

## CHEMISTRY - MODULE - 6

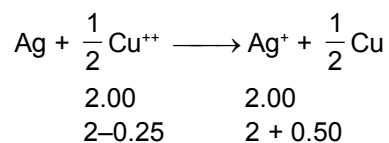
### ELECTROCHEMISTRY & SOLUBILITY OF SALTS

#### Single Correct Type Solution

6.  $\lambda_{\text{Ag}^+} = 62.3 \text{ Scm}^2 \text{ mol}^{-1}$ ,  $\lambda_{\text{Cl}^-} = 67.7 \text{ Scm}^2 \text{ mole}^{-1}$   
 $K_{\text{AgCl}} = 3.4 \times 10^{-6} \text{ Scm}^{-1}$   
 $\Lambda_{\text{AgCl}}^\infty = (62.3 + 67.5) = \frac{1000 \times 3.4 \times 10^{-6}}{S}$   
 $S = \frac{3.4 \times 10^{-3}}{(62.3 + 67.5)} = 2.6 \times 10^{-5} \text{ M}$
7. The S.R.P. of  $\text{Ag}^+$  is higher than that of  $\text{Ni}^{2+}$  hence it is a better oxidising agent. Like wise the S.O.P. of Ni is higher than that of Ag hence it is a better reducing agent.
11.  $\text{SrF}_2 \rightleftharpoons \text{Sr}^{2+} + 2\text{F}^-$        $K_{\text{sp}} = 7.9 \times 10^{-11}$   
 $\text{NaF} \rightleftharpoons \text{Na}^+ + \text{F}^-$        $[\text{F}^-] = 0.01$   
 Then  $7.9 \times 10^{-11} = x [0.01]^2$   
 $x \approx 7.9 \times 10^{-7} \text{ M}$ .
13. pH changes from 0 to 7.  
 $\therefore [\text{H}^+]$  changes from 1 to  $10^{-7} \text{ M}$ .  
 Accordingly  $E_{\text{red}}$  decreases by  $0.059 \log 10^{-7}$  i.e.  $0.059 \times (-7)$   
 $= -0.41 \text{ volt}$ .
14.  $m \text{ (theoretical)} = \frac{63.5 \times 0.1 \times 7200}{96500} = 0.4738 \text{ g}$   
 $\therefore \% \text{ efficiency} = \frac{0.3745}{0.4738} \times 100 = 79 \%$
15.  $R = \frac{1}{k} \frac{\ell}{A}$   
 The k is halved while the A is doubled. Hence R remains 50  $\Omega$ .  
 on substituting the value in above relation we get **B.E.<sub>(C=C)</sub> = 122 Kcal / mol**
16.  $2\text{CH}_2 = \text{CHCN} + 2\text{H}^+ + 2\text{e}^- \longrightarrow \text{CN} - (\text{CH}_2)_4 \text{CN}$   
 $m = \frac{\left(\frac{108}{2}\right) \times 10^{-3} \times 3000 \times 9.65 \times 3600}{96500} = 58.32 \text{ kg}$   
**Ans. 58 kg**
17.  $\Lambda_m = \Lambda_m^\infty - b\sqrt{C}$   
 $260 = \Lambda_m^\infty - 0.5 b$  ..... (1)  
 $250 = \Lambda_m^\infty - b$  ..... (2)  
 On solving (1) & (2), we get  
 $\Lambda_m^\infty = 270 \Omega^{-1} \text{ cm}^2 \text{ mol}^{-1}$

18.  $Q = 10 \times 4825$   
 $= 48250 \text{ C}$

$$\text{no. of mole} = \frac{48250}{96500} = 0.5$$

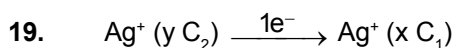


$$E_{\text{cell}} = E^{\circ}_{\text{Cell}} - \frac{0.0591}{1} \log \frac{[\text{Ag}^+]}{[\text{Cu}^{++}]^{1/2}}$$

$$E_1 = E^{\circ}_{\text{Cell}} - \frac{0.0591}{1} \log \frac{2.00}{(2.00)^{1/2}}$$

$$E_2 = E^{\circ}_{\text{Cell}} - \frac{0.0591}{1} \log \frac{2.50}{(1.75)^{1/2}}$$

$$\begin{aligned} \Delta E = E_2 - E_1 &= \frac{0.0591}{1} \left[ \log \sqrt{2} - \log \frac{2.50}{\sqrt{1.75}} \right] \\ &= \frac{0.0591}{1} [\log 1.41 - \log 1.88] \\ &= \frac{0.0591}{1} [0.1492 - 0.2742] \\ &= - \frac{0.0591}{1} \times 0.125 \\ &= - .00738 \text{ V} \end{aligned}$$



$$E = \frac{0.059}{1} \log \frac{y \text{ C}_2}{x \text{ C}_1}$$

$$E = 0.059 \log \frac{10 y}{x}$$

$$\log \frac{10 y}{x} = \frac{E}{0.059}$$

$$\log 10 + \log \frac{y}{x} = \frac{E}{0.059}$$

$$\log \frac{y}{x} = \frac{E}{0.059} - 1$$

$$20. \quad \frac{m_X}{m_Y} = \frac{\frac{A_X \times Q}{2}}{\frac{A_Y \times Q}{1}} \Rightarrow \frac{m_X}{m_Y} = 1 \quad \therefore \quad A_X = 2A_Y$$

$$21. \quad [Ag^+] = \frac{10^{-10}}{10^{-1}} = 10^{-9} \text{ M for AgCl ppt}$$

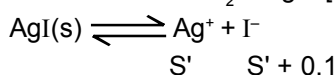
$$[Ag^+] = \frac{10^{-13}}{10^{-2}} = 10^{-11} \text{ M for AgBr ppt}$$

$$[Ag^+] = \frac{10^{-17}}{10^{-3}} = 10^{-14} \text{ M for AgI}$$

22. Let solubility of AgI in water is 'S' mole/lit.

$$\therefore K_{sp} = S^2 \quad \dots\dots\dots (1)$$

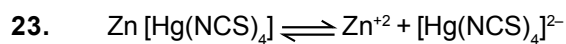
Now in 0.05 M BaI<sub>2</sub> we get [I<sup>-</sup>] = 0.05 x 2 = 0.1 M



$$K_{sp} = S'(S' + 0.1) = S' \times 0.1$$

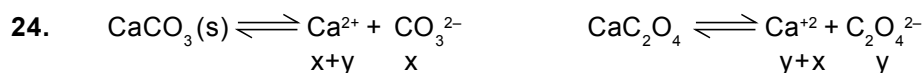
$$S^2 = 1 \times 10^{-15} \times 0.1$$

$$S = 10^{-8} \text{ M}$$



$$\Rightarrow K_{sp} = S^2$$

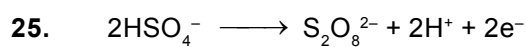
$$\Rightarrow S = \sqrt{22 \times 10^{-8}} = 4.69 \times 10^{-4} \text{ mol/L}$$



$$K_{sp_1} = (x + y) x$$

$$K_{sp_2} = (y + x) y$$

$$\text{Given } (x + y) = 8.426 \times 10^{-5} \text{ and } \frac{K_{sp_1}}{K_{sp_2}} = \frac{x}{y} = 2.087$$



$$\frac{2F}{0.75} \longrightarrow 1 \text{ mole } S_2O_8^{2-}$$

$$\therefore I = \frac{\frac{2}{0.75} \times 96500}{3600} = 71.5 \text{ amp}$$

26. The solubility of AgBr in presence of 10<sup>-7</sup> molar AgNO<sub>3</sub> is 3 × 10<sup>-7</sup> M.  
Therefore [Br<sup>-</sup>] = 3 × 10<sup>-4</sup> m<sup>3</sup>, [Ag<sup>+</sup>] = 4 × 10<sup>-4</sup> m<sup>3</sup> and [NO<sub>3</sub><sup>-</sup>] = 10<sup>-4</sup> m<sup>3</sup>

$$\text{Therefore } \kappa_{total} = \kappa_{Br^-} + \kappa_{Ag^+} + \kappa_{NO_3^-} = 39 \text{ Sm}^{-1}$$

27. 90 gm Hg has 10 gm Na

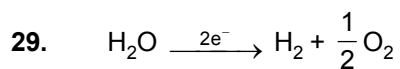
$$\therefore 10 \text{ gm Hg} = \frac{10}{90} \times 10 = \frac{10}{9} \text{ gm Na}$$

$$\therefore \text{weight of Na} = \frac{M}{n} \times \frac{i \times t}{96500}$$

$$\frac{10}{9} = \frac{23}{1} \times \frac{10 \times t}{96500} \quad [ \because \text{Na}^+ + e \rightarrow \text{Na} ]$$

$$\therefore t = \frac{10 \times 96500}{9 \times 10 \times 23} = 7.77 \text{ min}$$

**28.** Electrode potential values depend on reference electrode chosen but not cell potential

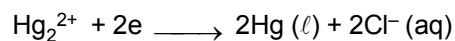
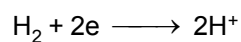


$$2 \times 96500 \text{ C} \rightarrow 22.4 \times 1000 \text{ cm}^3 \text{ H}_2$$

$$\frac{2 \times 96500 \times 224}{22.4 \times 1000} \text{ C} \leftarrow \therefore 224 \text{ cm}^3 \text{ H}_2$$

$$= 1930 \text{ C}$$

**30.** Considering the cell reaction



$$\Delta S = nF = \left( \frac{\partial E}{\partial T} \right)_p = 2 \times 96500 \times 3.4 \times 10^{-4} = 65.223 \text{ J/K/mole}$$

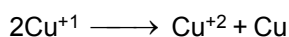
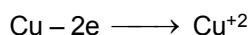
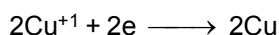
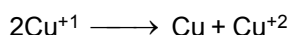
## CHEMISTRY MODULE -6 - MULTIPLE CHOICE QUESTIONS

### SOLUTIONS

#### Electrochemistry & Solubility of Salts

1. (A, B, D)

2. (A,C)



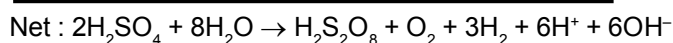
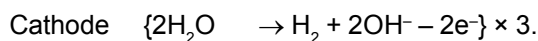
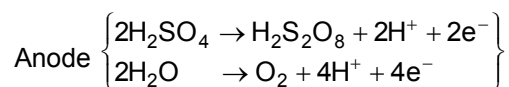
$$\therefore E^{\circ} = \frac{2 \times 0.521 + 2(-0.337)}{2} = 0.184$$

3. (A,B,C,D)

4. (B,C,D)

The SRP should be high for a species to be a good oxidising agent.

5. (A,B)



Hence ratio of  $n_{\text{O}_2}$  and  $n_{\text{H}_2}$  is 1 : 3.

6. (B,C)

7. (A,B)

$$K_{\text{sp}} = 1.1 \times 10^{-11} = (1.4 \times 10^{-4})^{x+y} x^x \cdot y^y$$

so we have  $x + y = 3$  (by comparing values)

$$\text{so, } x^x \cdot y^y = \frac{1.1 \times 10^{-11}}{1.4 \times 1.4 \times 1.4 \times 10^{-12}} = \frac{110}{1.96 \times 1.4} = 4$$

$$\text{Hence } x = 1, y = 2 \quad \text{or} \\ y = 1, x = 2$$

8. (A,B,D)

Arrange these values in decreasing order of SRP and then it can be seen that  $\text{Sn}^{+2}$ ,  $\text{H}^+$  and  $\text{Cu}^{+2}$  can oxidize Fe but not Hg.

9. (C)

$\text{CH}_3\text{COOH} + \text{NaOH} \longrightarrow \text{Na}^+ + \text{CH}_3\text{COO}^- + \text{H}_2\text{O}$  to conductance 1<sup>st</sup> increases slowly since no. of ions increases after end point it increases sharply due to  $\text{OH}^-$  ions.

10. (A, B, C) Reduction Potential of Cu is higher than that of Zn.

11. **(B,C)**  
Lower S.R.P. containing ion can displace higher S.R.P. containing ion.

12. **(B,C)**

13. **(A,B)**  
The SOP should be high for a species to be a good reducing agent

14. **(B,C,D)**  
 $\text{Fe}^{2+} + 2\text{e}^- \longrightarrow \text{Fe [in FeSO}_4\text{]}; \quad \text{Fe}^{3+} + 3\text{e}^- \longrightarrow \text{Fe [in Fe}_2\text{(SO}_4\text{)}_3\text{]}$   
 $\text{Fe}^{3+} + 3\text{e}^- \longrightarrow \text{Fe [in Fe (NO}_3\text{)}_3\text{]}$

$$\text{Amount of Fe deposited in FeSO}_4 = \frac{Q}{96500} \times \frac{56}{2}$$

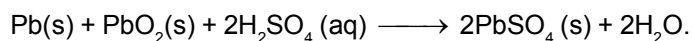
$$\text{Amount of Fe deposited in Fe}_2\text{(SO}_4\text{)}_3 = \frac{Q}{96500} \times \frac{56}{3}$$

15. **(A,C,D)**  
Resistance of cell is not due to vibrations of ion but actually it is due to collisions of ions.

16. **(A,C,D)**  
(A) because  $E_{\text{Cu}^{2+}/\text{Cu}}^\circ > E_{\text{Fe}^{2+}/\text{Fe}}^\circ$ .

17. **(B,C)**  
Reduction and electronation take place at cathode electrode, so it become positive electrode.

18. **(A,D)**  
Discharging reaction



19. **(A,B)**

20. **(B,D)**

## CHEMISTRY MODULE-6 – PARAGRAPH TYPE QUESTIONS

### SOLUTIONS

#### Electrochemistry & Solubility of Salts

1.  $[\text{Sr}^{2+}]_i = 0.0011 = 11 \times 10^{-4} \text{ M}$   
 $[\text{Sr}^{2+}]_f = 2 \times 10^{-4} \text{ M}$   
 $\therefore [\text{Sr}^{2+}] \text{ precipitated} = (11 - 2) \times 10^{-4} \text{ M}$   
 $= 9 \times 10^{-4} \text{ M}$   
 $[\text{F}^-] \text{ needed for this precipitation} = 2 \times 9 \times 10^{-4} = 18 \times 10^{-4} \text{ M}$   
Also,  $[\text{Sr}^{2+}]_f [\text{F}^-]_f^2 = K_{sp} = 2.048 \times 10^{-9}$   
But,  $[\text{Sr}^{2+}]_f = 2 \times 10^{-4} \text{ M}$   
 $\therefore [\text{F}^-]_f = 3.2 \times 10^{-3} \text{ M}$   
 $\therefore \text{Total } [\text{F}^-] \text{ needed} = 3.2 \times 10^{-3} + 18 \times 10^{-4} = 5 \times 10^{-3} \text{ M}$   
 $\therefore \text{NaF needed for 100 ml solution} = \frac{5 \times 10^{-3} \times 42}{10} = \mathbf{0.021 \text{ g}}$

2.  $\text{SrF}_2 \longrightarrow \underset{s}{\text{Sr}^{2+}} + \underset{2s}{2\text{F}^-}$   
 $\text{F}^-$  will react with  $\text{H}^+$  to produce HF  
 $\text{F}^- + \text{H}^+ \rightleftharpoons \text{HF}$   
 $K = \frac{1}{K_a} = \frac{[\text{HF}]}{[\text{H}^+][\text{F}^-]} = \frac{7}{10^{-5}}$   
 $\therefore [\text{HF}] = 7 \times 10^5 [\text{F}^-] [\text{H}^+] \quad (\because [\text{H}^+] = 10^{-5}; \text{pH}=5)$   
 $= 7 \times 10^5 [\text{F}^-] \times 10^{-5} = 7 [\text{F}^-]$   
Here  $[\text{F}^-] + [\text{HF}] = 2s$   
 $\therefore [\text{F}^-] = \frac{s}{4}$   
 $K_{sp} = s \left(\frac{s}{4}\right)^2 = 2.048 \times 10^{-9}$   
 $\therefore s = 3.2 \times 10^{-3} \text{ mol/L}$

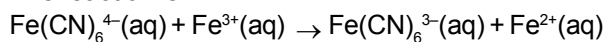
3. On increasing concentration of  $[\text{H}^+]$  ions the solubility of basic hydroxide  $\text{Fe}(\text{OH})_2$ , will increase.

4.  $[\text{OH}^-] = \sqrt{1.62 \times 10^{-5} \times 5} = 9 \times 10^{-3} \text{ M}$   
 $\text{Al}(\text{OH})_3 + \text{OH}^- \rightleftharpoons \underset{\text{aq.}}{[\text{Al}(\text{OH})_4]^-} \quad K = 40$   
 $(s) \quad \text{aq.}$   
 $\therefore 40 = \frac{[\text{Al}(\text{OH})_4]^-}{9 \times 10^{-3}}$   
 $\therefore [\text{Al}(\text{OH})_4]^- = 0.36 \text{ M}$   
 $\therefore \text{mgms of Al}(\text{OH})_3 \text{ dissolved} = \frac{0.36 \times 78 \times 1000}{10}$   
 $= 2808$

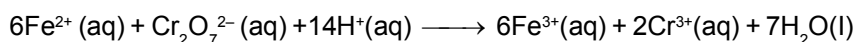
$$5. \quad E_{\text{cell}}^{\ominus} = E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^{\ominus} - E_{\text{Fe}(\text{CN})_6^{3-}/\text{Fe}(\text{CN})_6^{4-}}^{\ominus}$$

$$= +0.77 \text{ V} - 0.36 \text{ V} = +0.41 \text{ V}$$

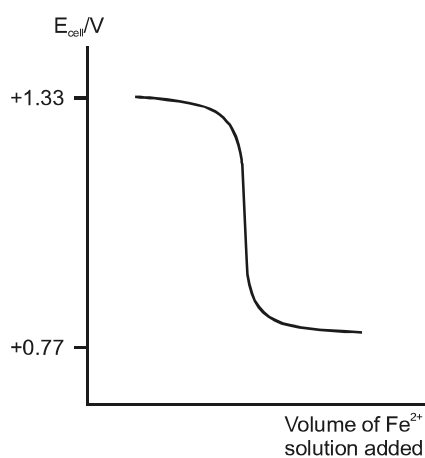
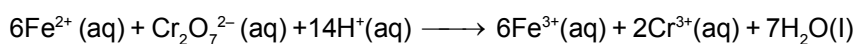
The reaction is



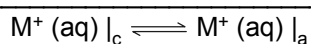
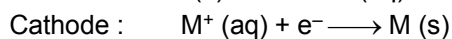
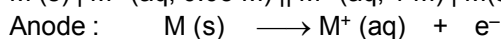
6. After one drop of iron(II) solution is added the beaker will contain a mixture of  $\text{Cr}_2\text{O}_7^{2-}$ ,  $\text{Cr}^{3+}$  and  $\text{Fe}^{3+}$  ions. The e.m.f. will be near to  $E_{\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}}^{\ominus} = +1.33 \text{ V}$ . At the end of the titration there will be  $\text{Cr}^{3+}$ ,  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$  ions. The e.m.f will be near to  $E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^{\ominus} = +0.77$ . The reaction is :



The apparatus would be like that in figure. The graph is shown in figure.



7.  $\text{M}(\text{s}) | \text{M}^+(\text{aq}, 0.05 \text{ M}) || \text{M}^+(\text{aq}, 1 \text{ M}) | \text{M}(\text{s})$



$$E_{\text{cell}} = E_{\text{cell}}^{\ominus} - \frac{0.0591}{1} \log \frac{\text{M}^+(\text{aq}) |_{\text{a}}}{\text{M}^+(\text{aq}) |_{\text{c}}}$$

$$= 0 - \frac{0.0591}{1} \log \left\{ \frac{0.05}{1} \right\}$$

$$= +ve = 70 \text{ mV and hence } \Delta G = -nFE_{\text{cell}} = -ve.$$

8. 
$$E_{\text{cell}} = \frac{-0.0591}{1} \log \left\{ \frac{0.0025}{1} \right\} = -\frac{0.0591}{1} \log \left\{ \frac{0.05}{20} \right\}$$

$$= 70 \text{ mV} + \frac{0.0591}{1} \log 20 = 140 \text{ mV}.$$



## CHEMISTRY MODULE-6 – MATRIX-MATCH TYPE QUESTIONS

### SOLUTIONS

#### Electrochemistry & Solubility of Salts

1. (A)  $\text{Al(OH)}_3 + \text{aq} \rightleftharpoons \text{Al}^{3+} \text{aq} + 3\text{OH}^- \text{aq}$   

$$\begin{matrix} & & S & & 3S \\ & & \text{Al(OH)}_3 & & \text{Al}^{3+} & & \text{OH}^- \end{matrix}$$

$$K_{sp} = (s)(3s)^3 = 27 \times s^4$$

$$27 \times s^4 = 2.7 \times 10^{-31}$$

$$s_4 = 10^{-32} \Rightarrow s = 10^{-8} \text{ mole/lit}$$
- (B) Total  $[\text{OH}^-]$  governed by NaOH only.  
 $\therefore [\text{OH}^-] = 0.1$   
 $[\text{Al}^{3+}] [\text{OH}^-]^3 = 2.7 \times 10^{-31}$   
 $[\text{Al}^{3+}] = \text{solubility} = 2.7 \times 10^{-28}$
- (C)  $\text{pH} = 5 \Rightarrow \text{pOH} = 9 \Rightarrow [\text{OH}^-] = 10^{-9}$   
 $[\text{Al}^{3+}] [\text{OH}^-]^3 = 2.7 \times 10^{-31}$   
 $[\text{Al}^{3+}] = 2.7 \times 10^{-28}$
- (D) Simultaneous solubility  
 $\text{Al(OH)}_3(\text{s}) \rightleftharpoons \text{Al}^{3+} + 3\text{OH}^-$   

$$\begin{matrix} & & S_1 & + & 3S_1 & + & 2S_2 \\ & & \text{Al(OH)}_3 & & \text{Al}^{3+} & & \text{OH}^- \end{matrix}$$

$$[\text{OH}^-] = 3s_1 + 3s_2 \sim 2s_2$$

$$K_{sp}(\text{Mg(OH)}_2) = (S_2)(2s_2)^2 = 4s_2^3 = 4 \times 10^{-12}$$

$$s_2 = 10^{-4} \text{ mole/lit}$$

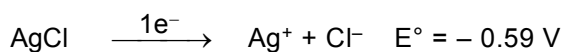
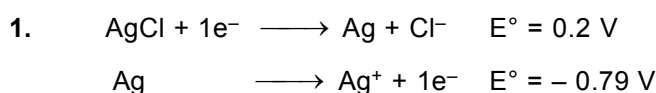
$$[\text{Al}^{3+}] [\text{OH}^-]^3 = 2.7 \times 10^{-31}$$

$$[\text{Al}^{3+}] = 2.7 \times 10^{-19}$$
2. (A) For AgBr,  $K_{sp} = [\text{Ag}^+] [\text{Br}^-]$   
 so on increase in  $[\text{Br}^-]$ , solubility of AgBr decreases.
- (B) For AgCN,  $K_{sp} = [\text{Ag}^+] [\text{CN}^-]$   
 $\text{CN}^-$  undergoes hydrolysis so solubility in water is more than expectation.  
 In acidic solution,  $\text{CN}^-$  gets protonated so solubility of AgCN increases.  
 AgCN forms complex  $\text{Ag(CN)}_2^-$  with excess of  $\text{CN}^-$ .
- (C) For  $\text{Fe(OH)}_3$ ,  $K_{sp} = [\text{Fe}^{3+}] [\text{OH}^-]^3$ .  
 In acidic solution  $\text{OH}^-$  decreases so solubility increases.  
 Solubility decreases in presence of  $\text{OH}^-$ .
- (D) For  $\text{Zn(OH)}_2$ ,  $K_{sp} = [\text{Zn}^{2+}] [\text{OH}^-]^2$   
 $\text{Zn(OH)}_2$  is amphoteric in nature so its solubility increases in both acidic as well as strongly basic solution.
3. A- Q,R; B - S; C - P; D-P
4. 1<sup>st</sup> fact.

## CHEMISTRY - MODULE - 6

### ELECTROCHEMISTRY & SOLUBILITY OF SALTS

#### Integer Type Solution



$$E^\circ = \frac{0.059}{n} \log K \quad \Rightarrow \quad -0.59 = \frac{0.059}{1} \log K_{\text{SP}}$$

$$\Rightarrow K_{\text{SP}} = 10^{-10}$$

Now solubility of AgCl in 0.1 M AgNO<sub>3</sub>

$$S(S + 0.1) = 10^{-10} \quad \Rightarrow \quad S = 10^{-9} \text{ mol/L}$$

Hence 1 mole dissolves in 10<sup>9</sup> L solution

hence in 10<sup>6</sup> L amount that dissolves ins 1 m mol.

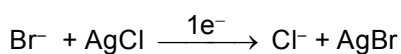
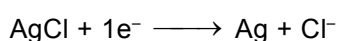
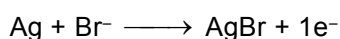
2.  $\lambda_m = \frac{1000 \times 4.95 \times 10^{-5}}{0.00099} = 50 \text{ S cm}^2 \text{ mol}^{-1}$

$$\alpha = \frac{50}{400} = 0.125$$

3.  $E_{\text{Br}^-/\text{AgBr}/\text{Ag}}^0 = E_{\text{Ag}^+/\text{Ag}}^0 + \frac{0.059}{1} \log K_{\text{SP}} \text{ AgBr} = E_{\text{Ag}^+/\text{Ag}}^0 - 0.7257$

and  $E_{\text{Cl}^-/\text{AgCl}/\text{Ag}}^0 = E_{\text{Ag}^+/\text{Ag}}^0 + \frac{0.059}{1} \log K_{\text{SP}} \text{ AgCl} = E_{\text{Ag}^+/\text{Ag}}^0 - 0.59$

Now cell reaction is



$$0 = (0.7257 - 0.59) + \frac{0.059}{1} \log \frac{[\text{Br}^-]}{[\text{Cl}^-]}$$

$$\Rightarrow \frac{[\text{Br}^-]}{[\text{Cl}^-]} = 0.005$$

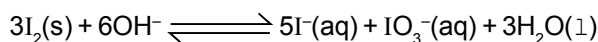
4.  $\alpha = \frac{\lambda_m}{\lambda_m^\infty} = \frac{7.8}{390} = 0.02$

$$\Rightarrow K_a(\text{CH}_3\text{COOH}) = C \alpha^2 = 0.04 \times (4.8)$$

$$\Rightarrow \text{p}K_a - \log 2^4 = 6 - 4 \times 0.3 = 4.8$$

$$\Rightarrow \text{p}K_b (= \text{H}_3\text{COO}^-) = 14 - \text{p}K_a = 9.2$$

5. Balanced equation will be



$$\Delta G^\circ = -172.5 \text{ kJmole}^{-1}$$

$$= -\frac{25}{3} \times 300 \times 2.3 \times 10^{-3} \log k$$

$$\log k = 30$$

$$10^{30} = \frac{10^{-5} \times 10^{-1}}{[\text{OH}^-]^6}$$

$$\text{so } [\text{OH}^-] = 10^{-6}$$

6. (q) Dissolved  $[\text{Zn}(\text{OH})_2] = [\text{Zn}^{+2}]_{\text{aq}} + [\text{Zn}(\text{OH})^+]_{\text{aq}} + (\text{Zn}(\text{OH})_2)_{\text{aq}} + [\text{Zn}(\text{OH})_3^-] + [\text{Zn}(\text{OH})_4^{2-}]$

$$\text{Now, } [\text{Zn}(\text{OH})_2]_{\text{aq}} = 10^{-6} \text{ M in saturated solution. so, } [\text{Zn}(\text{OH})^+] = \frac{10^{-6} \times 10^{-7}}{[\text{OH}^-]} = \frac{10^{-13}}{[\text{OH}^-]}$$

$$\text{Similarly, } [\text{Zn}^{+2}] = \frac{10^{-17}}{[\text{OH}^-]^2}, [\text{Zn}(\text{OH})_3^-] = 10^{-3} [\text{OH}^-],$$

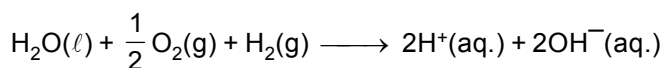
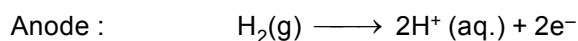
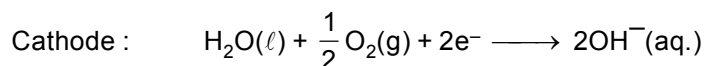
$$[\text{Zn}(\text{OH})_4^{2-}] = K_5 [\text{Zn}(\text{OH})_3^-] [\text{OH}^-] = (10^{-2} \text{ M}^{-1}) [\text{OH}^-]^2$$

$$\text{Dissolved Zn}(\text{OH})_2 = \frac{10^{-17}}{[\text{OH}^-]^2} + \frac{10^{-13}}{[\text{OH}^-]} + 10^{-6} + 10^{-3} [\text{OH}^-] + 10^{-2} [\text{OH}^-]^2$$

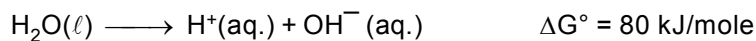
$$= \frac{10^{-17}}{10^{-16}} + \frac{10^{-13}}{10^{-8}} + 10^{-6} + 10^{-3} \times 10^{-8} + 10^{-18} = 10^{-1} + 10^{-5} + 10^{-6} + 10^{-11} = 10^{-1}$$

$$-\log \text{Zn}(\text{OH})_2(\text{aq}) = 1$$

7. Cell reaction



Also we have



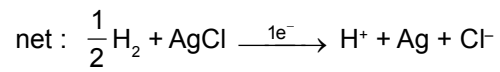
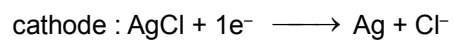
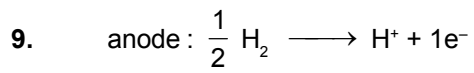
Hence for cell reaction

$$\Delta G^\circ = -77.20 \text{ kJ/mole}$$

$$\text{So, } E^\circ = -\frac{\Delta G^\circ}{nF} = \frac{77200}{2 \times 96500} = 0.40 \text{ V}$$

$$8. \quad R = \frac{1}{\kappa} \frac{\ell}{A} \quad \Rightarrow \quad \kappa = \frac{1}{250} \times \frac{7}{7} = 0.004 \Omega^{-1} \text{ cm}^{-1}$$

$$\Lambda_{\text{eq}} = \frac{0.004}{0.1} \times 1000 \Omega^{-1} \text{ cm}^2 \text{ eq}^{-1} = 40 \Omega^{-1} \text{ cm}^2 \text{ eq}^{-1}$$



$$E_{\text{cell}} = + 0.222 + \frac{0.059}{1} \log \frac{1}{[\text{H}^+][\text{Cl}^-]}$$

$$= + 0.222 + 0.059 \log \frac{[\text{OH}^-]}{[10^{-14}][\text{Cl}^-]}$$

$$= + 0.222 + 0.059 (14)$$

$$= + 1.048 \text{ volt}$$

10. (3)