
Eq.wt = mol.wt./ n factor =  $\frac{28}{6} = 4.67$ 

**2.** 
$$BrO_3 \longrightarrow Br_2$$
  
+5 0  $\therefore$  (V.f.)  $BrO_3 = 5$   $\therefore$  Eq wt = M/5

3. In acidic medium KMnO<sub>4</sub> shows following reaction –

$$\begin{array}{l}
\overset{+}}{\mathsf{Mn}} O_4^- + 8\mathsf{H}^+ + e^- \to \mathsf{Mn}^{2+} + 4\mathsf{H}_2\mathsf{O} \\
\mathsf{V}.\mathsf{f.} = +5 \\
\mathsf{Equivalent weight} = \frac{\mathsf{molecular wt.}}{\mathsf{v.f.}} = \frac{\mathsf{M}}{5} \\
\mathbf{4.} \qquad \overset{\mathsf{NH}_2\mathsf{OH}}{\longrightarrow} \overset{\mathsf{N}_2\mathsf{O}}{\overset{-1}{\longrightarrow} +1} \\
\overset{-}{\ldots} \mathsf{V}.\mathsf{f. of NH}_2\mathsf{OH} = 2 \quad \therefore \qquad \mathsf{Eq wt} = \mathsf{M}/2 \\
\mathbf{5.} \qquad \overset{\mathsf{SO}_2}{\mathsf{SO}_2} \quad \text{is reducing agent} \\
O_3 \quad \text{is oxidising agent} \\
O_3 \quad \longrightarrow O_2 + \mathsf{O}^{2-} \\
(\mathsf{v.f.} = 2) \\
\mathsf{Equivalent weight} = \frac{\mathsf{M}}{2} \\
\mathbf{6} \qquad \qquad \overset{\mathsf{molecular wt.}}{\mathsf{Mn}} \\
\mathbf{6} \qquad \qquad \overset{\mathsf{molecular wt.}}{\mathsf{Mn}} \\
\end{array}$$

**6.** Equivalent wt of acid =  $\frac{1}{\text{No. of H}^+ \text{ replaced per acid molecule}}$ 

	Equivalent wt. = $\frac{98}{1} = 1$					
7.	Equivalent wt. = $\frac{M}{2}$ l.e., valency factor = 2 MnSO, has valence factor 2 in the following reaction.					
	$\therefore \xrightarrow{^{+2}}{MnSO_4} \longrightarrow \xrightarrow{^{+4}}{MnO_2}$ v.f. = 2					
8.	Valency factor = 1					
	Equivalent weight = $\frac{M}{1}$					
9.2	$HNO_{3} \longrightarrow NH_{3}^{-3} \therefore V.f. \text{ of } HNO_{3} = 8$ Eq. wt. = M/8.					
10.	$Na_2 \stackrel{_{+2}}{S}_2 O_3 \longrightarrow Na_2 \stackrel{_{+6}}{S} O_4$					
	the total change in oxidation number = $4 \times 2 = 8$ $\therefore$ $E_{Na_2S_2O_3} = \frac{mol. wt.}{V. f} = \frac{M}{8}$					
11.	Normality = Molarity × v.f. $\therefore$ 1M H <sub>3</sub> PO <sub>4</sub> = 3N H <sub>3</sub> PO <sub>4</sub>					
12.	In CO (CO ) 16 g O = 12 g C or $8 g O = 6 g C$ In CO <sub>2</sub> (CO <sub>2</sub> )					
	$32 \text{ g O} = 12 \text{ g C}$ $\therefore$ $8 \text{ g O} = 3 \text{ g C}$ $\therefore$ Ratio $6:3 = 2:1$ .					
13.১	V.F. of $As_2S_3$ is 28. Reason of this is change in oxidation state of both As and S. V.F. = $2(5-3) + 3[6 - (-2)] = 4 + 24 = 28$ . $\therefore$ Eq. wt = M/28					
SECTION	ν (D)					
1.	m. eq. of $H_2SO_4 = m$ . eq. of $Na_2CO_3$					
	$0.1 \times \frac{v}{1000} = \frac{0.123}{106} \times 2$					
	V = 23.6 mL					
2.	Eq. of $Na_2CO_3 = Eq. of H_2SO_4$					
	$\frac{1.06}{106} \times 2 = \frac{25}{1000} \times N$					
	N = 0.8 N					
3.2	22400 mL volume contains = 1 mole gas					
	$\therefore$ 224 mL volume contains = $\frac{1}{22400} \times 224 = \frac{1}{100}$ mole CO <sub>2</sub>					
	Eq of $CO_2 = Eq$ of HCl					
	$\frac{1}{100} \times 2 = \frac{200}{1000} \times N$					
	N = 0.1 N					
4.	$E_{MCl_2} = E_{M+2} + E_{Cl^-} = 32.7 + 35.5 = 68.2$					
	$\therefore \qquad \text{Molecular mass} = 2 \times 68.2 = 136.4$					
5.2	(1) Explanation : No. of meq of $H^+ = 10 \times 1 + 20 \times 2 = 50$ [:: $H_2SO_4$ , N = 2 M]					
	No. of meq of $OH^- = 30 \times 1 = 30$ No. of meq of $H^+$ left unreacted = 50 - 30 = 20 meq					

Hence, (1) is correct, (2), (3) and (4) are ruled out. SECTION (E) 1. 1.5 N H<sub>2</sub>O<sub>2</sub> VS = ?VS = Normality × 5.6 = 1.5 × 5.6 = 8.4 V mole of  $CaCO_3 = 2/100 = 0.02$ 2.2 so mole of CaO =  $\frac{1}{2} \times 0.02 = 0.01$ so mass of CaO = 0.01 × 56 = 0.56 g  $H_2C_2O_4$ .  $2H_2O = 2 + 24 + 64 + 36 = 126$  and Equivalent wt. =  $\left[\frac{126}{2}\right]$ 3.  $0.2 = \frac{W \times 1000}{\left(\frac{126}{2}\right) \times 50}$ ∴ W = 0.63 g NaOH 4.2  $H_2C_2O_4$  $M = \frac{40 \times 1000}{1000}$  $M = \frac{63 \times 1000}{126 \times 100}$ 40×100 M = 5M = 10N = 10 N = 10V = 125 mL V = 125 mL  $N_1V_1 = 1250$  milli equivalent  $N_2V_2 = 1250$  milli equivalent so, resulting solution will be neutral. SECTION (F) m eq of  $KMnO_4 = 0.1 \times 5 \times V = 0.5 V$ 1.2  $\& m eq K_2 Cr_2 O_7 = 0.1 \times 6 \times V = 0.6 V$ So, K<sub>2</sub>Cr<sub>2</sub>O<sub>7</sub> will oxidise more Fe<sup>2+</sup> 2. Milli equivalents of  $FeC_2O_4 = 0.1 \times 3 \times 25 = 7.5$ From choice (4), milli equivalents of  $KMnO_4 = 0.1 \times 5 \times 15 = 7.5$  $\therefore$  m. eq. of FeC<sub>2</sub>O<sub>4</sub> = m. eq. of KMnO<sub>4</sub> 3.  $1.68 \times 10^{-3} \times 6 = 3.36 \times 10^{-3} \times X$ x = 3So, oxidation number of A increases by 3. .:. New oxidation number of A = -n + 3 = 3 - n. (1) Formula of silver chromate (VI) will be Ag<sub>2</sub> CrO<sub>4</sub> 4.2 (2) Minimum mass of zinc required for reduction of 0.1 mole of  $Cr^{3+}$  to  $Cr^{2+} = \frac{0.1}{2}$  moles of Zn  $=\frac{6.54}{2}$  g = 3.27 g  $(v.f. of Zn = 2 \& v.f. of Cr^{3+} = 1)$ (3)  $CrO_{4^{-}} \longrightarrow Cr_{2}O_{7^{-}}$  in both ions, chromium is in +6 oxidation state. (4) Given reaction is correct. Equivalents of  $C_2 O_4^{2-}$  = equivalents of  $MnO_4^{-}$ 5.2  $x(mole) \times 2 = 1 \times 5$  $x = \frac{5}{2}$  mole EXERCISE # 2 PART - I

- 1.∞ In 1<sup>st</sup> reaction, oxisdation number of N is + 4, in 2<sup>nd</sup> reaction it is -3 & in 3<sup>rd</sup> reaction it is +5, none of which changes during reaction.
  2. Valency factor ratio is inversely related to molar ratio. (V.f.)HI: (V.f.)HNO<sub>3</sub> = 1 : 3 = 2 : 6 ∴ Molar ratio = 6 : 2
  3. MnO<sub>4</sub><sup>-</sup> + C<sub>2</sub>O<sub>4</sub><sup>2-</sup> + H<sup>+</sup> Mn<sup>2+</sup> + CO<sub>2</sub> + H<sub>2</sub>O V.f. = 5 V.f. = 2 ∴ Balanced equation : 2MnO<sub>4</sub><sup>-</sup> + 5C<sub>2</sub>O<sub>4</sub><sup>2-</sup> + 16H<sup>+</sup> → 2Mn<sup>2+</sup> + 10CO<sub>2</sub> + 8H<sub>2</sub>O
  4. Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup> + SO<sub>2</sub> → Cr<sup>3+</sup> + (<sup>+6)</sup>/<sub>SO<sub>3</sub></sub>
  5.∞ SO<sub>2</sub> + (<sup>+5)</sup>/<sub>O<sub>3</sub><sup>-</sup></sub> → SO<sub>3</sub> + I<sup>-</sup>
- 6. For a completely balanced equation, net charge on reactant side & product side must be equal. -1 + 4 + x (-1) = 0x = +3
- 7. In Mg<sub>3</sub>N<sub>2</sub> 3(+2) +2 (x) = 0 2x = -6x = -3
- 8. Equivalent of  $K_2Cr_2O_7$  = equivalent of  $N_2H_4$ also equivalent of KMnO<sub>4</sub> = equivalent of  $N_2H_4$ So, equivalent of  $K_2Cr_2O_7$  = equivalent of KMnO<sub>4</sub>  $0.1 \times 6 \times V_1 = 0.3 \times 5 \times V_2$ so  $V_2$  = 2/5 V<sub>1</sub>

9.2 2Li + 
$$\mathring{H}_2 \longrightarrow 2Li\overset{-1}{H}$$
  
H itself is reduced, so it act as oxidizing agent.

10. Reduction Na + H<sub>2</sub>O  $\longrightarrow$  NaOH +  $\frac{1}{2}$ H<sub>2</sub> Oxidation

**11.** In KMnO<sub>4</sub>,  $K_2Cr_2O_7 \& C_2H_5OH$ Mn(+7), Cr(+6), C(+4) are present in their max. oxidation state, so they cannot be further oxidised but in  $H_2S$ , S(-2) is present in its minimum oxidation state so it can be oxidised by SO<sub>2</sub>.

**12.** 
$$\begin{array}{c} \overset{(+2)}{\text{CaCO}}_3 \longrightarrow \overset{(+2)}{\text{CaO}} + \overset{(+4)}{\text{CO}}_2 \end{array}$$

There is no change is oxidation state of any element so it is not a redox reaction.

13.	(+4) NO <sub>2</sub>	&	$N_{2}^{(+4)}O_{4}$	difference = 0		
	$P_{2}^{+5}O_{5}$	&	${}^{+5}P_4O_{10}$	difference = 0		
	$\overset{+1}{N_2}O$	&	+2 NO	difference = 1		
	$\overset{_{+4}}{\mathrm{SO}}_2$	&	*6 SO <sub>3</sub>	difference = 2		
14.æ	$K_{2}Cr_{2}O_{7} +$		D <sub>4</sub> + 3SO <sub>2</sub> —	$\rightarrow \text{K}_2\text{SO}_4 + \text{Cr}_2(\text{SO}_4)_3 + \text{H}_2\text{O}_4$		
15.	Here Ni is oxidised to Ni <sup>2+</sup> So it acts as reducing agent					

16. +7 +6  
KMnQ<sub>4</sub> → K<sub>2</sub>MnQ<sub>4</sub>  
v.t. = 1  
equivalent wt. = 
$$\frac{M}{1}$$
 = M.  
17. 2H, S + SO<sub>2</sub> → 3S + 2H<sub>2</sub>O  
S in SO<sub>2</sub> is reduced from +4 to 0 state. Therefore it act as oxidising agent.  
18. 4 $K^{C}O_3 \rightarrow 3K^{C}O_4 K^{C}I$   
 $SO_2 + H_5^{C}S \rightarrow 2H_2O + 3$   
 $BaO_4 + H_5^{C}S \rightarrow 2H_2O + 3$   
 $C^{C}S \rightarrow 2H_2O + 3I_5 + 4H_5O \rightarrow 2R^{O_5} + 6I + 8H^+$   
N is oxidised so it acts a reducing agent.  
Reduction  
(oxidation (reducing agent))  
20.  $2FeCI_5 + H_5^{C}S \rightarrow 2FeCI_5 + 2HCI + 5$   
 $Cxidation (reducing agent)$   
21.  $Fe^{t^2} + (\frac{+7}{1})^{-1} \longrightarrow Mn^{t^2} + Fe^{t^2}$   
Reduction  
22. In formation of nitrogen oxides from N<sub>2</sub> & O<sub>2</sub>  
N is oxidised & O is reduced.  
 $H^{C}S (-H_5) = H^{C}S - H^{C}S + H^+ \longrightarrow Mn^{2+} + CO_2 + H_2O$   
 $V, I_1 = 5 V, I_1 = 2$   
 $\therefore$  Balanced equation :  $2MnO_4 + 5C_2O_4^{2-} + 16H^+ \longrightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$   
25.  $M^{H+} - 3e^- \longrightarrow M^{H+}$   
26.  $[Fe(CO_3)] : x + 5(0) = 0 \implies x = 0$   
27. SRP values for KMnO<sub>4</sub> & K<sub>2</sub>Cr<sub>2</sub>O<sub>4</sub> are 1.57 V & 1.33 V respectively so greater the SRP better agent.  
28.  $MnO_4^- + C_2O_4^{2-} + H^+ \longrightarrow Mn^{2+} + CO_2 + H_2O$   
 $V, I_1 = 5 V, I_1 = 2$   
 $\therefore$  Balanced equation :  $2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \longrightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$   
29.  $C I^+ + MnO_4^- \longrightarrow Mn^{2+} + CI_2$ 

oxidising

Eq of  $Cl_2 = Eq$  of  $KMnO_4$ 2 [mole of  $Cl_2$ ] = 5  $\left[\frac{10}{158}\right]$ mole of  $CI_2 = \frac{50}{2 \times 158} = 0.15823$  mole volume of Cl<sub>2</sub> at STP = 0.15823 × 22.4 = 3.54 L PART - II 1. When O. No. decreases i.e. reduced or oxidizing agent. 2.  $Cu^+ \longrightarrow Cu^{2+}$  (Oxidation)  $Cu^+ \longrightarrow Cu$  (Oxidation) For disproportion oxidation & reduction both occur which is not possible for F2. 3. 5. Sn  $(0 \rightarrow + 4)$ It undergoes oxidation, so it is the reducing agent.  $S_2O_3^{2-} + 6H^+ + 4e^- \longrightarrow 2S + 3H_2O.$  $MnO_4^- \longrightarrow Mn^{2+}$ 6. (+7)(+2) MnO<sub>4</sub> - acts as oxidising agent. *.*.. meq of NaOH = N × V = 0.1 × 200 = 20 ; meq of  $H_2SO_4 = N \times V = 0.1 \times 200 = 20$ . 7. Resulting solution is neutral. *.*.. EXERICSE # 3 PART - I 3. Number of equivalents of  $KMnO_4$  = Number of equivalents of  $H_2O_2$ Moles of  $KMnO_4 \times Valency$  factor = Moles of  $H_2O_2 \times Valency$  factor  $H_2O_2 = 1 \times \frac{5}{2} = \frac{5}{2}$ Moles of 5. In alkaline medium,  $MnO_4^- + I^- \longrightarrow MnO_2^- + IO_3^-$ (v.f.=3) (v.f.=6) Moles of  $KMnO_4 \times v.f. = Moles of KI \times v.f.$ Moles of KMnO<sub>4</sub> =  $\frac{1 \times 6}{3}$  = 2. 6. In acidic medium MnO<sub>4</sub><sup>-</sup> oxidises ferrous oxalate as follows :  $2\mathsf{MnO}_{4}^{-} + 5\mathsf{C}_{2}\mathsf{O}_{4}^{2-} + 16\mathsf{H}^{+} \longrightarrow 2\mathsf{Mn}^{2+} + 10\mathsf{CO}_{2} + 8\mathsf{H}_{2}\mathsf{O}$ 5 moles of oxalate ions are oxidised by 2 moles of MnO<sub>4</sub>-. ÷ 1 mole of oxalate ion is oxidised by =  $\frac{2}{5}$  mole of MnO<sub>4</sub><sup>-</sup> = 0.4 mole of MnO<sub>4</sub><sup>-</sup> *.*..  $\begin{array}{c} \text{Cl}_2\\ (0) \end{array}$  + OH<sup>-</sup>  $\longrightarrow \begin{array}{c} \text{NaCl}\\ (-1) \end{array}$  +  $\begin{array}{c} \text{NaClO}_3\\ (+5) \end{array}$ 7. Compound Oxidation number of nitrogen 8.  $N_2H_4$ -2 = NH<sub>3</sub> -3 = N<sub>3</sub>H -1/3 = NH<sub>2</sub>OH -1 =

- 9.  $KCIO_3 + H_2C_2O_4 + H_2SO_4 \longrightarrow K_2SO_4 + KCI + CO_2 + H_2O$ Maximum change in oxidation number is observed in CI (+5 to -1).
- **10.**  $O_3$  is reduced into  $O^{-2}$  ion and  $Ag_2O$  is reduced to Ag so  $H_2O_2$  is reducing agent in both (a) and (b)
- 12. 1st reaction is not a redox reaction as the oxidation number of elements remains unchanged.
- KMnO<sub>4</sub> is an oxidising so it can oxidise SO<sub>2</sub> readily.
   KMnO<sub>4</sub> + SO<sub>2</sub> ? Mn<sup>2+</sup> + SO<sub>3</sub>
   NO<sub>2</sub> is strong oxidising agent, CO<sub>2</sub> is neither oxidising agent nor reducing agent,
- 14.  $MnO_{4}^{+7} + C_{2}O_{4}^{+3} + H^{+} \longrightarrow Mn^{+2} + CO_{2}^{+4} + H_{2}O$ Reduction half = [MnO<sub>4</sub><sup>-</sup> + 8H<sup>+</sup> + 5e<sup>-</sup>  $\longrightarrow Mn^{2+} + 4H_{2}O$ ] x 2 Oxidation Half = [C<sub>2</sub>O<sub>4</sub><sup>2-</sup>  $\longrightarrow 2CO_{2} + 2e^{-}$ ] x 5
  - PART II
- $^{+5}_{I}$  O  $^{+-}_{3}$  + I<sup>-</sup> + H<sup>+</sup>  $\longrightarrow$  H<sub>2</sub>O +  $^{0}_{I_2}$ 1. (i) Oxidation half cell (1) Balancing the numbers of atoms  $2I^{-} \longrightarrow I_{2}$ (2) Balancing charge  $2I^{-} \longrightarrow I_{2} + 2e^{-}$ ... (1) (ii) (1) Reduction half reaction  $IO_3^- + H^+ \longrightarrow H_2O + I_2$ (1) Balancing number of atoms  $2IO_3^- + 12H^+ \longrightarrow 6H_2O + I_2$ (2) Balancing the charge  $2IO_3^- 12H^+ + 10 \longrightarrow 6H_2O + I_2$ ... (2) Multiplying Eq. (1) by 5 and adding it to Eq. (2)  $2I^{-} \longrightarrow I_{2} + 2 \times 5$  $2IO_3^- + 10I^- \longrightarrow 6H_2O + I_2$  $2IO_3^- + 10I^- + 12H^+ \longrightarrow 6I_2 + 6H_2O$  $IO_3^- + 5I^- + 6H^+ \longrightarrow 3I_2 + 3H_2O$ or Hence, a = 5, b = 6, c = 3, d = 3
- 2. Potassium dichromate is used in volumetric analysis mostly in oxidation titration because it is not deliquescent. It forms compound with other elements and precipitate out. So the assertion is true but the reason is false.
- 3. Potassium dichromate react with alcohol and the reduction of potassium dichromate takes place and dichromate (orange red) change of  $Cr^{3+}$  (green)  $Cr_2O^{2-}_7 + 14H^+ + 6e^- 2Cr^{3+} + 7H_2O$ So, the reason is false.
- 4. Oxidation state of I in  $HIO_4$  is + 7 as : 1 + x + 4 (-2) = 0 x = + 7Oxidation state of I in  $H_3IO_5$  is + 7 as. 3 + x + 5 (-2) = 0x = + 7

Oxidation state of I in  $H_5IO_6$  is + 7 as 5 + x + 6 (-2) = 0

- 5. Acidity is defined as the number of replaceable hydroxyl group is one molecule of the base.
- 6. The reaction of oxidation of ferrous oxalate by potassium dichromate in acidic medium is written as  $2FeC_2O_4 + Cr_2O_7^{2-} + 14H^- \longrightarrow 2Fe^+ + 2Cr^{3+} + 4CO_2 + 7H_2O$ 
  - $\therefore$  2 moles of FeC<sub>2</sub>O<sub>4</sub> are oxidised by = 1 mole of Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup>
  - $\therefore$  1 moles of FeC<sub>2</sub>O<sub>4</sub> will be oxidised by = 1/2 = 0.5 mole of Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup>
- 7. The balanced chemical reaction is as  $3UO^{2+} + Cr_2O_7^{2-} + 8H^+ \longrightarrow 3UO_2^{2+} + 2Cr^{3+} 4H_2O$ Hence, the value of x, y and z are respectively 3, 8, 2.
- 8. Meq of A = Meq of B Meq of KMnO<sub>4</sub> = 20×0.5=10 Meq of 50 ml of 0.1 M H<sub>2</sub>C<sub>2</sub> O<sub>4</sub> = 50×0.2 =10  $(0.1 \text{ M H}_2 \text{ C}_2 \text{ O}_4 = 0.2 \text{ N H}_2 \text{ C}_2 \text{ O}_4)$
- 9.  $[Fe^{2+} \rightarrow Fe^{3+} + e^{-}] \times 6$ From Mohr's salt  $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$   $6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$ Millimoles of  $Fe^{2+} = 750 \times 0.6 = 450$ Moles of  $Fe^{2+} = \frac{450}{1000} = 0.450 = 0.450$  moles.  $6 \text{ mol } Fe^{2+} = 1 \text{ mol } Cr_2O_7^{2-} \therefore 0.450 \text{ mole } Fe^{2+} = \frac{0.450}{6} = 0.075 \text{ mol } Cr_2O_7^{2-} = 0.075 \times 294 \text{ g} = 22.05 \text{ g}$ 11 MnO<sub>2</sub> + NaCl  $\xrightarrow{H_2SO_4} MnCl_2 + Cl_2$

### PART - III

- 2. Dilute nitric acid converts chromate into dichromate and H<sub>2</sub>O.  $2K_{2}CrO_{4} + 2HNO_{3} \longrightarrow K_{2}Cr_{2}O_{7} + 2KNO_{3} + H_{2}O.$   $\stackrel{+3}{Mn_{2}O_{3}} \xleftarrow{+4e^{-}} [KMnO_{4}] \xrightarrow{+e^{-}} [MnO_{4}]^{2-}$   $\stackrel{+3e^{-}}{\swarrow} \xleftarrow{+4e^{-}} [KMnO_{2}]^{2-}$   $\stackrel{+6}{\swarrow} \xleftarrow{+5e^{-}} \xleftarrow{+4e^{-}} MnO_{2}$ 3.
- 4. HCl reduces  $MnO_4^-$  to  $Mn^{2+}$  and itself oxidises to  $Cl_2$ .

- 5. Weight of organic compound = 29.5 mg  $NH_{2} + HCI \longrightarrow NH_{4}CI$ HCI (remaining) + NaOH -----> NaCI + H<sub>a</sub>O (1.5 m mole) Total milimole of HCI = 2mili mole of HCl used by  $NH_2 = 2 - 1.5 = 0.5$ mili mole of  $NH_3 = 0.5$ weight of  $NH_3 = 0.5 \times 17 \text{ mg} = 8.5 \text{ mg}$ weight of nitrogen =  $\frac{14}{17} \times 8.5$  mg = 7 mg ∴ % of Nitrogen in compound =  $\frac{7}{29.5}$  × 100 = 23.7 %.  $MnO_4^-$  +  $C_2O_4^{2-}$  +  $H^+$   $\longrightarrow$   $Mn^2$  +  $CO_2$  +  $H_2O$ 6. vf = 1(7 - 2)vf = 2(3-2)= 5 = 2 **Balanced Equation :**  $2\mathsf{MnO}_{4}^{-} + 5\mathsf{C}_{2}\mathsf{O}_{4}^{2-} + 16 \mathsf{H}^{+} \longrightarrow 2\mathsf{Mn}^{2+} + 10 \mathsf{CO}_{2} + 8\mathsf{H}_{2}\mathsf{O}$ So, x = 2, y = 5 & z = 16. 7. Mass of organic compound = 1.4 g let it contain x mmole of N atom. organic compound  $\longrightarrow NH_3$ x m mole  $2NH_3 + H_2SO_4 \longrightarrow (NH_4)_2 SO_4$ . (1st) 6 mmole initially taken.  $H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$  (2nd) 2 mmole reacted Hence m moles of  $H_2SO_4$  reacted in 2nd equation = 1  $\Rightarrow$  m moles of H<sub>2</sub>SO<sub>4</sub> reacted from 1st equation = 6 - 1 = 5 m moles  $\Rightarrow$  m moles of NH<sub>3</sub> in 1st equation = 2 x 5 = 10 m moles  $\Rightarrow$  m moles of N atom in the organic compound = 10 m moles  $\Rightarrow$  mass of N = 10 x 10<sup>-3</sup> x 14 = 0.14 g  $\Rightarrow$  % of N =  $\frac{0.14}{1.4} \times 100 = 10$  %  $Na_{2}C_{2}O_{4} + H_{2}SO_{4} \longrightarrow Na_{2}SO_{4} + H_{2}O + CO + CO_{2}$ 8.  $Na_{2}C_{2}O_{4} + CaCl_{2} \longrightarrow CaC_{2}O_{4} \downarrow + 2NaCl$  $8H_2SO_4 + 5CaC_2O_4 + 2KMnO_4 \longrightarrow 5CaSO_4 + K_2SO_4 + 2MnSO_4 + 8H_2O_4 + 10CO_2$ 9. As 1L solution have 10<sup>-3</sup> mol CaSO<sub>4</sub> Eq. of  $CaSO_4 = eq.$  of  $CaCO_3$ in 1L solution  $n_{CaSO_4} \times v.f. = n_{CaCO_3} \times v.f.$  $10^{-3}\times 2=n_{CaCO_3}\times 2$  $n_{CaCO_3} = 10^{-3}$  mol in 1L  $\therefore w_{CaCO_3} = 100 \times 10^{-3}$  g in 1L solution  $\therefore$  hardness in terms of CaCO<sub>3</sub>  $= \frac{w_{CaCO_3}}{w_{Total}} \times 10^6 = \frac{100 \times 10^{-3} \text{g}}{1000 \text{g}} \times 10^6 = 100 \text{ppm}$
- **10.** Volume strength =  $11.35 \times M = 11.35$  (STP)