Solutions

S1. Ans. (d)

$4OH^- \longrightarrow 2H_2O + O_2 + 4e^-$

for 2 mole of H_2O = 4F charge is required

for 1 mole of $H_2O = \frac{4F}{2}$ = 2F required

$\overset{+7}{MnO_4^-} \longrightarrow \overset{+2}{Mn^{2+}}$

for 1 mole MnO_4^- 5F charge is required

$Ca^{2+} \xrightarrow{+2e^-} Ca$

for 1 mole Ca^{2+} ion required = 2F

1.5 mole Ca^{2+} ion required $= \frac{2}{1} \times 1.5$ = 3F

$\overset{+2}{FeO} \longrightarrow \overset{+3}{Fe_2O_3}$

for 1 mole FeO, 1F charge is required.

S2. Ans. (c)

$2NO(g) \rightleftharpoons N_2(g) + O_2(g)$

$K_c = \frac{[N_2][O_2]}{[NO]^2}$

$= \frac{3\times10^{-3}\times4.2\times10^{-3}}{2.8\times10^{-3}\times2.8\times10^{-3}} = 1.607$

	$2NO(g)$	$\rightleftharpoons$	$N_2(g)$	+	$O_2(g)$
t = 0	0.1		0		0
	$0.1 - 0.1\,\alpha$		$0.05\,\alpha$		$0.05\,\alpha$

$K_c = \frac{0.05\alpha\times0.05\alpha}{(0.1-0.1\alpha)^2}$

$K_c = \frac{0.5\alpha\times0.05\alpha}{0.01(1-\alpha)^2}$

$1.607 = \frac{(0.05)^2\alpha^2}{0.01(1-\alpha)^2}$

$\frac{\alpha^2}{(1-\alpha)^2} = \frac{1.607\times(0.1)^2}{(0.05)^2}$

$\frac{\alpha}{1-\alpha} = \frac{1.27\times0.1}{0.05}$

$\frac{\alpha}{1-\alpha} = 2.54$

$3.54\,\alpha = 2.54$

$\alpha = \frac{2.54}{3.54} = 0.717$

S3. Ans. (b)

Tl^3 act as an oxidising agent not reducing agent.

S4. Ans. (d)

B and D statements are correct.

S5. Ans. (a)

$E^\circ_{cell} = E^\circ_C - E^\circ_A$

= (1.33) – (–0.44)

= +1.77 V

S6. Ans. (b)

Centimolar solution = 1/100 M = 0.01 M

Conductivity (k) = 0.0210 $ohm^{-1}\ cm^{-1}$

Resistance (R) = 60 ohm

$k = 1/R\left(\frac{l}{A}\right)$

$\Rightarrow 0.0210 = 1/60\left(\frac{l}{A}\right) \Rightarrow \frac{l}{A} = 1.26\ cm^{-1}$

S7. Ans. (d)

$\Delta_{rG} = -nFE_{cell}$

E_{cell} is an intensive property and Δ_{rG} is an extensive property as it depends on number of e^- transferred in cell reaction.

S8. Ans. (d)

Since, E°_{OP} of Al is more than Co^{2+}, so at anode Al will oxidise and at cathode Co^{3+} will reduce.

$E^\circ_{cell} = (E^\circ_{Cathode})_{RP} - (E^\circ_{Anode})_{RP}$

$= E^\circ_{Co^{3+}/Co^{2+}} - E^\circ_{Al^{3+}/Al}$

= (1.81) – (–1.66)

= +3.47 V

S9. Ans. (c)

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

$E^\circ_{cell} = 1.1$ V

$\Delta G^\circ = -nFE^\circ_{cell}$

$\therefore n = 2$

$\Delta G^\circ = -2\times 96487 \times 1.1$

$\Delta G^\circ = -212271.4\ J\ mol^{-1}$

$\Delta G^\circ = -212.27\ kJ\ mol^{-1}$

S10. Ans. (d)

For a reaction to be spontaneous, E°_{cell} must be positive.

■ For, $FeSO_4(aq) + Zn(s) \longrightarrow ZnSO_4(aq) + Fe(s)$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

= -0.44 V – (-0.76 V) = 0.32 V

■ For, $2CuSO_4(aq) + 2Ag(s) \longrightarrow 2Cu(s) + Ag_2SO_4(aq)$

E°_{cell} = 0.34 V – 0.80 V

= –0.46 V

■ For, $CuSO_4(aq) + Zn(s) \longrightarrow ZnSO_4(aq) + Cu(s)$

E°_{cell} = 0.34 V – (–0.76 V)

= 1.1 V

■ Foe, $CuSO_4(aq) + Fe(s) \longrightarrow FeSO_4(aq) + Cu(s)$

E°_{cell} = 0.80 V – (–0.44 V)

= 1.24 V

S11. Ans. (a)

■ $MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$ -(i)

$E^{\circ}_{MnO_4^-/Mn^{2+}} = -E^{\circ}_{Mn^{2+}/MnO_4^-} = 1.51$ V

■ $H_2O \longrightarrow \frac{1}{2}O_2 + 2H^+ + 2e^-$ -(ii)

$E^{\circ}_{O_2/H_2O} = 1.223$ V

Using 2 × (i) + 5 × (ii), net cell reactions is

$2MnO_4^- + 6H^+ \longrightarrow 2Mn^{2+} + \frac{5}{2}O_2 + 3H_2O$

$E^{\circ}_{cell} = E^{\circ}_{C} - E^{\circ}_{A} = E^{\circ}_{MnO_4^-/Mn^{2+}} - E^{\circ}_{O_2/H_2O}$

= 1.51 – 1.223 = 0.287 V

Since $E^{\circ}_{cell} > 0$, therefore net cell reaction is spontaneous and so MnO_4^- liberate O_2 from H_2O in presence of an acid.

S12. Ans. (c)

$Ni(s) + 2Ag^+(0.001\ M) \longrightarrow Ni^{2+}(0.001\ M) + 2Ag(s)$

E°_{cell} = 1.05 V

$E^{\circ}_{cell} = E^{\circ}_{cell} - \frac{0.059}{n}\log\frac{[Ni^{2+}]}{[Ag^+]^2}$

$\Rightarrow 1.05 - \frac{0.059}{2}\log\frac{(10^{-3})}{(10^{-3})^2}$

$\Rightarrow 1.05 - \frac{0.059}{2}\log(10)^3$

$\Rightarrow$ 1.05 – 0.0295 × 3

$\Rightarrow$ 1.05 – 0.0885

$\Rightarrow$ 0.9615 V

S13. Ans.(a)

$\Lambda_{NaCl} = \Lambda_{Na^+} + \Lambda_{Cl^-}$

$\Lambda_{HCl} = \Lambda_{H^+} + \Lambda_{Cl^-}$

$\Lambda_{CH_3COONa} = \Lambda_{Na^+} + \Lambda_{CH_3COO^-}$

Let, $\Lambda_{Na^+} = x, \Lambda_{Cl^-} = y, \Lambda_{H^+} = \Lambda_{CH_3COO^-=w}$

Given,

x + y = 126.45 ...(i)

y + z = 426.16 ...(ii)

x + w = 91 ...(iii)

From the above 3 equations, value of

z +w = 390.71d

S14. Ans.(b)

$\Lambda^{\circ}_{M(CH_3COOH)} = \Lambda^{\circ}_{M(H^+)} + \Lambda^{\circ}_{M(CH_3COO^-)}$

= 350 + 50 = 400 S cm^2 mol^{-1}

$\alpha = \frac{\Lambda^C_M}{\Lambda^{\circ}_M}$

$a = \frac{20}{400} = 5 \times 10^{-2}$

$K_{a(CH_3COOH)} = C\alpha^2$

$= 0.007 \times (5 \times 10^{-2})^2 = 1.75 \times 10^{-5} mol\ L^{-1}$

S15. Ans.(a)

The following reactions take place on electrolysis of dilute sulphuric acid on using pt electrodes are:

At cathode:

$4H^+ + 4e^-(aq) \rightarrow 2H_2(g)$

At anode:

$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

Thus, the product obtained at anode is O_2.

S16. Ans.(d)

1 equivalent of any substance is deposited by 1 F of charge.

We have, 20 g calcium

The balance reaction will

The charge on Ca in $CaCl_2$

Cl has –1 charge so that

Ca + 2(–1) = 0

Ca = 2

We have to get Ca from Ca^{2+}

Number of required moles = mass/molar mass

Molar mass of Ca is 40 g/mol and required of Ca is 20 g

Hence number of moles = 20/40

= 0.5 mol

Electricity required to produce 1 mol of calcium = 2F

The electricity required to produce 0.5 mol of calcium = 0.5 × 2F

= 1F

S17. Ans.(b)

$Au^{3+} + 3e^- \rightarrow Au(s)\ E^\circ = 1.40\ V$

$Fe^{2+} + 2e^- \rightarrow Fe(s)\ E^\circ = -0.44\ V$

$Mg^{2+} + 2e^- \rightarrow Mg(s)$ E°=-2.36 V

$K^+ + 1e^- \rightarrow K(s)\ E^\circ = -2.93\ V$

Au^{3+} occupies the top position in the electrochemical series.

S18. Ans.(a)

Cell reaction involved in hydrogen-oxygen fuel cell is

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

Thus, $R = H_2(g), O_2(g); P = H_2O(l)$

S19. Ans.(c)

$E_{cell} = E^\circ{}_{cell} - \frac{0.059}{n} logQ$ ...(i)

At equilibrium, $Q = K_{eq}$ and $E_{cell} = 0)$

$0 = E^\circ{}_{cell} - \frac{0.059}{1} logK_{eq}$ (from equation (i))

$logK_{eq} = \frac{E^\circ{}_{cell}}{0.059} = \frac{0.59}{0.059} = 10$

$K_{eq} = 10^{10} = 1 \times 10^{10}$

S20. Ans.(a)

$\Delta G^\ominus = -nF\ E^\ominus_{cell}$

$= -2 \times 96500 \times 0.24$

$= -46320\ Jmol^{-1} = \frac{-46320}{1000}$

$= -46.32\ kJ/mol$

S21. Ans.(c)

$\overset{+1}{HBrO} \rightarrow \overset{0}{Br_2}, E^0_{HBrO/Br_2} = 0.595\ V'$

$\overset{+1}{HBrO} \rightarrow \overset{+5}{BrO_3^-}, E^\circ_{BrO_3/HBrO} = 1.5\ V$

E°_{cell} for the disproportionation of $HBrO$,

$E^\circ_{cell} = E^\circ_{HBrO/Br_2} - E^\circ_{BrO_3^-/HBrO}$

$= 1.595 - 1.5$

$= 0.095\ V = +ve$

Hence, option (c) is correct answer.

S22. Ans.(d)

$Zn|ZnSO_4(0.01M)||CuSO_4(1.0M)|Cu$

$\therefore\ E_1 = E^\circ_{cell} - \frac{2.303RT}{2\times F} \times \log\frac{(0.01)}{1}$

When concentrations are changed

$\therefore\ E_2 = E^\circ_{cell} - \frac{2.303RT}{2F} \times \log\frac{1}{0.01}$

$i.e,\ E_1 > E_2$

S23. Ans.(c)

$\Lambda^\circ_m(AgCl) = \Lambda^\circ_m(AgNO_3) + \Lambda^\circ_m(KCl) - \Lambda^\circ_m(KNO_3)$

$= (133.4 + 149 - 144.9)\ S\ cm^2\ mol^{-1}$

$= 138.4\ S\ cm^2\ mol^{-1}$

S24. Ans.(c)

$Zn^{2+} + 2e^- \rightarrow Zn, E^\circ = -0.760\ V$

$Ag_2O + H_2O + 2e^- \rightarrow 2Ag + 2OH^-,$

$E^\circ = 0.344\ V$

Both are reducing potential :

As, $E^\circ_{Ag_2O/2Ag} > E^\circ_{Zn^{2+}/Zn}$

$\therefore$ Cell reaction will be

At anode; $Zn \rightarrow Zn^{2+} + 2e^-$ E° = +0.760 V

At cathode;

$Ag_2O + H_2O + 2e^- \rightarrow 2Ag + 2OH^-$

$E^\circ = +0.344\ V$

Cell reaction:

$Zn + Ag_2O + H_2O \rightarrow Zn^{2+} + 2Ag + 2OH^-$

$n = 2$ $E^\circ_{cell} = 1.104\ V$

$\because\ \Delta G^\circ = -nFE^\circ_{cell}$

$\Delta G^\circ = -2 \times 96500 \times 1.10$

$= -213072\ J\ mol^{-1}$

$= -213.072\ kJ\ mol^{-1}$

$|\Delta G^\circ| = 213.072\ kJ\ mol^{-1}$

S25. Ans.(d)

According to Faraday's first law:

w = z.i.t

$z = \frac{E}{96500}(molar\ mass)$

$0.1 \times 71 = \frac{35.5}{96500} \times 3 \times t$ as n-factor is 2

$t = 6433$ sec = 107.2 min

$\simeq$ 110 min

S26. Ans.(b)

Zn have a higher (–ve) electrode potential that is more reactive than Fe. It is coated on iron substances to provide resistance against rusting such a process is called galvanization. But in reverse, that is Fe cannot be coated on Zn, as corrosion will occur. In above, Zn displaces Fe from its salt solution.

S27. Ans.(a)

According to Faraday's law

Q = ne

Q = it

ne = it

$n = \frac{1\times60}{1.6\times10^{-19}} = 3.75\times10^{20}\ electrons$

S28. Ans.(c)

$\Delta G^\circ = -nFE^\circ\ cell \qquad E^\circ\ cell = (-ve)$

So, $\Delta G^\circ = (+ve) \quad \Delta G > 0$

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

$\therefore \quad K_{eq} < 1$

S29. Ans.(d)

Concentration = 0.5 mol dm^{-3},

K = 5.76×10^{-3} S cm^{-1}

T = 298 K

$\lambda_m = \frac{K\times1000}{M}$

$= \frac{5.76\times10^{-3}}{0.5} = 11.52\ S\ cm^2\ mol^{-1}$

S30. Ans.(b)

$2H^+ + 2e^- \rightarrow H_2$

$E = E^\circ - \frac{0.059}{2}\log\frac{P_{H_2}}{(H^+)^2}$

$0 = 0 - \frac{0.059}{2}\log\frac{P_{H_2}}{(10^{-7})^2}$

$\log 1 = 0$

$P_{H_2} = (10^{-7})^2 = 10^{-14}\ atm$.

S31. Ans.(d)

S32. Ans.(b)

Sn^{2+} cannot reduce Fe^{2+}, so $FeCl_2$ and $SnCl_2$ can exists together.

S33. Ans.(a)

$\Delta G^\circ = -2.303\ RT\log K_{sp}$

$63300 = -2.303\times8.314\times298\log K_{sp}$

$K_{sp} \sim 8\times10^{-12}$

S34. Ans.(c)

$\overset{+6}{MnO_4^{-2}} \quad\rightarrow\quad MnO_4^- + e^-$

0.1 mole   0.1 mole

Q = nF = 0.1 × 96500 = 9650 C

S35. Ans.(c)

$W_{O_2} = \frac{5600}{22400}\times32 = 8g = 1\ equivalent$

$= 1\ equivalent\ of\ Ag$

$= 108\ g$

S36. Ans.(c)

According to Kohlrausch's law of limiting molar conductivity:

$\alpha = \frac{\Lambda_m}{\Lambda_m^\circ} = \frac{9.54}{238} = 0.04008$

% Dissociation (α) = 4.008%

S37. Ans.(b)

pH = 10

Oxidation potential of standard hydrogen electrode is given by:

$E_{op} = 0.0592\ (pH)_{Anode}$

= 0.0592 × 10

E_{op} = 0.59 V

S38. Ans.(a)

$E^\circ cell = E^\circ_{cathode} - E^\circ_{anode}$

$= 0.76 - (-0.34)$

$= 1.1\ V$