CHAPTER 17

Electrochemistry

1. **(a)** Oxidising tendency $\propto \dfrac{1}{\text{Electrode potential}}$

$\text{TX} \longrightarrow \text{No reaction}$

$\text{TY} \longrightarrow \text{X, Z}$

$\text{TZ} \longrightarrow \text{X}$

$\Rightarrow$ order of electrode potential is

$\text{TY} < \text{TZ} < \text{TX}$

$\Rightarrow$ Order of oxidation of the anion is

$\text{Y}^- > \text{Z}^- > \text{X}^-$

2. **(a)** $2\text{Ag}^+ + \text{H}_2 \longrightarrow 2\text{H}^+ + 2\text{Ag}$

$$E = E° - \frac{0.0591}{2} \log \frac{[\text{H}^+]^2}{\text{P}_{\text{H}_2} \times [\text{Ag}^+]^2}$$

$$0.222 = 0.7995 - \frac{0.0591}{2} \log \frac{1}{[\text{Ag}^+]^2}$$

$[\text{Ag}^+] = 10^{-9.8}$

$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = (10^{-9.8}) \times (1) = 10^{-9.8}$

3. **(b)** $\text{Sn}^{4+} + 2\text{e}^- \longrightarrow \text{Sn}^{2+};\ E° = 0.13\text{V}$

$\text{Br}_2 + 2\text{e}^- \longrightarrow 2\text{Br}^-;\ E° = 1.08\text{V}$

$E°$ value shows Br_2 has higher reduction potential.

Hence

$E_{cell} = E_{\text{Br}_2/\text{Br}^-} - E_{\text{Sn}^{+4}/\text{Sn}^{+2}}$

$= 1.08 - 0.13 = 0.95\ \text{V}$

Now $-\Delta G = nF\,E_{\text{cell}}$

$n = 2, F = 96500$

$-\Delta G = 2 \times 96500 \times 0.95\ \text{kJ/mol.}$

Also, $\Delta G = -2.303\,\text{RT} \log K_{\text{eq}}$

$$\log K_{\text{eq}} = -\frac{\Delta G}{2.303 \times R \times T}$$

$$= \frac{-(-2 \times 96500 \times 0.95)}{2.303 \times 8.314 \times 298} = 32.13$$

$K_{\text{eq}} = \text{antilog } 32.682 \approx 10^{32}$

4. **(a)** When $\text{pH} = 14\ [\text{H}^+] = 10^{-14}$ and $[\text{OH}^-] = 1\text{M}$

$K_{sp} = [\text{Cu}^{2+}][\text{OH}^-]^2 = 10^{-19}$

$$\therefore\quad [\text{Cu}^{2+}] = \frac{10^{-19}}{[\text{OH}^-]^2} = 10^{-19}$$

The half cell reaction

$\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}$

$$E = E° - \frac{0.059}{2} \log \frac{1}{[\text{Cu}^{2+}]}$$

$$= 0.34 - \frac{0.059}{2} \log \frac{1}{10^{-19}} = -0.22\ \text{V}$$

5. **(c)** The reaction

$2\,\text{Cu}^+(aq) \longrightarrow \text{Cu}(s) + \text{Cu}^{2+}(aq)$

$$E_{\text{cell}} = E°_{\text{cell}} - \frac{0.059}{1} \log \frac{[\text{Cu}^{2+}]}{[\text{Cu}^+]^2}$$

At equilibrium $E_{\text{cell}} = 0$

$\therefore\quad E°_{\text{cell}} = 0.059 \log K_{\text{eq}}$

$$\text{or,}\quad \log K_{\text{eq}} = \frac{0.52 - 0.16}{0.059}$$

$\therefore\quad K_{\text{eq}} = 1.2 \times 10^6$

6. **(b)** $\text{Cu}^{2+} + 1\text{e}^- \longrightarrow \text{Cu}^+$

$E_1^{\text{o}} = 0.15\text{V};\ \Delta G_1^{\text{o}} = -n_1 E_1^{\text{o}} F$

$\text{Cu}^+ + 1\text{e}^- \longrightarrow \text{Cu}$

$E_2^{\text{o}} = 0.50\text{V};\ \Delta G_2^{\text{o}} = -n_2 E_2^{\text{o}} F$

$\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu} \qquad \Delta\text{G}° = \Delta\text{G}°_1 + \Delta\text{G}°_2$

$$-nE^\circ F = -1\, n_1 E_1^\circ F + (-1)\, n_2 E_2^\circ F$$
$$-nE^\circ F = -1\, (n_1 E_1^\circ F + n_2 E_2^\circ F)$$

$$E^\circ = \frac{n_1E_1^o + n_2E_2^o}{n} = \frac{0.15\times1+0.50\times1}{2}$$

$\Rightarrow 0.325$

7. **(d)** The cell reactions are :

At anode : $\mathrm{Fe}(s) \longrightarrow \mathrm{Fe}^{2+}(aq) + 2e^-$

At cathode : $\mathrm{Cu}^{2+}(aq) + 2e^- \longrightarrow \mathrm{Cu}\ (s)$

We know that :

$$\Delta G^\circ = -nF\ E^\circ_{cell}\ ;\ n = 2\ \text{mol}$$

$$E^\circ_{cell} = [E^\circ_{(Cu^{2+}/Cu)} - E^\circ_{(Fe^{2+}/Fe)}]$$
$$= (+0.34\ \text{V}) - (-0.44\ \text{V}) = +0.78\ \text{V}$$

$F = 96500\ \text{C mol}^{-1}$

$$\therefore\ \Delta G^\circ = -nFE^\circ_{cell}$$
$$= -(2\ \text{mol}) \times (96500\ \text{C mol}^{-1}) \times (+0.78\ \text{V})$$
$$= -150540\ \text{CV} = -150540\ \text{J}\ \ (\because\ 1\ \text{CV} = 1\ \text{J})$$
$$= -150.54\ \text{kJ}$$

8. **(b)** For this cell, reaction is;

$\mathrm{Zn + Fe^{2+} \rightarrow Zn^{2+} + Fe}$

$$E = E^\circ - \frac{0.0591}{n}\log\frac{c_1}{c_2};\ E^\circ = E + \frac{0.0591}{n}\log\frac{c_1}{c_2}$$

$$E^\circ = 0.2905 + \frac{0.0591}{2}\log\frac{10^{-2}}{10^{-3}} = 0.32\ \text{V}.$$

$$E^\circ = \frac{0.0591}{2}\log K_{eq}$$

$$\log K_{eq} = \frac{0.32\times2}{0.0591} = \frac{0.32}{0.0295}$$

$$\therefore K_{eq} = 10^{\frac{0.32}{0.0295}}.$$

9. **(a)** We know

$$2H_2(g) \longrightarrow 4H^+ + 4e^-;\ E = 0.0\ \text{V}$$

$$\underline{O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l);\ E = 1.23\ \text{V}}$$
$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l);\ E^\circ_{cell} = 1.23\ \text{V}$$

$$\Delta G^\circ_{298} = -nFE^\circ = -4\times96500\times1.23 = -474.78\ \text{kJ}$$

$$\Delta_f H^\circ_{298} = \Delta_r G^\circ_{298} + T.\Delta_r S^\circ_{298}$$

$$= (-474.78) + 298 \times (-0.32)$$
$$= -570.14\ \text{kJ/mol}$$

$$\therefore\ \ \Delta_f H^\circ_{298} = -285.07\ \text{kJ/mol}$$

10. **(d)** Molarity $= 0.01$ M ; Resistance $= 40$ ohm;

Cell constant $\frac{l}{A} = 0.4\,\text{cm}^{-1}$.

Specific conductivity (κ)

$$= \frac{\text{cell constant}}{\text{resistance}} = \frac{0.4}{40} = 0.01\ \text{ohm}^{-1}\ \text{cm}^{-1}$$

$$\text{Molar conductance } (\wedge_m) = \frac{1000\kappa}{\text{Molarity}}$$

$$= \frac{1000\times0.01}{0.01} = 10^3\ \text{ohm}^{-1}\ \text{cm}^2\ \text{mol}^{-1}$$

11. **(a)** $2H^+ + 2e^- \longrightarrow H_2$

$$E_H\,(\text{Eq. wt}) = \frac{2}{2} = 1\ \text{g}$$

$$= \frac{22400}{2} = 11200\ \text{mL (STP)}$$

$$\text{Total charge passed} = \frac{96500\times112}{11200} = 965$$

$Q = It = 965$

$$I = \frac{965}{965} = 1\,\text{amp.}$$

12. **(c)** $\mathrm{ClCH_2COONa + HCl \rightarrow}$

$\mathrm{ClCH_2COOH + NaCl}$

$$\Lambda_{ClCH_2COONa} + \Lambda_{HCl} = \Lambda_{ClCH_2COOH} + \Lambda_{NaCl}$$

$$224 + 203 = \Lambda_{ClCH_2COOH} + 38.2$$

$$\Lambda_{ClCH_2COOH} = 427 - 38.2$$

$$= 388.8\ \text{ohm}^{-1}\text{cm}^2\text{g eq}^{-1}$$

13. **(a)** $\Lambda^\infty_{eq}(NH_4OH) = \Lambda^\infty_{eq}(NH_4Cl) +$

$\Lambda^\infty_{eq}(NaOH) - \Lambda^\infty_{eq}(NaCl)$

$$= 129.8 + 217.8 - 109.3 = 238.3\ \text{ohm}^{-1}\text{cm}^2\text{eq}^{-1}$$

$$\alpha = \frac{\Lambda_{eq}}{\Lambda^\infty_{aq}} = \frac{9.30}{238.3} = 0.04$$

14. **(b)** $$\underset{c(1-\alpha)}{HA(aq)} \rightleftharpoons \underset{c\alpha}{H^+(aq)} + \underset{c\alpha}{A^-(aq)}$$

$$K_a = \frac{c\alpha^2}{1-\alpha}; \text{ where } \alpha = \frac{\Lambda_m}{\Lambda_m^\infty}$$

$$\therefore \quad K_a = \frac{c\left(\frac{\Lambda_m}{\Lambda_m^\infty}\right)^2}{\left(1-\frac{\Lambda_m}{\Lambda_m^\infty}\right)}$$

$$= \frac{c\Lambda_m^2}{\Lambda_m^\infty\left(\Lambda_m^\infty - \Lambda_m\right)}$$

15. **(a)** (Eq. = Charge on the ion/Total charge)

$$[K^+] = \frac{1}{8} \text{ mole} \times 2 = \frac{1}{4} \text{ mole} = \frac{1}{4} \text{ Eq.}$$

$$[Al^{3+}] = \frac{6}{8} = \frac{3}{4} \text{ Eq.}$$

$$[SO_4^{2-}] = \frac{8}{8} = 1 \text{ Eq.}$$

Λ_{eq}° $K_2SO_4.Al_2(SO_4)_3.24H_2O$

$= \lambda_{eq}^\circ(K^+) + \lambda_{eq}^\circ(Al^{3+}) + \lambda_{eq}^\circ(SO_4^{2-})$

$$= \frac{1}{4} \times 73.5 + 149 \times \frac{3}{4} + 85.8 \times 1$$

$= 18.375 + 111.75 + 85.8 = 215.92$

16. **(a)** By Faraday's Ist law of electrolysis,

$$\frac{W}{E} = \frac{Q}{96500} \qquad (\text{where } Q = it = \text{charge of ion})$$

We know that no. of gram equivalent

$$= \frac{W}{E} = \frac{it}{96500} = \frac{1 \times 965}{96500} = \frac{1}{100}$$

(where i = 1 A, t = 16 × 60 + 5 = 965 sec.)

Since, we know that

$$\text{Normality} = \frac{\text{No. of gram equivalent}}{\text{Volume (in litre)}} = \frac{\frac{1}{100}}{1}$$

=0.01 N

17. **(b)** The molar conductivity of the dissociated form of crotonic acid is

$\Lambda_m(HC) = \Lambda_m(HCl) + \Lambda_m(NaC) - \Lambda_m(NaCl)$

$= (426 + 83 - 126)\ \Omega^{-1}\ cm^2 mol^{-1}$

$= 383\ \Omega^{-1}\ cm^2 mol^{-1}$

The molar conductivity of HC,

$$\Lambda_m(HC) = \frac{\kappa}{C} = \frac{3.83 \times 10^{-5} \Omega^{-1} cm^{-1}}{0.001} \times 1000$$

$= 38.3\ \Omega^{-1}\ cm^2 mol^{-1}$

The degree of dissociation,

$$\alpha = \frac{\Lambda_m(HC)}{\Lambda_m^\infty(HC)} = \frac{\left(38.3\Omega^{-1}cm^2mol^{-1}\right)}{\left(383\Omega^{-1}cm^2mol^{-1}\right)} = 0.1$$

$$K_a = \frac{C\alpha^2}{1-\alpha} = \frac{\left(10^{-3}\right)(0.1)^2}{1-0.1} = 1.11 \times 10^{-5}$$

18. **(d)**
$$\begin{array}{l} Pb + SO_4^{--} \longrightarrow PbSO_4 + 2e^- \\ PbO_2 + 4H^+ + SO_4^{--} + 2e^- \longrightarrow PbSO_4 + 2H_2O \\ \hline Pb + PbO_2 + 2H_2SO_4 \longrightarrow 2PbSO_4 + 2H_2O \end{array}$$

The reaction indicates that 2 moles of H_2SO_4 corresponds to 2 × 96500 C and 2 mol H_2SO_4 $\equiv$ 4 equiv. of H_2SO_4.

2 × 96500 C consumed 4 equiv. of H_2SO_4

and $100 \times 10^{-3} \times 9.65 \times 10^5$ C consumed

$$= \frac{4 \times 100 \times 10^{-3} \times 9.65 \times 10^5}{2 \times 96500} = 2 \text{ equiv. } H_2SO_4$$

$$\therefore \quad \text{Decrease in normality} = \frac{2}{5} = 0.40$$

19. **(c)** Gold deposited in the first cell = 9.85 g

At. wt. of Gold = 197, Oxidation number of gold = +3

$$\text{Eq. Wt. of Gold} = \frac{197}{3}$$

W = Zit

(where W stands for the weight of ions deposited, i for current and t for time and Z for electro-chemical equivalent of the electrolyte.)

∵ Charge required to deposit 1 g eq. of gold = 1F = 96,500 C

∴ Charge required to deposit 9.85 g of gold or

$$\frac{9.85}{197/3} \text{ g eq. of gold} = \frac{96{,}500 \times 9.85 \times 3}{197} \text{ C}$$

= 14475 C

According to Faraday's second law,

$$\frac{\text{Wt. of Cu}}{\text{Eq. wt. of Cu}} = \frac{\text{Wt. of Gold}}{\text{Eq. wt. of Gold}}$$

$\Rightarrow$ Wt. of Cu deposited

$$= \frac{9.85 \times 3}{197} \times \frac{63.5}{2} = 4.76\,\text{g}$$

$$\text{Current} = \frac{Q}{t} = \frac{14475}{5 \times 3600}\,\text{A} = \frac{193}{240}\,\text{A} = 0.80\,\text{A}$$

20. **(c)** $Cu^{2+}(aq) + H_2 \rightleftharpoons 2H^+ + Cu(s)$

For the given electrochemical cell

$Pt \mid H_2\,(p = 1\text{ atm}) \mid$

$H^+\,(aq.\ x\text{M}) \parallel Cu^{2+}\,(aq.\ 1.0\text{ M}) \mid Cu(s)$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{n}\log\frac{[H^+]^2}{[Cu^{2+}]}$$

$$0.49 = 0.37 - \frac{0.0591}{2}\log\frac{x^2}{1}$$

$$0.49 = 0.37 - \frac{0.0591}{2} \times 2\log x$$

$0.15 = -0.0591 \times \log x$

$2.54 = -\log x$

Also, $-\log[H^+] = \text{pH}$

$\therefore\ \ \text{pH} = -\log x$

$\Rightarrow\ \ \text{pH} = 2.54 \approx 2.5$

21. **(1.39)** $Cr^{2+}(aq) \longrightarrow Cr^{3+}(aq) + e^-, E^\circ = 0.41\text{ V}$ ...(i)

$HClO(aq) + H^+(aq) + 2Cr^{2+}(aq) \longrightarrow$

$2Cr^{3+}(aq) + Cl^-(aq) + H_2O(l), E^\circ = 1.8\text{V}$ ...(ii)

[Reaction (i)] × – 2 + [Reaction (ii)] we get.

$$HClO(aq) + H^+(aq) + 2e^- \longrightarrow Cl^-(aq) + H_2O(l)$$

$$\Delta G^\circ = -2FE^\circ - 2(-1 \times FE^\circ)$$

$$-2FE^\circ = -2F(1.8) + 2F(0.41)$$

$E^\circ = +1.8 - 0.41 = 1.39\text{ V}$

22. **(0.84)** After addition of Cd and its oxidation into Cd^{2+}.

$$NO_3^-(aq) + 4H^+(aq) + 3e^- \longrightarrow NO(g) + 2H_2O(l)$$

$0.1 - x \quad 0.4 - 4x;$

where $x = 0.08$

$\left[NO_3^-\right]$remaining $= 0.02$ M; $\left[H^+\right]$remaining $= 0.08$M

$$E_{NO_3^-|NO} - E^\circ_{NO_3^-|NO} - \frac{0.0591}{3}\log\frac{1}{\left[NO_3^{0-}\right]\left[H^+\right]^4}$$

$$= 0.96 - \frac{0.0591}{3}\log\frac{1}{(0.02)(0.08)^4} = 0.84\text{ V}$$

23. **(0.0529)** Using Faraday's second law of electrolysis,

$$\frac{\text{Weight of Cu deposited}}{\text{Weight of Ag deposited}} = \frac{\text{Equ. wt. of Cu}}{\text{Equ. wt. of Ag}}$$

$$\Rightarrow\ \frac{w_{Cu}}{0.18} = \frac{63.5}{2} \times \frac{1}{108}$$

$$\Rightarrow\ w_{Cu} = \frac{63.5 \times 18}{2 \times 108 \times 100} = 0.0529\text{ g.}$$

24. **(1.5)**

For the cell:

(anode) (cathode)

$Cr \mid Cr^{3+}\,(1\text{ M}) \parallel H^+\,(1\text{ M}) \mid Pt\,(H_2)\,(1\text{ bar})$

$$E^o_{\text{cell}} = E^o_{H^+/H_2} - E^o_{Cr^{3+}/Cr}$$

$$0.74 = 0 - E^o_{Cr^{3+}/Cr}$$

$$E^o_{Cr^{3+}/Cr} = -0.74\text{ V}$$

For the cell:

(anode) (cathode)

$(1\text{ bar})\ Pt\,(H_2) \mid H^+\,(1\text{ M}) \parallel Ag^+\,(1\text{ M}) \mid Ag$

$$E^o_{\text{cell}} = E^o_{Ag^+/Ag} - E^o_{H^+/H_2}$$

$$0.80 = E^o_{Ag^+/Ag} - 0$$

$$E^o_{Ag^+/Ag} = +0.80\text{ V}$$

For the cell:

(anode) (cathode)

$Ag \mid Ag^+\,(0.1\text{ M}) \parallel Cr^{3+}\,(0.1\text{ M}) \mid Cr$

Reduction at cathode:

(i) $Cr^{3+} + 3e^- \rightarrow Cr;\ E^o_{Cr^{3+}/Cr} = -0.74\text{ V}$

Oxidation at anode:

(ii) $Ag \rightarrow Ag^+ + e^-;\ E^o_{Ag^+/Ag} = +0.80\text{ V}$

eq. (i) + 3 × eq. (ii)

$Cr^{3+} + 3\,Ag \rightarrow Cr + 3\,Ag^{+}; n = 3$

$$E^{o}_{cell} = E^{o}_{Cr^{3+}/Cr} - E^{o}_{Ag^{+}/Ag}$$

$$E^{o}_{cell} = -0.74 - 0.80 = -1.54\,V$$

$$E_{cell} = E^{o}_{cell} - \frac{0.0591}{n}\log\frac{[Ag^{+}]^{3}}{[Cr^{3+}]}$$

$$= -1.54 - \frac{0.06}{3}\log\frac{(0.1)^{3}}{(0.1)} = -1.54 - \frac{0.06}{3} \times (-2)$$

$= -1.5\,V$

25. **(0.15)**

Cell reaction for Daniell's cell:

$Zn\,(s) + Cu^{2+}\,(aq) \rightarrow Cu\,(s) + Zn^{2+}\,(aq)$

$$\text{Reaction quotient } Q = \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

At start, $Q = 1$ as $[Zn^{2+}] = [Cu^{2+}]$

also $\log Q = 0$ at point O for $Q = 1$

This is the maximum EMF, the cell can generate. As the reaction proceeds, $[Zn^{2+}]$ will increase and $[Cu^{2+}]$ will decrease. Thus, the value of Q will increase and the cell EMF will decrease.

$$E_{cell} = E^{o}_{cell} - \frac{0.0591}{n}\log\frac{[Zn^{2+}]}{[Cu^{2+}]}$$

$E_{cell} = 1.07\,V,\ E^{o}_{cell} = 1.10\,V, n = 2$

$$1.07 = 1.10 - \frac{0.06}{2}\log\frac{1.5}{[Cu^{2+}]}$$

$$0.03 = 0.03\log\frac{1.5}{[Cu^{2+}]}$$

$$\log\frac{1.5}{[Cu^{2+}]} = 1$$

$$\Rightarrow \quad \frac{1.5}{[Cu^{2+}]} = 10$$

$\Rightarrow \quad [Cu^{2+}] = 0.15\,mol/L$

26. **(76.43)**

Reduction at cathode:

$$\overset{o}{O_2} \rightarrow 2H_2\overset{-2}{O}$$

$O_2 + 4\,H^{+} + 4e^{-} \rightarrow 2\,H_2O,\ \ E^{o} = +1.23\,V$

Oxidation at anode:

$$\overset{-2}{N_2}H_4 \rightarrow \overset{o}{N_2}$$

$N_2H_4 \rightarrow N_2 + 4\,H^{+} + 4e^{-}\ \overset{o}{E},\ = -0.9\,V$

Overall cell reaction:

$N_2H_4\,(aq) + O_2(g) \rightarrow N_2(g) + 2H_2O\,(l); n = 4$

$$E^{o}_{cell} = E^{o}_{\frac{1}{2}O_2/H_2O} - E^{o}_{N_2/N_2H_4}$$

$$= (1.23 - 0.9)V = +0.33\,V$$

Maximum available potential $E^{o}_{cell} = 0.33\,V$

Maximum amount of energy $= -\Delta G$

$= nFE^{o} = 4 \times 96500 \times 0.33\,J = 1{,}27{,}380\,J$

Assuming 60% efficiency

$$W_{max} = 127380 \times \frac{60}{100}\,J = 76.43\,KJ$$

27. **(6)** Molecular mass of $H_2SO_4 = 98$ g/mol

Molarity of initial solution

$$= \frac{\%W/W \times 10 \times d}{\text{moleculer mass}} = \frac{38 \times 10 \times 1.274}{98} = 4.94$$

Molarity of final solution

$$= \frac{19.6 \times 10 \times 1.2}{98} = 2.4\,M$$

Decrease in amount of H_2SO_4 during this discharging process

$= (4.94 - 2.4) \times 98 \times 2.36\,g = 6 \times 98\,g$

Overall reaction during discharging

$Pb(s) + PbO_2(s) + 2H_2SO_4(aq.) \rightarrow$

$2PbSO_4(s) + 2H_2O(l); n = 2$

$\therefore$ For 1 mole of H_2SO_4 +1 F is required

6×98 g $H_2SO_4 = 6$ moles of H_2SO_4

Thus, number of Faradays required = 6

28. **(142)**

Sodium metal :

$E = E_0 + (KE)_{max};\ E^{0}_{cell} = 0.22\,V$

Cell reaction

Cathode : $AgCl(s) + e^{-} \longrightarrow Ag(s) + Cl^{-}(aq)$

Anode : $\frac{1}{2}H_2(g) \longrightarrow H^+(aq) + e^-$

Overall :

$$AgCl(s) + \frac{1}{2}H_2(g) \longrightarrow Ag(s) + H^+(aq) + Cl^-(aq)$$

$$E_{cell} = E^0_{cell} - \frac{0.06}{1}\log[H^+][Cl^-]$$

$$E_{cell} = 0.22 - \frac{0.06}{1}\log[10^{-1}][10^{-1}]$$

$$= 0.22 + 0.12 = 0.34\text{ V}$$

$(KE)_{max} = E_{cell} = 0.34$ eV

So, E = 2.3 + 0.34 = 2.64 eV = Energy of photon incident

For potassium metal :

$E = E_0 + (KE)_{max}$

$2.64 = 2.25 + (KE)_{max}$

$(KE)_{max} = 0.39 = E_{cell}$

Cell reaction

Cathode : $AgCl(s) + e^- \longrightarrow Ag(s) + Cl^-(aq)$

Anode : $\frac{1}{2}H_2(g) \longrightarrow H^+(aq) + e^-$

Overall : $AgCl(s) + \frac{1}{2}H_2(g) \longrightarrow$

$$Ag(s) + H^+(aq) + Cl^-(aq)$$

$$E_{cell} = E^0_{cell} - \frac{0.06}{1}\log[H^+][Cl^-]$$

$0.39 = 0.22 - 0.06\log[H^+]^2$

$0.39 = 0.22 - 0.12\log[H^+]$

$0.17 = 0.12 \times pH$

$pH = 17/12 = 1.4166 \simeq 1.42 = 142 \times 10^{-2}$

29. **(2.15)**

At equilibrium state $E_{cell} = 0$; $E^0_{cell} = 0.01$ V

$Sn + Pb^{2+} \longrightarrow Sn^{2+} + Pb$

$$E = E^0_{cell} - \frac{0.06}{n}\log\frac{[P]}{[R]}$$

$$0 = 0.01 - \frac{0.06}{2}\log\frac{[Sn^{2+}]}{[Pb^{2+}]}$$

$$-0.01 = -\frac{0.06}{2}\log\frac{[Sn^{2+}]}{[Pb^{2+}]}$$

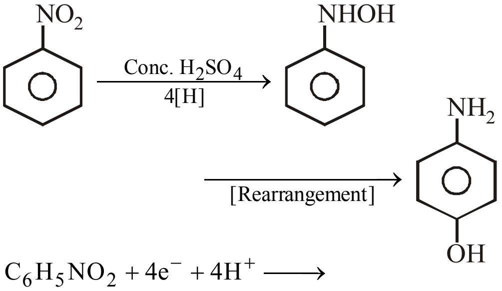
$$\frac{1}{3} = \log\frac{[Sn^{2+}]}{[Pb^{2+}]}$$

$$\frac{[Sb^{2+}]}{[Pb^{2+}]} = 10^{1/3} = 2.15$$

30. **(9.81)** 9.65 ampere current was passed for 1.0 hour (3600 seconds)

Number of moles of electrons passed

$$= \frac{I(A) \times t(s)}{96500} = \frac{9.65A \times 3600s}{96500} = 0.36\text{ moles}$$



$C_6H_5NO_2 + 4e^- + 4H^+ \longrightarrow$

p-Aminophenol + H_2O

$\because$ 4 moles of electrons reduces 1 mole of nitrobenzene to *p*-aminophenol.

$\therefore$ 0.36 moles of electrons will reduce $\frac{0.36}{4} = 0.09$ moles of nitrobenzene to *p*-aminophenol.

p-aminophenol molar mass = 109.14 g/mol

Mass of *p*-aminophenol obtained

= 109.14 g/mol × 0.09 mol = 9.81 g