

## <H1> Additional Objective Questions

### <H2> Single Correct Choice Type

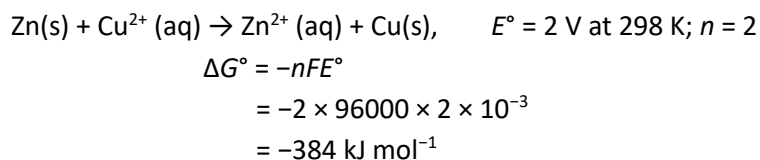
1.(B) Concept based.

2.(C) Faraday's second law states that "The number of Faradays passed is equal to the number of gram equivalents of electrolytes discharged".

3.(A) Iron loses electrons to form  $\text{Fe}^{2+}$  only when it acts as anode.

4.(D) Impure copper is made anode and cathode is a sheet of pure copper. As the current is passed, the positive  $\text{Cu}^{2+}$  ions are attracted to the negative cathode where they take up the electrons and deposit themselves as neutral copper atoms.

5.(A) For reaction



6.(A) Best reducing agent is one which is itself oxidized most easily, namely, A.

7.(C) We know

$$\text{Cell constant} = \frac{\text{Specific conductance of KCl}}{\text{Conductance of KCl}} = \frac{xS}{yS}$$

For 0.01 M  $\text{H}_2\text{SO}_4$ ,

Specific conductance = Cell constant  $\times$  Conductance =  $xz/y$ .

8.(C) For the given cell reaction

$$\begin{aligned}\Delta G^\circ &= -nFE_{\text{cell}}^\circ = -RT \ln K_{\text{eq}} \\ -5 \times 96500 \times 0.59 &= 2.303 \times 8.314 \times 298 \log K_{\text{eq}}\end{aligned}$$

Solving, we get  $K_{\text{eq}} = 7.8 \times 10^{49} \approx 10^{50}$ .

9.(A) We know

$$E_{\text{cell}}^\circ = (\text{SRP})_{\text{cathode}} - (\text{SRP})_{\text{Anode}}$$

To get the highest  $E^\circ_{\text{cell}}$  the S.R.P value of cathode must be highest. So,  $\text{Au}^{3+}/\text{Au}(\text{s})$  has highest value of SRP 1.40 but per electron S.R.P values is  $\frac{1.40}{3} = 0.467 \text{ V}$ . For  $\text{Ag}^+/\text{Ag}$  per electron S.R.P values is 0.80 V which is highest among the given electrodes.

**10.(A)** Greater the value of standard reduction potential, greater will be its tendency to undergo reduction. So, the sequence of deposition of metals on cathode will be Ag, Hg, Cu. Here, magnesium will not be deposited because its standard reduction potential is negative. So, it is a strong tendency to undergo oxidation. Therefore, on electrolysis of  $\text{Mg}(\text{NO}_3)_2$  solution,  $\text{H}_2$  gas will be evolved at cathode.

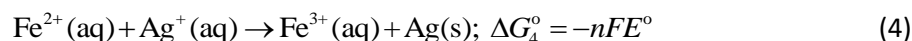
**11.(C)**  $\text{Cu}^{2+}$  ions gain electrons and are deposited on copper.

**12.(B)** Concept based.

**13.(C)** We have



Thus, for the given cell reaction



Equating Eq. (1), (2), and (3) with Eq. (4), we get

$$\Delta G_4^\circ = \Delta G_1^\circ + \Delta G_2^\circ + \Delta G_3^\circ$$

$$-nFE^\circ_{\text{cell}} = -Fx - 2y + 3z$$

$$E^\circ_{\text{cell}} = (x + 2y - 3z) \text{ V}$$

**14.(B)** Given



For  $\text{Cu}^{2+} + e^- \rightarrow \text{Cu}^+; E^\circ_{\text{Cu}^{2+}/\text{Cu}^+} = x \text{ V}$

On reversing Eq. (2), we get



From Eq. (1) and Eq. (3), we get

$$\begin{aligned} \Delta G_3^\circ &= \Delta G_1^\circ + \Delta G_2^\circ \\ \Rightarrow E_3^\circ &= \frac{F(2E_1^\circ + E_2^\circ)}{F} \\ \Rightarrow E_3^\circ &= 2 \times 0.34 + (-0.522) \\ \Rightarrow E_3^\circ &= 0.68 - 0.522 = 1.58 \text{ V} \end{aligned}$$

**15.(D)** According to Kohlrausch law of independent migration of ions, the incorrect equation is:

$$\left(\Lambda_m^0\right)_{\text{NaBr}} - \left(\Lambda_m^0\right)_{\text{NaI}} = \left(\Lambda_m^0\right)_{\text{KBr}} - \left(\Lambda_m^0\right)_{\text{NaBr}}$$

From L.H.S

$$\begin{aligned} \left(\Lambda_m^0\right)_{\text{NaBr}} - \left(\Lambda_m^0\right)_{\text{NaI}} &= \lambda_{m\text{Na}^+}^0 + \lambda_{m\text{Br}^-}^0 - \lambda_{m\text{Na}^+}^0 - \lambda_{m\text{I}^-}^0 \\ &= \lambda_{m\text{Br}^-}^0 - \lambda_{m\text{I}^-}^0 \end{aligned}$$

From R.H.S

$$\begin{aligned} \left(\Lambda_m^0\right)_{\text{KBr}} - \left(\Lambda_m^0\right)_{\text{NaBr}} &= \lambda_{m\text{K}^+}^0 + \lambda_{m\text{Br}^-}^0 - \lambda_{m\text{Na}^+}^0 - \lambda_{m\text{Br}^-}^0 \\ &= \lambda_{m\text{K}^+}^0 - \lambda_{m\text{Na}^+}^0 \end{aligned}$$

Hence, L.H.S is not equal to R.H.S.

**16.(B)** Given that  $i = 0.193 \text{ A}$ ,  $t = 1 \text{ h} = 60 \times 60 \text{ s}$ . Hence,

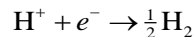
$$Q = it = 0.193 \times 3600 = 694.8 \text{ C} = 7.2 \times 10^{-3} \text{ F}$$

$$[\text{OH}^-] = \frac{7.2 \times 10^{-3}}{0.1} = 7.2 \times 10^{-2} \text{ M}$$

Therefore,  $\text{pOH} = 2 - \log 7.2$ ;  $\text{pH} = 14 - \text{pOH}$ ;  $= 12 + \log 7.2 = 12.82$

**17.(C)** For water at 298 K,  $[\text{H}^+] = 10^{-7} \text{ M}$

Reduction reaction is

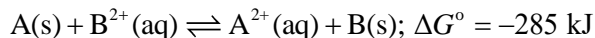


Therefore,

$$\begin{aligned} E_{\text{cell}} &= -\frac{RT}{F} \ln \frac{(p_{\text{H}_2})^{1/2}}{[\text{H}^+]} = -0.591 \log \frac{(p_{\text{H}_2})^{1/2}}{[\text{H}^+]} \\ E_{\text{cell}} &= -0.0591 \log \frac{1}{10^{-7}} = -0.4137 \approx 0.414 \text{ V} . \end{aligned}$$

**18.(A)** Among the given electrolytes, HCOOH is the strongest electrolyte and CH<sub>3</sub>COOH is the weakest electrolyte. Hence, the correct order of electrical conductivity is HCOOH > C<sub>6</sub>H<sub>5</sub>COOH > CH<sub>3</sub>COOH.

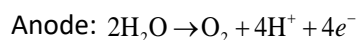
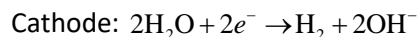
**19.(D)** For the given reaction,



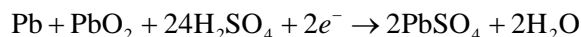
But  $\Delta G^\circ = -nFE^\circ \Rightarrow -285 \times 1000 = -2 \times 96500 E^\circ \Rightarrow E^\circ = 1.48 \text{ V}$

Since efficiency is 84%, actual  $E^\circ = 1.48 \times \frac{84}{100} = 1.24 \text{ V}$ .

**20.(A)** The reactions at the cathode and anode are



**21.(C)** The following reactions take place in a lead storage battery during discharging process:



In the overall reaction, H<sub>2</sub>SO<sub>4</sub> is consumed and water is formed (Pb, PbSO<sub>4</sub>, and PbO<sub>2</sub> are solids). Therefore, the density of H<sub>2</sub>SO<sub>4</sub> solution decreases.

**22.(A)** In the given reaction, I<sup>-</sup> has been oxidized to I<sub>2</sub> and Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup> ions have been reduced to Cr<sup>2+</sup>.

Therefore,

$$E_{\text{cell}}^\circ = E_{\text{Cr}_2\text{O}_7^{2-}}^\circ - E_{\text{I}_2}^\circ$$

$\Rightarrow$

$$0.79 = 1.33 - E_{\text{I}_2}^\circ \text{ or } E_{\text{I}_2}^\circ = 0.54 \text{ V}$$

**23.(B)** We know

$$\text{Conductivity} \propto \frac{\text{Area} \times \text{Concentration}}{\text{Length}}$$

or  $\text{Conductivity} = \kappa \times \frac{\text{Area} \times \text{Concentration}}{\text{Length}}$

or  $\kappa = \frac{\text{Conductivity} \times \text{Length}}{\text{Area} \times \text{Concentration}} = \frac{\text{S} \times \text{m}}{\text{m}^2 \times \text{mol m}^{-3}} = \text{S m}^2 \text{ mol}^{-1}$

**24.(B)** For the given reaction,

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.059}{2} \log K$$

At equilibrium,  $E_{\text{cell}} = 0$ . Therefore,  $E_{\text{cell}}^\circ = \frac{0.059}{2} \log K$

$$0.23 = \frac{0.059}{2} \log K \Rightarrow K = 6.26 \times 10^7$$

**25.(A)** For the cell reaction  $\text{Ag(s)} + \frac{1}{2} \text{Cl}_2(\text{g}) \rightarrow \text{AgCl(s)}$ ;  $E^\circ = -1.14\text{V}$

or 
$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0591}{1} \log p_{\text{Cl}_2}^{1/2}$$

Under standard conditions  $p_{\text{Cl}_2} = 1$ , therefore  $\log p_{\text{Cl}_2}^{1/2} = 0$ .

Useful work =  $-W_{\text{max}} = -nFE = (-1) \times (-1.14) \times 96500 \times 10^{-3} \text{ kJ} = 110 \text{ kJ mol}^{-1}$ .

**26.(B)** For the given reaction,

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{RT}{nF} \ln \frac{[\text{Zn}^{2+}][\text{Cl}^-]^2}{p_{\text{Cl}_2}}$$

**27.(B)** Given that  $\text{Cu}^+ + e^- \rightarrow \text{Cu}$   $E_1^\circ = 0.522 \text{ V}$   $n = 1$  (1)

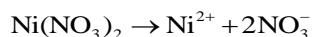
$\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$   $E_2^\circ = 0.3402 \text{ V}$   $n = 2$  (2)

$\text{Cu}^{2+} + e^- \rightarrow \text{Cu}^+$   $E_3^\circ = ?$   $n = 1$  (3)

Equation (2)–(1) gives Eq. (3), therefore,

Since 
$$E_3^\circ = \frac{n_1 E_1^\circ + n_2 E_2^\circ}{n_3} = 1 \times \frac{(-0.522) + 2 \times 0.3402}{1} = -0.522 + 0.6802 = 0.1584 \text{ V}$$

**28.(A)** We know



Reaction takes place at cathode:  $\text{Ni}^{2+} + 2e^- \rightarrow \text{Ni}$

On passing 2 F of electricity 1 mol of mol of Ni gets deposited.

Hence, 0.1 F of electricity will deposit  $\frac{1}{2} \times 0.1 = 0.05 \text{ mol Ni}$ .

**29.(C)** Faraday's second law states that  $\frac{W}{E} = \text{constant}$

So, g equiv. of  $\text{Ag}^+ = \text{g equiv. of Ca}^{2+} = \text{g equiv. of Al}^{3+} = a$  (say)

Now, mole of  $\text{Ag}^+ = a^+$  ( $\text{Ag}^+ + e^- \rightarrow \text{Ag}$ )

$$\text{Mole of Ca}^{+2} = \frac{W}{E} \quad (\text{Ca}^{+2} + 2e^{-} \rightarrow \text{Ca})$$

$$\text{Mole of Al}^{3+} = \frac{a}{3} \quad (\text{Al}^{3+} + 3e^{-} \rightarrow \text{Al})$$

So, mole ratio of  $\text{Ag}^{+} : \text{Ca}^{+2} : \text{Al}^{3+} = a : \frac{W}{E} : \frac{W}{E} = 6 : 3 : 2$

**30.(A)** The reaction is  $\text{Zn} + \text{MgCl}_2 \rightarrow \text{ZnCl}_2 + \text{Mg}$ . Therefore,

$$E_{\text{cell}}^{\circ} = E_{\text{Zn/Zn}^{2+}}^{\circ} + E_{\text{Mg}^{2+}/\text{Mg}}^{\circ} = +0.762 - 2.37 = -1.608 \text{ V}$$

Therefore  $E_{\text{cell}}^{\circ} = E_{\text{Zn/Zn}^{2+}}^{\circ} + E_{\text{Mg}^{2+}/\text{Mg}}^{\circ}$  is negative so no reaction will take place.

**31.(C)** Relation between equilibrium constant and  $E_{\text{cell}}^{\circ}$  is

$$E_{\text{cell}}^{\circ} = 2.303 \frac{RT}{nF} \log K_C \text{ at } 298 \text{ K}$$

[ $n = 2$  from the reaction]

$$\Rightarrow E_{\text{cell}}^{\circ} = \frac{0.059}{2} \log(10 \times 10^{15})$$

$$\Rightarrow E_{\text{cell}}^{\circ} = \frac{0.059}{2} \times 16 = 0.472 \text{ V}$$

**32.(D)**  $\Delta G^{\circ} = -nFE_{\text{cell}}^{\circ} = -2 = -2 \times 96500 \times 0.6753$  (at  $25^{\circ}\text{C}$ ) =  $-130333 \text{ J}$

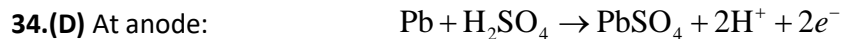
$\Delta S = nf \left( \frac{\partial E}{\partial T} \right)$  where  $\left[ \frac{\partial E}{\partial T} \right]$  is called temperature coefficient of emf.

$$\left( \frac{\partial E}{\partial T} \right)_p = \frac{E_2 - E_1}{T_2 - T_1} = \frac{0.6753 - 0.6915}{298 - 273} = -6.48 \times 10^{-4}$$

Therefore,  $\Delta S = 2 \times 96500 \times (-6.48 \times 10^{-4}) = -125.064 \text{ JK}^{-1} \text{ mol}^{-1}$ .

Substituting the values in  $\Delta G = \Delta H - T\Delta S$ , we get  $\Delta H = -167.6 \text{ kJ}$ .

**33.(A)** Equivalent conductance is the conducting power of all the ions produced by one gram equivalent of an electrolyte in a given solution. In  $\text{SrCl}_2$ , gram equivalent mass is larger as a result of which the compound shows maximum value of equivalent conductance in a fused state.



Net discharging equation  $\text{Pb} + \text{PbO}_2 + 2\text{H}_2\text{SO}_4 \rightarrow 2\text{PbSO}_4 + 2\text{H}_2\text{O}$

So,  $\text{H}_2\text{SO}_4$  is consumed during discharging of lead storage battery.

**35.(B)** Cell reaction:  $3\text{Ni} + 2\text{Au}^{3+} \rightarrow 3\text{Ni}^{2+} + 2\text{Au}$

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{6} \log \frac{[\text{Ni}^{2+}]^3}{[\text{Au}^{3+}]^2} = (0.25 + 1.5) - \frac{0.0591}{6} \log \frac{(0.1)^3}{(1)^2}$$

$$= 1.75 + \frac{0.0591}{2} \times \log(1) = 1.75 + 0.295 = +1.7795 \text{ V}$$

**36.(D)** We know

$$\Delta G^{\circ} = -RT \ln K \quad (1)$$

$$\Delta G^{\circ} = -nFE_{\text{cell}}^{\circ} \quad (2)$$

From Eq. (1) and (2), we get

$$\ln K = \frac{nFE^{\circ}}{R \times T} = \frac{2 \times 96000 \times 2}{8 \times 300}$$

$$\ln K = 160$$

or

$$K = e^{160}$$

**37.(D)** The reaction involved is  $\text{BaSO}_4 \rightleftharpoons \text{Ba}^{2+} + \text{SO}_4^{2-}$

Given that  $\kappa(\text{BaSO}_4) = 3.06 \times 10^{-6} \text{ S cm}^{-1}$  and  $\lambda(\text{BaSO}_4) = 1.53 \text{ S cm}^2 \text{equiv.}^{-1}$

Let the solution contain  $x$  mol of Ba and  $x$  mol of  $\text{SO}_4$  and they are produced from  $x \text{ mol L}^{-1}$  of  $\text{BaSO}_4$

$$\lambda = \kappa \times V = 3.06 \times 10^{-6} \times \frac{1000}{x} = 1.53$$

where  $x = 2 \times 10^{-3}$ ,  $[\text{Ba}^{2+}] = 2 \times 10^{-3}$ ;  $K_{\text{sp}} = 4 \times 10^{-6}$ .

**38.(A)** Given that  $\left(\frac{\partial E}{\partial T}\right)_p = -4.0 \times 10^{-5} \text{ V deg}^{-1}$ ,  $n = 2$  and  $E_{\text{cell}}^{\circ} = 1.0181 \text{ V}$ .

Hence,

$$\Delta G^{\circ} = -2FE_{\text{cell}}^{\circ} = -2 \times 96500 \times 1.018 = -196474 \text{ J} = -196.5 \text{ kJ}$$

$$\Delta S^{\circ} = nF \left(\frac{\partial E}{\partial T}\right) = 2 \times 96500 \times 4 \times 10^{-5} = -7.72 \text{ J}$$

$$\Delta H^{\circ} = nF \left( T \left(\frac{\partial E}{\partial T}\right)_p - E_{\text{cell}}^{\circ} \right) = 198.8 \text{ kJ}$$

**39.(B)** We know

$$\Lambda_m^\circ(\text{CaCl}_2) = \lambda_m^\circ(\text{Ca}^{2+}) + 2\lambda_m^\circ(\text{Cl}^-) \quad (1)$$

Substituting the values in Eq. (1), we get

$$\Lambda_m^\circ(\text{CaCl}_2) = 118.88 \times 10^{-4} + 2(77.33 \times 10^{-4})$$

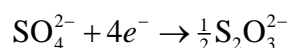
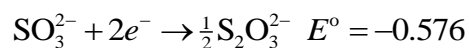
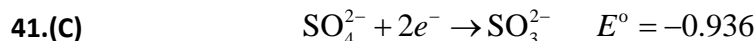
$$\Lambda_m^\circ(\text{CaCl}_2) = 273.54 \times 10^{-4} \text{ S mm}^2 \text{ mol}^{-1}$$

**40.(B)** Number of Faradays passed =  $\frac{9.65 \times 1 \times 3600}{96500}$

Let  $W_{\text{Zn}}$  grams of zinc be discharged at anode and  $W_{\text{Cu}}$  grams of copper be discharged at cathode.

$$\text{Therefore } \frac{W_{\text{Zn}}}{65.4} \times \frac{W_{\text{Cu}}}{63.5} \times = \frac{9.65 \times 3600 \times 1}{96500}$$

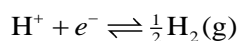
$$W_{\text{Zn}} = 11.77, W_{\text{Cu}} = 11.43 \text{ g.}$$



$$E^\circ = \frac{-2 \times 0.936 - 2 \times 0.576}{4} = -0.756 \text{ V}$$

**42.(B)** At infinite dilution, the mobility of large ion is more than small ions, therefore, the correct order of equivalent conductance at infinite dilution is iii > ii > i.

**43.(D)** For the given hydrogen electrode,



From Nernst equation, we have

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0591}{n} \log Q \quad (1)$$

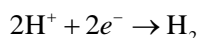
Substituting the values in Eq. (1), we get

$$E_{\text{cell}} = 0.0 - \frac{0.0591}{1} \log \frac{(p\text{H}_2)^{1/2}}{[\text{H}^+]}$$

$$\Rightarrow E_{\text{cell}} = \frac{0.0591}{1} \log \frac{1}{10^{-7}}$$

$$\Rightarrow E_{\text{cell}} = -0.0591 \times 7 \times \log 10 = -0.413 \text{ V}$$

**44.(D)** The reaction occurring is





Thus, 1 mol of H<sub>2</sub>, that is, 22400 cm<sup>3</sup> at NTP requires 2F = 2 × 96500 C

Therefore, 4480 cm<sup>3</sup> requires =  $\frac{2 \times 96500 \times 4480}{22400} = 38600 \text{ C}$

As  $Q = I \times t \Rightarrow t = \frac{Q}{I} = \frac{38600}{2.14} = 18037.38 \text{ s} = 5 \text{ h}$

**45.(C)** From Nernst equation, we have

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log Q \quad (1)$$

Substituting the values in Eq. (1), we get

$$0.62 = 0.8 - \frac{0.06}{1} \log \frac{[\text{H}^+]}{[\text{Ag}^+]}$$

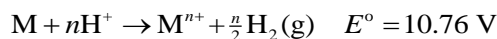
$$\Rightarrow -0.18 = 0.06 \log \frac{1}{[\text{Ag}^+]} \Rightarrow [\text{Ag}^+] = 10^{-3} \text{ M}$$

Weight of Ag =  $10^{-3} \times 108 = 0.108 \text{ g}$

Therefore, weight% of Ag =  $\frac{0.108}{1.08} \times 100 = 10\%$

**46.(B)** At the anode:  $\text{M} - ne^- \rightarrow \text{M}^{n+}$   $E^{\circ} = +0.76$   
0.02M

At the cathode:  $n\text{H}^+ + ne^- \rightarrow \frac{n}{2} \text{H}_2(\text{g})$   $E^{\circ} = 0$   
(1 M)



$$E_{\text{cell}} = 0.81 = +0.76 - \frac{0.0591}{n} \log \frac{[\text{M}^{n+}]}{[\text{H}^+]^n} = 0.81 = +0.76 - \frac{0.0591}{n} \log \frac{[\text{M}^{n+}]}{[\text{H}^+]^n}$$

$$\frac{0.0591}{n} \times 1.7 = 0.81 - 0.76 \Rightarrow n = \frac{0.0591 \times 1.7}{0.05} = 2.$$

**47.(B)** The reaction involved is  $\text{MnO}_4^- + 5e^- + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$

According to Nernst equation,

$$E_{\text{red}} = E_{\text{red}}^{\circ} - \frac{0.059}{5} \log \left[ \frac{[\text{Mn}^{2+}]}{[\text{MnO}_4^-][\text{H}^+]^8} \right]$$

Let  $[\text{H}^+]_{\text{initial}} = x/100 = x/10^2$ , therefore,

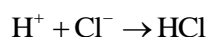
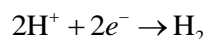
$$\begin{aligned} E_{\text{red}(\text{initial})} &= E_{\text{red}}^{\circ} - \frac{0.0591}{5} \log \left[ \frac{[\text{Mn}^{2+}] \times 10^{16}}{[\text{MnO}_4^-][X]^8} \right] \\ &= \frac{-0.0591}{5} \log 1016 = -0.1891 \text{ V} \end{aligned}$$

This  $E_{\text{red}(\text{initial})}$  decreases by 0.189 V. The tendency of the half-cell to get reduced is its oxidizing power. Hence, the oxidizing power decreases by 0.189 V.

$$48.(B) \alpha = \frac{\Lambda_m}{\Lambda_m^0} = \frac{\Lambda_m}{425} = 0.9 \Rightarrow \Lambda_m = 382.5 \text{ S cm}^{-1} \text{ g. equiv}^{-1}$$

$$\text{Therefore, } \Lambda_m = \kappa \times N \Rightarrow N = \frac{382.5}{3.825} = 10 \text{ N}$$

49.(A) HCl is a strong electrolyte. The reactions involved are



50.(A) Given that  $\kappa = 1.12 \times 10^{-2} \text{ Scm}^{-1}$  for  $N/10$  KCl solution.

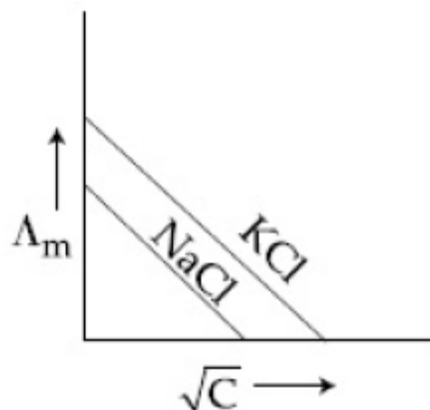
We know

$$\kappa = \frac{1}{R} \times \frac{l}{A}$$

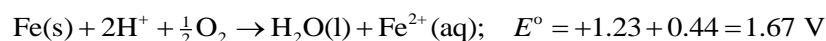
$$\Rightarrow 1.12 \times 10^{-2} = \frac{1}{65} \times \frac{l}{A}$$

$$\Rightarrow \frac{l}{A} = 0.725 \text{ cm}^{-1}$$

51.(A) KCl has higher molar conductance than that of NaCl hence, the correct graph will be



52.(A) The expression is  $\Delta G^0 = -nFE^0$ . Reversing the second equation and adding to the first equation, we get the net reaction as



Therefore,  $\Delta G^\circ = -nFE^\circ = -2 \times 96500 \times 1.67 = 322310 \text{ J mol}^{-1} = 322.31 \text{ kJ mol}^{-1}$

**53.(B)** Specific conductance ( $\kappa$ )<sub>AgCl</sub> =  $\kappa_{\text{solution}} - \kappa_{\text{water}} = 1.86 \times 10^{-6} - 6 \times 10^{-8} = 180 \times 10^{-8}$

Therefore, solubility is found as

$$S_{\text{AgCl}} = \kappa_{\text{AgCl}} \times \frac{1000}{\Lambda_{\text{eq}}^\circ} = 180 \times 10^{-8} \times \frac{1000}{137.2} = 1.3 \times 10^{-5} \text{ mol L}^{-1}$$

**54.(C)** From Faraday's second law, we have

$$\frac{W}{E} = \text{constant}$$

For given electrodes reactions,



Therefore, 
$$\frac{W_{\text{Ag}}}{E_{\text{Ag}}} = \frac{W_{\text{Cu}}}{E_{\text{Cu}}} \quad (1)$$

Substituting the values in Eq. (1), we get

$$\frac{W_{\text{Ag}}}{E_{\text{Ag}}} = \frac{W_{\text{Cu}}}{E_{\text{Cu}}} = \frac{107.9}{63.54/2} = \frac{107.9}{31.77}$$

**55.(B)** The reaction is  $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$

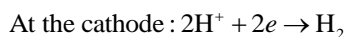
Now, 
$$\text{Volume} = \frac{\text{Mass}}{\text{Density}} = \text{Area} \times \text{Thickness deposited} = 100 \times 10^{-2} \text{ cm}^3$$

Mass of  $\text{Cu}^{2+}$  ion deposited on plate =  $1 \times 8.94 \text{ g}$

Now, 
$$\text{Mass} = \frac{EIt}{96500} = \frac{Eq}{96500}$$

$$8.94 = \frac{63.5 \times Q}{2 \times 96500} \Rightarrow Q = 27172 \text{ C}$$

**56.(A)** The half-cell reactions in the cell are:



The EMF of the cell is given by

$$E_{\text{cell}} = \frac{0.0591}{2} \log \frac{[\text{H}^+]_{\text{cathode}}^2}{[\text{H}^+]_{\text{anode}}^2} \quad (1)$$

Substituting the values in Eq. (1), we get

$$E_{\text{cell}} = \frac{0.0591}{2} [2\log[\text{H}^+]_{\text{cathode}} - 2\log[\text{H}^+]_{\text{anode}}]$$

$$\Rightarrow E_{\text{cell}} = \frac{0.0591}{2} (-2\text{pH}_{\text{cathode}} + 2\text{pH}_{\text{anode}})$$

$$\Rightarrow E_{\text{cell}} = \frac{0.0591}{2} \times 6 = 0.177 \text{ V}$$

**57.(D)** The reaction is  $\text{C}_6\text{H}_5\text{NO}_2 + 6\text{H}^+ + 6\text{e}^- \rightarrow \text{C}_6\text{H}_5\text{NH}_2 + 3\text{H}_2\text{O}$

6 F is required to reduce 1 mol of nitrobenzene to aniline. So,  $6 \times 96500 = 57900 \text{ C}$ .

**58.(B)** Using Kohlrausch's law

$$\Lambda_{m, \text{BaCl}_2}^0 = \lambda_{\text{Ba}^{2+}}^0 + 2\lambda_{\text{Cl}^-}^0 = x_1 \quad (1)$$

$$\Lambda_{m, \text{H}_2\text{SO}_4}^0 = 2\lambda_{\text{H}^+}^0 + 2\lambda_{\text{SO}_4^{2-}}^0 = x_2 \quad (2)$$

$$\Lambda_{m, \text{HCl}}^0 = \lambda_{\text{H}^+}^0 + 2\lambda_{\text{Cl}^-}^0 = x_3 \quad (3)$$

Adding the Eqs. (1) and (2) and subtracting  $2 \times \text{Eq. (3)}$ , we get

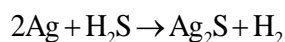
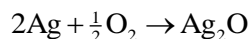
$$\Lambda_{m, \text{BaSO}_4}^0 = x_1 + x_2 - 2x_3$$

### <H2> Multiple Correct Choice Type

**1.(A, D)** Reduction potential means the means the ability of the species to accept electrons. Metals are good contributor of electrons. That is why down the group (which is arranged in order of increasing reduction potential) reactivity decreases

**2.(A, B)** The standard reduction potential of copper and silver is greater than the standard reduction potential of hydrogen. So, hydrogen gas reduces cupric oxide and silver oxide.

**3.(A, B)** The tarnishing of silver ornaments in atmosphere is due to the formation of silver oxide and silver sulphide. The reaction involved is



**4.(A, D)** The reaction involved is  $2\text{Tl} + \text{Cu}^{2+} \rightarrow 2\text{Tl}^+ + \text{Cu}$

$$E_{\text{cell}} = E_{\text{cell}}^0 - \frac{0.0591}{2} \log \frac{[\text{Tl}^{2+}]}{[\text{Cu}^{2+}]}$$

**5.(A, C)** Due to the formation of  $\text{Cu}^{2+}$  ions in the solution.

**6.(A, B)** Concept based.

**7.(A, D)** With silver, lead and zinc ions present in electrolyte, KCl cannot be used as these ions are insoluble in water and precipitation will take place.

8.(A, B)

$$\text{Option (A): } E_{\text{cell}} - E_{\text{cell}}^{\circ} = \frac{0.0591}{2} \log \frac{[\text{Ag}^+]^2}{[\text{Zn}^{2+}]} = \frac{0.0591}{2} \log \frac{[0.1]^2}{[0.01]} = 0$$

$$\text{Option (B): } E_{\text{cell}} - E_{\text{cell}}^{\circ} = \frac{0.0591}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{H}^+]^2} = \frac{0.0591}{2} \log \frac{[0.01]}{[0.1]^2} = 0$$

$$\text{Option (C): } E_{\text{cell}} - E_{\text{cell}}^{\circ} = \frac{0.0591}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{H}^+]^2} = \frac{0.0591}{2} \log \frac{[1]}{[0.1]^2} = \frac{0.0591}{2} \log 100 = 0.0591 \text{ V}$$

$$\text{Option (D): } E_{\text{cell}} - E_{\text{cell}}^{\circ} = \frac{0.0591}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{H}^+]^2} = \frac{0.0591}{2} \log \frac{[0.01]}{[0.01]^2} = \frac{0.0591}{2} \log 100 = 0.0591 \text{ V}$$

9.(A, C) If  $\left(\frac{\partial E}{\partial T}\right) < 0$

$$\Rightarrow \Delta H > nFE$$

Therefore, the process inside the cell is exothermic.

10.(A, C) We know

$$\frac{W}{E} = \frac{It}{n \times 96500} = \frac{9.65 \times 3600}{96500n} = 0.36$$

equiv. of  $\text{Ag}^+ = 0.36$  mol of  $\text{Ag}^+$  ( $n = 1$ ) and 0.18 mol of  $\text{Cu}^{2+}$  ( $n = 2$ ).

Now, Ag will oxidize to  $\text{Ag}^+$  and  $\text{Cu}^{2+}$  will reduce to Cu.

11.(B, C) According to Faraday's law of electrolysis, the amount ( $W$ ) element deposited at electrode is directly proportional to its equivalent weight ( $E$ ).

$$\frac{W}{E} = \text{Constant}$$

Or  $W \propto E$

For,  $\overset{+2}{\text{Fe}}\text{SO}_4 \rightarrow \overset{0}{\text{Fe}}$  (Deposited); the equivalent weight of Fe =  $\frac{M}{2}$

For,  $\overset{+3}{\text{Fe}_2}(\text{SO}_4)_3 \rightarrow \overset{0}{\text{Fe}}$  (Deposited); the equivalent weight of Fe =  $\frac{M}{3}$

For,  $\overset{+3}{\text{Fe}}(\text{NO}_3)_3 \rightarrow \overset{0}{\text{Fe}}$  (Deposited); the equivalent weight of Fe =  $\frac{M}{3}$

Therefore, the amount of iron deposited in  $\text{Fe}_2(\text{SO}_4)_3$  and  $\text{Fe}(\text{NO}_3)_3$  is equal.

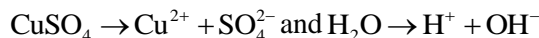
In all three electrolytes, water get oxidized to oxygen gas.

12.(A, B, D) Conductance of an electrolyte solution increases if the interionic forces of attraction are high.

13.(A, B, C) It contains platinum electrode which is inert.

**14.(A, B, D)** Since, concentration cells have the same electrodes with different concentrations of metal-ion, or gas-ion or different electrode pressure. Here in option, **(C)**, the electrons as well as concentration both are different, while other have same electrode and different concentrations.

**15.(B, C, D)** The reactions involved are



At the cathode:  $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu(s)}$

At the anode:  $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu(s)}$

At the cathode:  $\text{Ni} - 2e^- \rightarrow \text{Ni}^{2+}(\text{aq})$

**16.(A, D)**  $\Delta G = \Delta H - T\Delta S$  and  $\Delta G = \Delta H + T\left(\frac{\partial \Delta G}{\partial T}\right)_p$

Therefore, 
$$\left(\frac{\partial \Delta G}{\partial T}\right)_p = \frac{\Delta G - \Delta H}{T} = -\frac{T\Delta S}{T} = -\Delta S$$

Also, 
$$-nEF = \Delta H + T \times \left(\frac{\partial E}{\partial T}\right)_p$$

Therefore, 
$$\left(\frac{\partial E}{\partial T}\right)_p = \frac{\Delta H + nEF}{T}$$

**17.(C, D)** Saturated solution of  $\text{KNO}_3$  is used to make 'salt-bridge' because it is well soluble in water and dissociates completely in  $\text{K}^+$  and  $\text{NO}_3^-$  ions. Velocities of both  $\text{K}^+$  and  $\text{NO}_3^-$  are nearly same. A salt bridge maintains the electrical neutrality of the system as  $\text{K}^+$  has very low reduction potential and  $\text{NO}_3^-$  has very low oxidation potential.

**18.(A, B, C)** Salt bridge is used to eliminate liquid junction potential raised due to different speed of ions present in cathodic and anodic compartments.

**<H2> Assertion–Reasoning Type**

**1.(D)** On dilution, the number of ions per unit volume decreases degree of ionization of weak electrolyte increases and mobility of ions increases.

**2.(A)** Standard reduction potential of zinc electrode is less than that of copper electrode.

**3.(A)** For specific conduction, the number of charge-carrying particles is less compared to the increases in area. For equivalent conductance the increases in volume is more comparable to the decrease in specific conductance.

**4.(B)** Assertion is correct, and reason is correct but not the proper explanation. Assertion is convention based but reason is related to Nernst equation.

**5.(C)** The reactions involved are  $\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^-$  and  $\text{H}_2\text{O} \rightarrow \text{H}^+ + \text{OH}^-$

At the cathode:  $2\text{H}^+ + 2e^- \rightarrow \text{H}_2(\text{g})$

At the anode:  $2\text{Cl}^- - 2e^- \rightarrow \text{Cl}_2(\text{g})$

**6.(C)** Salt bridge electrolytes maintain the electrical neutrality without mixing with the solution.

**7.(B)** On applying potential greater than 1.1 V, reaction proceeds in opposite direction and the cell now behaves as electrolytic cell.

**8.(D)** On dilution conductivity decreases, because the number of ions that carry the current in an electrolyte decreases. Hence, statement S1 is wrong.

Molar conductivity ( $\Lambda_m$ ) increases on decreasing the concentration ( $\Lambda_m$ ) increasing, because volume of electrolyte increases. Hence, statement S2 is correct.

$$\Lambda_m = \frac{1000 \times \kappa}{C}$$

### <H2> Comprehension Type

**1.(A)** We know

$$\begin{aligned} \Delta G &= -nFE^\circ \\ \Rightarrow -827 \times 1000 &= -4 \times 96500 E^\circ \\ \Rightarrow E^\circ &= 2.14 \text{ V.} \end{aligned}$$

**2.(A)** We know

$$E_{\text{cell}}^\circ = \frac{0.059}{n} \log K_C.$$

This represents a straight line passing through origin

**3.(A)** From Nernst equation at equilibrium, we have

$$E_{\text{cell}}^\circ = \frac{0.059}{n} \log K \quad (1)$$

Substituting the values in Eq. (1), we get

$$\begin{aligned} 0.46 &= \frac{0.059}{2} \log K \\ \Rightarrow \log K &= 15.6 \\ \Rightarrow K &= \text{antilog } 15.6 \end{aligned}$$

**4.(B)**  $\text{Pb} + \text{SO}_4^{2-} \rightarrow \text{PbSO}_4 + 2e^-$

$\text{PbO}_2 + 4\text{H}^+ + \text{SO}_4^{2-} + 2e^- \rightarrow \text{PbSO}_4 + 2\text{H}_2\text{O}$

Overall cell reaction is  $\text{Pb} + \text{PbO}_2 + 4\text{H}^+ + 2\text{O}_4^{2-} \rightarrow 2\text{PbSO}_4 + 2\text{H}_2\text{O}$

From the above equation, it is clear that  $E_{\text{H}_2\text{SO}_4} = \frac{98}{1}$

$$\mathbf{5.(A)} \text{ Weight of } \text{H}_2\text{SO}_4 \text{ before discharging} = 1.294 \times \frac{39}{100} \times 3500 = 1766.31 \text{ g}$$

$$\text{Weight of } \text{H}_2\text{SO}_4 \text{ after discharging} = 1.139 \times \frac{20}{100} \times 3500 = 797.3 \text{ g}$$

$$\text{Normality of } \text{H}_2\text{SO}_4 \text{ before discharging} = \frac{1766.31}{98} \times \frac{1}{3.5} = 5.15 \text{ N}$$

$$\text{Normality of } \text{H}_2\text{SO}_4 \text{ after discharging} = \frac{797.3}{98} \times \frac{1}{3.5} = 2.32 \text{ N}$$

**6.(C)** We know

$$\text{Ampere hour} = \frac{\text{Coulomb}}{3600\text{s}} = \frac{954178}{3600} = 265.05$$

**7.(A)** The weight of  $\text{H}_2\text{SO}_4$  lost during discharge =  $1766.31 - 797.3 = 969.01 \text{ g}$

$$\text{Moles of } \text{H}_2\text{SO}_4 \text{ lost during discharge} = \frac{969.01}{98} = 9.88786$$

During the discharge reaction, 4 mol of  $\text{H}^+$  ions, that is, 2 mol of  $\text{H}_2\text{SO}_4$  require 2F,  $2 \times 96500$  coulombs.

$$\text{Therefore, } 9.88786 \text{ mol will require} = \frac{2 \times 96500 \times 9.88786}{2} = 954178 \text{ C}$$

**8.(B)** The reaction involved in discharge process at anode is



**9.(C)** Given that  $R = 1000 \Omega$ ,  $\Lambda_m^0(\text{Cl}^-) = 80 \text{ S cm}^2 \text{ mol}^{-1}$ ;  $\Lambda_m^0(\text{SO}_4^{2-}) = 160 \text{ S cm}^2 \text{ mol}^{-1}$ . Applying the equation

$\Lambda_m^0 = \Lambda_m + b\sqrt{C}$ , we get

$$\Lambda_{m(\text{NaCl})}^0 = 107 + b\sqrt{4 \times 10^{-4}} \tag{1}$$

$$\Lambda_{m(\text{NaCl})}^0 = 97 + b\sqrt{9 \times 10^{-4}} \tag{2}$$

Subtracting Eq. (2) from Eq. (1), we get

$$0 = 10 + b(2 \times 10^{-2}) - b(3 \times 10^{-2})$$



Solving, we get  $b = 10^3$ . Hence,  $\Lambda_{m(\text{NaCl})}^0 = 107 + 10^3 \sqrt{4 \times 10^{-4}} = 107 + 20 = 127 \text{ S cm}^2 \text{ mol}^{-1}$

**10.(D)** We have  $\Lambda_m^0 = \Lambda_m - b\sqrt{C} = 127 - 10^3 \sqrt{25 \times 10^{-4}} = 127 - 10^3 \times 5 \times 10^{-2} = 77$

$$\Lambda_m^0 = \kappa \times C = 77 = K \times V = K \times \frac{1000}{25 \times 10^{-4}}$$

Hence,  $\kappa = 77 \times 25 \times 10^{-7} = 1925 \times 10^{-7}$

Now,  $\kappa = \frac{1}{R} \times \frac{l}{A} \Rightarrow \frac{l}{A} = 1925 \times 10^{-7} \times 1000 = 0.1925 \text{ cm}^{-1}$

**11.(D)** We know that  $N = M \times 2 \Rightarrow M = \frac{5 \times 10^{-3}}{2}$ . We found that  $\frac{l}{A} = 0.1925 \text{ cm}^{-1}$

$$\begin{aligned} \Lambda &= \kappa \times V = \kappa \times \frac{1000 \times 2}{5 \times 10^{-3}} = \frac{1}{R} \times \frac{l}{A} \times \frac{1000 \times 2}{5 \times 10^{-3}} \\ &= \frac{1}{400} \times 0.1925 \times \frac{1000 \times 2}{5 \times 10^{-3}} \\ &= 0.009625 \times 10^4 = 192.5 \text{ S cm}^2 \text{ mol}^{-1} \end{aligned}$$

## <H2> Integer Answer Type

1.(1.525) For the half-cell reaction,

$$E_{\text{cell}} = E_{\text{red}}^{\circ} - \frac{0.059}{n} \log[\text{H}^+]^4 \quad (1)$$

Substituting the values in Eq. (1), we get

$$E_{\text{cell}} = 1.23 - \frac{0.059}{4} \log[\text{H}^+]^4$$

$$\Rightarrow E_{\text{cell}} = 1.23 + \frac{0.059 \times 4}{4} \text{ pH}$$

$$\Rightarrow E_{\text{cell}} = 1.23 + 0.059 \times 5 \quad [\text{Given: pH} = 5]$$

$$\Rightarrow E_{\text{cell}} = 1.525 \text{ V}$$

2.(2.15) For the given electrochemical cell



$$E_{\text{cell}}^{\circ} = E_{\text{Pb}^{2+}/\text{Pb}}^{\circ} - E_{\text{Sn}^{2+}/\text{Sn}}^{\circ}$$

$$= -0.13 - (-0.14) = 0.01 \text{ V}$$

From the Nernst equation for given electrochemical cell

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{2.303RT}{nF} \log \left[ \frac{\text{Sn}^{2+}}{\text{Pb}^{2+}} \right] \quad (1)$$

Substituting the given values in Eq. (1), we get

$$0 = 0.01 - \frac{0.06}{2} \log \left[ \frac{\text{Sn}^{2+}}{\text{Pb}^{2+}} \right] \quad [\text{At Equilibrium, } E_{\text{cell}} = 0]$$

$$\Rightarrow 0.01 = 0.03 \log \left[ \frac{\text{Sn}^{2+}}{\text{Pb}^{2+}} \right]$$

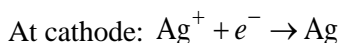
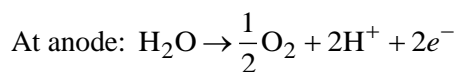
$$\Rightarrow \frac{1}{3} = \log \left[ \frac{\text{Sn}^{2+}}{\text{Pb}^{2+}} \right] \quad (2)$$

Taking antilog of Eq. (2), we get

$$\left[ \frac{\text{Sn}^{2+}}{\text{Pb}^{2+}} \right] = 10^{1/3} = 2.15$$

3.(5.67) The moles of  $\text{AgNO}_3$  deposited at cathode

$$n_{\text{AgNO}_3} = \frac{m}{M} = \frac{108}{108} = 1 \text{ mol}$$



Thus, 1 F required to deposit 1 mol of Ag and 2 F required to produce 0.5 mol of oxygen gas.

$\Rightarrow$  1 F will produce 0.25 mol of oxygen gas.

Thus, the volume of oxygen gas produced at 273 K and 1 bar pressure can be calculated using ideal gas equation.

$$\Rightarrow pV = nRT$$

$$V = \frac{0.25 \times 0.08314 \times 273}{1} = 5.67 \text{ L}$$

4.(2) The original emf is  $E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{0.1}{0.5}$

Now, changed or new emf is

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{0.01}{0.25}$$

Solving only the log terms, we get

In the first case  $\log 0.1/0.5 = -0.6989$

In the second case  $\log 0.01/0.25 = -1.3979$

Taking their ratios, we get  $1.3979/0.6989 = 2$ .

5.(6) At cathode:  $\text{Mg}^{2+} + 2e^- \rightarrow \text{Mg}$

At anode:  $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2e^-$

Therefore, the equivalent of Mg at cathode = Equivalent of  $\text{Cl}_2$  at anode

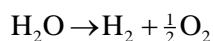
$$\text{Therefore, } \frac{6.5}{24.3/2} = \frac{W_{\text{Cl}_2}}{35.5} \Rightarrow W_{\text{Cl}_2} = 18.99 \text{ g}$$

At NTP,  $pV = \frac{W}{M}RT$ . Substituting values, we get

$$1 \times V = \frac{18.99}{71} \times 0.0821 \times 273$$

Solving, we get volume of  $\text{Cl}_2 = 5.99 \text{ L}$ .

6.(8) When water is electrolyzed, gram equivalent mass is liberated at cathode and anode.



Mass of  $\text{H}_2$  formed = 1.008 g

$$\text{Number of moles of } \text{H}_2 = \frac{1.008}{2} = 0.504$$

Mass of  $\text{O}_2$  formed =  $x$  g

Number of moles of  $\text{O}_2 = x/32$

According to the law of conservation of mass,

Number of moles of  $\text{H}_2 \times 2 =$  Number of moles of  $\text{O}_2$

$$0.504 = \frac{1}{2} \left( \frac{x}{32} \right) \Rightarrow x = \frac{32.256}{4} = 8.064 \text{ g}$$

$$7.(3) E_{\text{cell}}^{\circ} = E_{\text{Fe}^{2+}/\text{Fe}^{3+}}^{\circ} + E_{\text{Ag}^+/\text{Ag}}^{\circ} = -0.771 + 0.799 = 0.028 \text{ V}$$

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

$$0 = E_{\text{cell}}^{\circ} - \frac{0.0591}{1} \log \frac{[\text{Fe}^{3+}]}{[\text{Fe}^{2+}][\text{Ag}^+]}$$
$$E_{\text{cell}}^{\circ} = 0.0591 \log \frac{1}{[\text{Ag}^+]} \Rightarrow [\text{Ag}^+] = 0.34 \text{ M}$$

$$\text{Now, } \log K = \frac{nE^{\circ}}{0.0591} \Rightarrow K = 3.0$$

8.(2) 88 g Hg has 12 g Cd. Therefore, 2.56 g Hg require =  $\frac{12 \times 2}{88}$  g Cd = 0.349

Since  $\text{Cd}^{2+} + 2e^- \rightarrow \text{Cd}$ , hence the equivalent weight of Cd is  $E_{\text{Cd}} = \frac{112.40}{2}$

Now, 
$$W = \frac{Eit}{96500}$$

$$\Rightarrow 0.349 = \frac{112.4 \times 5 \times t}{2 \times 96500}$$

$$\Rightarrow t = 120 \text{ s} = 2 \text{ min}$$

9.(7) We know

$$E_{\text{cell}}^{\circ} = \frac{0.0591}{2} \log K_C$$

$$\Rightarrow K_C = \text{antilog} \left( \frac{2 \times 0.235}{0.0591} \right) = 8.966 \times 10^7$$

10.(4)  $\text{Ni} \rightarrow \text{Ni}^{2+} + 2e^-$ ;  $E_{\text{ox}}^{\circ} = 0.236 \text{ V}$

$2\text{H}^+ + 2e^- \rightarrow \text{H}_2$ ;  $E_{\text{red}}^{\circ} = 0$

Therefore,  $E_{\text{cell}} = 0$  so  $E_{\text{cell}}^{\circ} + \frac{0.059}{2} \log_{10} \frac{[\text{H}^+]^2}{[\text{Ni}^{2+}]} = 0 = 0.236 + \frac{0.059}{2} \log_{10} [\text{H}^+]^2$

Solving, we get  $-\log \text{H}^+ = 4$  therefore  $\text{pH} = 4$ .

### <H2> Matrix–Match Type

1.(A)  $\rightarrow$  (s); (B)  $\rightarrow$  (p); (C)  $\rightarrow$  (q, r); (D)  $\rightarrow$  (q)

Concept based.

2.(A)  $\rightarrow$  (r); (B)  $\rightarrow$  (q); (C)  $\rightarrow$  (p); (D)  $\rightarrow$  (s)

Concept based.

3.(A)  $\rightarrow$  (q); (B)  $\rightarrow$  (q); (C)  $\rightarrow$  (r, s); (D)  $\rightarrow$  (p, q)

Concept based.

4.(A)  $\rightarrow$  (p, r); (B)  $\rightarrow$  (s); (C)  $\rightarrow$  (p); (D)  $\rightarrow$  (t, q)

(A)  $E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \frac{[\text{Ag}^+]^2}{[\text{Cd}^{2+}]} = -1.20 - \frac{0.0591}{2} (-1) = -1.17 \text{ V}$$

As  $E_{\text{cell}}$  is negative,  $\Delta G^{\circ} = -nFE^{\circ}$  is positive, so the reaction should be non-spontaneous.

(B)  $[\text{Ag}^+]^2/[\text{Cd}^{2+}] = (0.1)^2/0.1 = 1/10$

(C)  $E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$

(D)  $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = -0.40 - 0.80 = -1.20 \text{ V}$ .

