

# electrochemistry-CBSE(Solutions)

## Explanations

1. The galvanic cell for the given reaction is



2. Refer to text on page 30. (1)

3. Electrochemical cell is a device that converts the chemical energy of a spontaneous redox reaction into electrical energy. (1)

If the external applied potential becomes greater than  $E_{\text{cell}}^{\circ}$  of electrochemical cell, the reaction gets reversed and electrochemical cell now functions as an electrolytic cell. (1)

4. Given  $E_{\text{Fe}^{2+}/\text{Fe}}^{\circ} = -0.44\text{V}$ ,  $E_{\text{X}^{2+}/\text{X}}^{\circ} = -2.36\text{V}$

$$E_{\text{Y}^{2+}/\text{Y}}^{\circ} = -0.14\text{V}$$

'X' is a better for coating the surface of iron because standard oxidation potential of X is greater than that of Fe so before oxidation of iron takes place, oxidation of X occurs. (2)

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

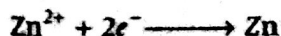
$$E^{\circ} = +1.1\text{V}, \Delta G^{\circ} = ?$$

$$1F = 96500\text{C mol}^{-1}, n = 2$$

$$\Delta G^{\circ} = -nFE^{\circ} = -2 \times 96500 \times 1.1 \quad (1)$$

$$= -212300\text{J mol}^{-1} = -212.3\text{kJ mol}^{-1} \quad (1)$$

6. The electrode reaction involved is



Thus,  $n = 2$

By using Nernst equation, we get,

$$E_{\text{Zn}^{2+}/\text{Zn}} = E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} - \frac{0.0591}{2} \log \frac{1}{[\text{Zn}^{2+}]} \quad (1)$$

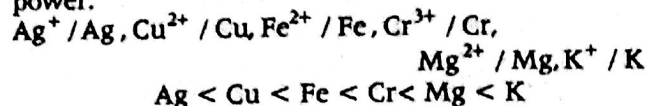
Give that 0.1 M  $\text{ZnSO}_4$  solution is 95% dissociated that means

$$[\text{Zn}^{2+}] = \frac{95}{100} \times 0.1 = 0.095\text{M}$$

$$\begin{aligned} \therefore E_{\text{Zn}^{2+}/\text{Zn}} &= -0.76 - \frac{0.0591}{2} \log \frac{1}{0.095} \\ &= -0.76 - 0.0295[\log 1000 - \log 95] \\ &= -0.76 - 0.0295 \times 1.022 \\ & \quad [\because \log 1000 = 3, \log 95 = 1.978] \\ &= -0.79015\text{V} \quad (1) \end{aligned}$$

7. Reducing power  $\propto \frac{1}{\text{Reduction potential}}$ , i.e. as the reduction potential decreases, reducing power increases. So, arrange the given metals on this basis.

Metals in an increasing order of their reducing power.



(More negative  $E^{\circ}$  means that the redox couple is a stronger reducing agent). (2)

8. Given,  $E_{\text{Cd}^{2+}/\text{Cd}}^{\circ} = -0.403\text{V}$

$$E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} = -0.763\text{V}$$

$$\therefore E_{\text{Cell}}^{\circ} = -0.403 - (-0.763)$$

$$E_{\text{Cell}}^{\circ} = 0.36\text{V}$$

$$\therefore \Delta_r G^{\circ} = -nFE_{\text{Cell}}^{\circ} \quad (1)$$

where  $n$  = number of moles of electrons used

$$n = 2 \text{ (Here)} \Rightarrow F = 96500\text{C mol}^{-1}$$

$$\therefore \Delta_r G^{\circ} = -2 \times 96500 \times 0.36 = 69480\text{J/mol} \quad (1)$$

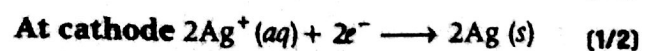
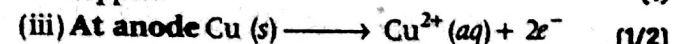
$$\text{Also, } E_{\text{Cell}}^{\circ} = \frac{0.0591}{n} \log K_c$$

$$\therefore \log K_c = \frac{n \times E_{\text{Cell}}^{\circ}}{0.0591} = \frac{2 \times 0.36}{0.0591} = 12.18$$

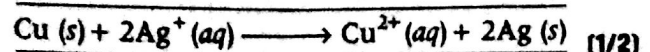
$$\log K_c = 12.18 \quad (1)$$

9. (i)  $\text{Cu}/\text{Cu}^{2+}(aq, 1\text{M}) || \text{Ag}^{+}(aq, 1\text{M}) | \text{Ag}$  (1)

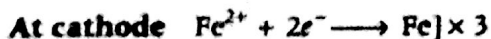
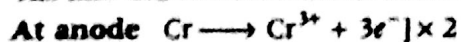
(ii) Direction of flow of current is from silver to copper. (1)



Net reaction



10. Since, oxidation of Cr is taking place in the given reaction, the chromium electrode is anode and as Fe is reduced in the reaction, Fe electrode is cathode. The half-cell reactions are as follows



Overall reaction



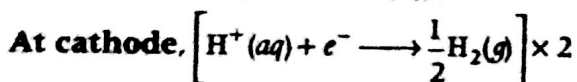
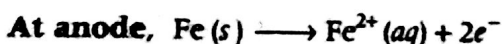
$$E^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ = -0.44 - (-0.74) = 0.3 \text{ V} \quad (1/2)$$

$$E = E^\circ - \frac{0.0591}{n} \log \frac{[\text{Cr}^{3+}]^2}{[\text{Fe}^{2+}]^3} \quad (1/2)$$

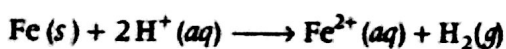
Here,  $n$  = number of electrons transferred, i.e. equal to 6.

$$\begin{aligned} &= 0.3 - \frac{0.0591}{6} \log \frac{(0.01)^2}{(0.1)^3} \\ &= 0.309 \approx 0.31 \text{ V} \quad (1) \end{aligned}$$

11. Write Nernst equation and find the number of electrons to find  $E_{\text{cell}}$ .



Overall reaction



Given,  $E_{\text{Fe}^{2+}/\text{Fe}}^\circ = -0.44 \text{ V}$ ,

$$E_{\text{H}^+/\text{H}_2}^\circ = 0.00 \text{ V}$$

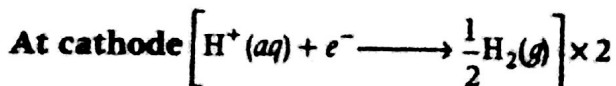
$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0591}{2} \log \frac{[\text{Fe}^{2+}]}{[\text{H}^+]^2} \quad (1)$$

$$= [0.00 - (-0.44)] - \frac{0.0591}{2} \log \frac{10^{-3}}{(10^{-2})^2} \quad (1)$$

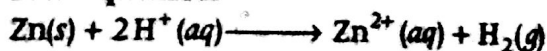
$$= 0.44 - \frac{0.0591}{2}$$

$$E_{\text{cell}} = 0.4104 \text{ V} \quad (1)$$

12. At anode  $\text{Zn}(s) \longrightarrow \text{Zn}^{2+}(aq) + 2e^-$



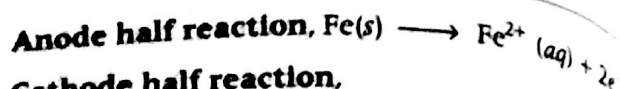
Net equation



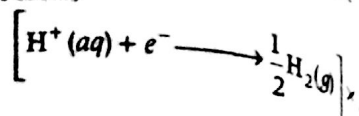
$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0591}{n} \log \frac{[\text{Zn}^{2+}]}{[\text{H}^+]^2}$$

$$= 0.76 - \frac{0.0591}{2} = 0.731 \text{ V} \quad (2)$$

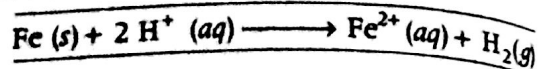
13. Write Nernst equation and find the number of electrons to find  $E_{\text{cell}}$ .



Cathode half reaction,



Net reaction



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

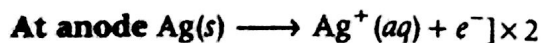
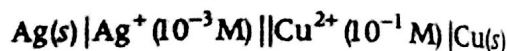
$$E_{\text{cell}} = E_{\text{cell}}^\circ - \left( \frac{0.05912}{n} \right) \log \frac{[\text{Fe}^{2+}]}{[\text{H}^+]^2}$$

$$= 0.44 - \frac{0.05912}{2} \log \frac{0.001}{1}$$

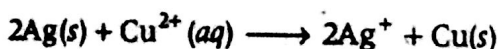
$$\begin{aligned} \therefore E_{\text{cell}} &= 0.44 - [0.02956 \times (-3)] \\ &= 0.44 + 0.08868 \\ &= 0.52868 \text{ V} \approx 0.53 \text{ V} \end{aligned}$$

14. Write Nernst equation and find the number of electrons to find  $E_{\text{cell}}$ .

As per the condition given,



Overall reaction



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

$$= 0.46 - \frac{0.0591}{2} \log \frac{(10^{-3})^2}{10^{-1}}$$

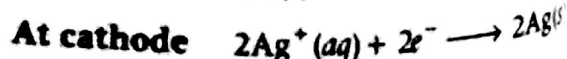
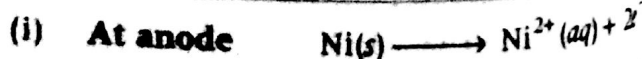
$$= 0.46 - \frac{0.0591}{2} \log 10^{-5}$$

$$E_{\text{cell}} = 0.608 \text{ V}$$

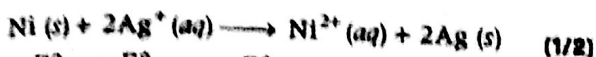
15. (i) Calculate  $E_{\text{cell}}^\circ$  by using the formula,

$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ$$

(ii) Calculate the number of electrons by writing electrode reactions then write Nernst equation and find  $E_{\text{cell}}$ .



**Overall reaction**



$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} \\ = +0.80 \text{ V} - (-0.25 \text{ V}) = 1.05 \text{ V} \quad (1)$$

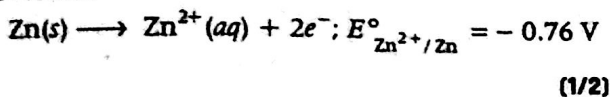
$$(ii) E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{[\text{Ni}^{2+}]}{[\text{Ag}^+]^2} \quad (1/2) \\ = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{(0.1)}{(1)^2} \\ = 1.05 - \frac{0.0591}{2} \times -1 \\ = 1.05 + 0.0295 = 1.0795 \text{ V} \quad (1)$$

16. (i) Calculate  $E_{\text{cell}}^{\circ}$  by using the formula,  
 $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$   
 (ii) Calculate  $\Delta_r G^{\circ}$  by using the formula,  
 $\Delta_r G^{\circ} = -nFE_{\text{cell}}^{\circ}$

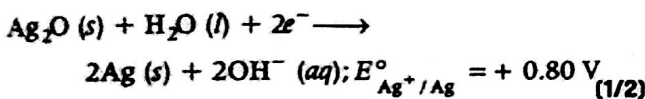
The given reaction, shows that zinc is oxidised and silver is reduced in the button cell.

The reactions occurring are :

**At anode**



**At cathode**



$$\text{Now, } E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} \quad (1/2) \\ = E_{\text{Ag}^+/\text{Ag}}^{\circ} - E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} \\ = 0.80 - (-0.76) = 1.56 \text{ V} \quad (1/2)$$

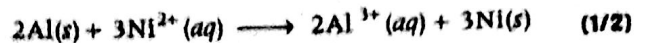
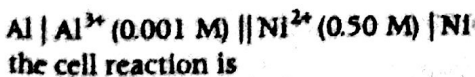
$$\text{We know that, } \Delta_r G^{\circ} = -nE_{\text{cell}}^{\circ} F \quad (1/2)$$

$$\text{Here, } n = 2, F = 96500 \text{ C mol}^{-1}$$

$$\Delta_r G^{\circ} = -2 \times 1.56 \times 96500 \text{ J mol}^{-1} \\ = -301080 \text{ J mol}^{-1} = -301 \text{ kJ mol}^{-1} \quad (1/2)$$

17. (i) Calculate  $E_{\text{cell}}^{\circ}$  by using the formula,  
 $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$   
 (ii) Calculate the number of electrons by writing the electrode reactions then write Nernst equation and find  $E_{\text{cell}}$ .

For the cells,



Thus,  $n = 6$

$$E_{\text{cell}}^{\circ} = E_{\text{Ni}^{2+}/\text{Ni}}^{\circ} - E_{\text{Al}^{3+}/\text{Al}}^{\circ} \\ = -0.25 \text{ V} - (-1.66 \text{ V}) = 1.41 \text{ V} \quad (1)$$

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log \frac{[\text{Al}^{3+}]^2}{[\text{Ni}^{2+}]^3} \quad (1/2)$$

$$E_{\text{cell}} = 1.41 \text{ V} - \frac{0.0591 \text{ V}}{6} \log \frac{(10^{-3})^2}{(0.50)^3} \\ = 1.41 \text{ V} - \frac{0.0591 \text{ V}}{6} \log \frac{8 \times 10^{-6}}{1} \left( \because \frac{1}{0.125} = 8 \right)$$

$$E_{\text{cell}} = 1.41 \text{ V} - \frac{0.0591 \text{ V}}{6} \times (-5.0969) \\ = 1.41 \text{ V} + 0.0502 \text{ V} = 1.4602 \text{ V} \quad (1)$$

18. (i) Find  $\Delta G^{\circ}$  using the formula,  
 $\Delta G^{\circ} = -nFE^{\circ}$   
 (ii) Find  $\log K_C$  by using the formula,  
 $-nFE^{\circ} = -2.303 RT \log K_C$   
 (iii) Find  $K_C$  using the antilog table.

$$\text{Here, } n = 2$$

$$\Delta G^{\circ} = -nFE^{\circ} \quad (1/2)$$

$$\Delta G^{\circ} = -2 \times 96500 \text{ C} \times 1.05 \text{ V} \\ \Delta G^{\circ} = -202.65 \text{ kJ mol}^{-1} \quad (1/2)$$

$$-nFE^{\circ} = -2.303 RT \log K_C \quad (1/2)$$

$$[\because \Delta G^{\circ} = -2.303 RT \log K_C]$$

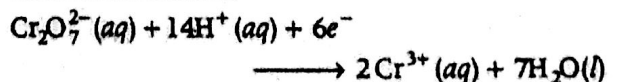
$$\log K_C = \frac{nE^{\circ}}{0.0591} \\ = \frac{2 \times 1.05}{0.0591} = 35.5329 \quad (1/2)$$

$$K_C = \text{antilog } 35.5329$$

$$K_C = 3.412 \times 10^{35} \quad (1)$$

19. Given,  $[\text{Cr}_2\text{O}_7^{2-}] = 0.10 \text{ M}$ ,  $[\text{Cr}^{3+}] = 0.20 \text{ M}$  and  
 $[\text{H}^+] = 1 \times 10^{-4} \text{ M}$ ,  $E_{\text{cell}}^{\circ} = 1.33 \text{ V}$ ,  $n = 6$

Half-cell reaction,



By using Nernst equation, we get

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591 \text{ V}}{n} \log \frac{[\text{Cr}^{3+}]^2}{[\text{Cr}_2\text{O}_7^{2-}] [\text{H}^+]^{14}} \quad (1) \\ = 1.33 \text{ V} - \frac{0.0591 \text{ V}}{6} \log \frac{(0.2)^2}{(0.1) (1.0 \times 10^{-4})^{14}} \quad (1)$$

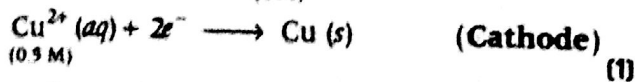
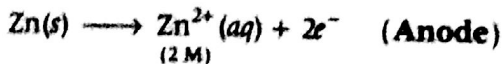
$$\begin{aligned}
 &= 1.33 \text{ V} - \frac{0.0591 \text{ V}}{6} \log \frac{4 \times 10^{-2}}{1.0 \times 10^{-37}} \\
 &= 1.33 \text{ V} - \frac{0.0591 \text{ V}}{6} \log 4 \times 10^{35} \\
 &= 1.33 \text{ V} - \frac{0.0591 \text{ V}}{6} \times 55.6021 \\
 &= 1.33 \text{ V} - 0.5477 \text{ V} = 0.7823 \text{ V} \quad (1)
 \end{aligned}$$

20. (i) Calculate  $E_{\text{cell}}^{\circ}$  by using the formula,

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$$

- (ii) Calculate the number of electrons by writing the electrode reactions then write Nernst equation and find  $E_{\text{cell}}$ .

- (i) The cell reactions are as



(ii)  $E_{\text{cell}}^{\circ} = E_{\text{right}}^{\circ} - E_{\text{left}}^{\circ}$

$$E_{\text{cell}}^{\circ} = 0.34 \text{ V} - (-0.76 \text{ V}) = 1.10 \text{ V} \quad (1/2)$$

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} \quad (1/2)$$

$$\begin{aligned}
 E_{\text{cell}} &= 1.10 \text{ V} - \frac{0.0591 \text{ V}}{2} \log \frac{2}{0.5} \\
 &= 1.10 \text{ V} - \frac{0.0591 \text{ V}}{2} \times 0.6021 \\
 &= 1.10 \text{ V} - 0.0178 \text{ V} = 1.0822 \text{ V} \quad (1)
 \end{aligned}$$

21. (i) Find  $\log K_C$  using the formula,

$$\Delta G^{\circ} = -2.303 \log K_C$$

- (ii) Find  $K_C$  using the antilog table.

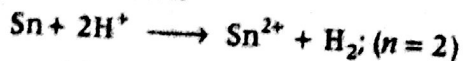
$$\Delta G^{\circ} = -2.303 RT \log K_C \quad (1)$$

$$-212300 = -2.303 \times 8.314 \times 298 \times \log K_C$$

$$\log K_C = \frac{212300}{2.303 \times 8.314 \times 298} = 37.2074 \quad (1)$$

$$K_C = \text{antilog } 37.2074, K_C = 1.6 \times 10^{37} \quad (1)$$

22. Cell reaction is



$$E_{\text{cell}}^{\circ} = E_R - E_L = 0 - (-0.144) = 0.144 \text{ V} \quad (1/2)$$

According to Nernst equation,

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591 \text{ V}}{n} \log \frac{[\text{Sn}^{2+}]}{[\text{H}^{+}]^2} \quad (1/2)$$

$$E_{\text{cell}} = 0.144 \text{ V} - \frac{0.0591 \text{ V}}{2} \log \frac{(0.05)}{(0.02)^2} \quad (1)$$

$$= 0.144 \text{ V} - \frac{0.0591 \text{ V} \times 2.0969}{2} = 0.078 \text{ V}$$

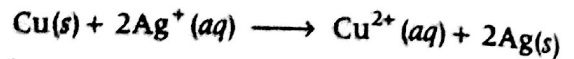
23. (i) Write the cell reaction for the given cell.

- (ii) Calculate  $E_{\text{cell}}^{\circ}$  by using the formula,

$$E_{\text{cell}}^{\circ} = E_{\text{right}}^{\circ} - E_{\text{left}}^{\circ}$$

- (iii) Calculate  $\text{Ag}^{+}$  ion concentration by using Nernst equation.

Cell reaction is



$$E_{\text{cell}}^{\circ} = E_R - E_L = +0.80 - (+0.34) = 0.46 \text{ V}$$

Number of electrons taking part,  $n = 2$

By using Nernst equation,

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591 \text{ V}}{n} \log \frac{[\text{Cu}^{2+}]}{[\text{Ag}^{+}]^2}$$

$$0.422 \text{ V} = 0.46 \text{ V} - \frac{0.0591}{2} \log \frac{[\text{Cu}^{2+}]}{[\text{Ag}^{+}]^2}$$

$$0.422 \text{ V} = 0.46 \text{ V} - 0.0295 \log \frac{[\text{Cu}^{2+}]}{[\text{Ag}^{+}]^2}$$

$$\log \frac{[\text{Cu}^{2+}]}{[\text{Ag}^{+}]^2} = \frac{0.46 - 0.422}{0.0295} = 1.288$$

$$\frac{[\text{Cu}^{2+}]}{[\text{Ag}^{+}]^2} = \text{antilog } 1.288$$

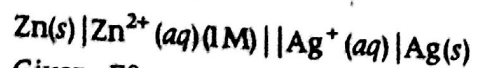
$$\frac{[\text{Cu}^{2+}]}{[\text{Ag}^{+}]^2} = 19.41$$

$$\frac{0.10}{[\text{Ag}^{+}]^2} = 19.41$$

$$[\text{Ag}^{+}]^2 = 0.00515$$

$$[\text{Ag}^{+}] = 0.0717 = 7.17 \times 10^{-2} \text{ M}$$

24. According to given information, cell can be represented as



Given,  $E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} = -0.763 \text{ V}$ ,  $E_{\text{Ag}^{+}/\text{Ag}}^{\circ} = +0.80 \text{ V}$

$$\therefore E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$$

$$= 0.80 - (-0.76) = 1.56 \text{ V}$$

By using Nernst equation, we get

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{Ag}^{+}]^2}$$

$$1.48 = 1.56 - \frac{0.0591}{2} \log \frac{1}{[\text{Ag}^{+}]^2}$$

$$0.08 = +0.0295 \log \frac{1}{[\text{Ag}^{+}]^2}$$

$$\frac{0.08}{0.0295} = -2\log[\text{Ag}^+]$$

$$-\frac{0.08}{0.0295 \times 2} = \log[\text{Ag}^+]$$

$$\log[\text{Ag}^+] = -1.356$$

$$[\text{Ag}^+] = \text{Antilog}[-1.356]$$

$$= 0.0441 = 4.4 \times 10^{-2} \text{ M}$$

25.  $E^\circ_{\text{cell}} = 1.05 \text{ V}, n = 2$

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

$$E_{\text{cell}} = 1.05 - \frac{0.0591}{2} \log \frac{(0.160)}{(0.002)^2}$$

$$E_{\text{cell}} = 1.05 - 0.0295 \log 4 \times 10^4$$

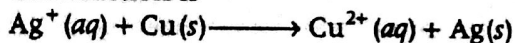
$$E_{\text{cell}} = 1.05 - 0.0295 \times 4.6021$$

$$= 0.91 \text{ V}$$

26. Given,  $E^\circ_{\text{cell}} = 0.46 \text{ V}, [\text{Ag}^+] = 0.001 \text{ M},$

$$[\text{Cu}^{2+}] = 0.10 \text{ M}$$

Cell reaction is



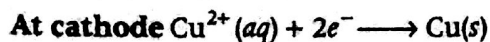
Thus,  $n = 2$

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

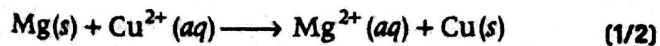
$$= 0.46 - \frac{0.0591}{2} \log \frac{0.10}{(10^{-3})^2}$$

$$= 0.46 - 0.148 = 0.312 \text{ V}$$

27. At anode  $\text{Mg}(s) \longrightarrow \text{Mg}^{2+}(aq) + 2e^-$



Overall reaction



$$E^\circ_{\text{cell}} = 0.34 - (-2.37) = + 2.71 \text{ V} \quad (1/2)$$

(i)  $\therefore n = 2$

$$\therefore E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{2} \log \frac{0.001}{0.0001} \quad (1)$$

$$E_{\text{cell}} = 2.71 - 0.0295 \log 10$$

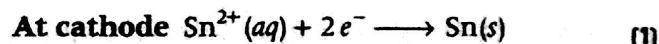
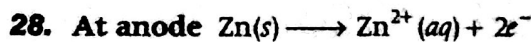
$$E_{\text{cell}} = 2.71 - 0.0295 \times 1$$

$$E_{\text{cell}} = 2.68 = 2.70 \text{ V} \quad (1)$$

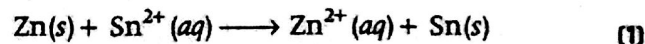
(ii)  $\Delta G = -nFE_{\text{cell}}$

$$\Delta G = -2 \times 96500 \times 2.70 \quad (\because 1F = 96500 \text{ C mol}^{-1})$$

$$\Delta G = 521.10 \text{ kJ mol}^{-1} \quad (2)$$



Overall reaction



$$E^\circ_{\text{cell}} = E^\circ_{\text{Sn}^{2+}/\text{Sn}} - (-E^\circ_{\text{Zn}^{2+}/\text{Zn}})$$

$$= -0.14 + 0.76$$

$$= 0.62 \text{ V} \quad (1)$$

$$\Delta G^\circ = -nFE^\circ$$

$$= -2 \times 96500 \times 0.62 \quad (1)$$

$$\Delta G^\circ = -119660 \text{ J mol}^{-1} \quad (1)$$

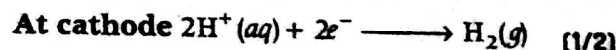
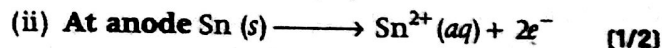
29.

(i) Calculate  $E^\circ_{\text{cell}}$  by using the formula,

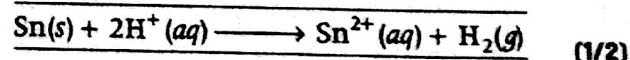
$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

(ii) Calculate the number of electrons by writing the electrode reactions, then write Nernst equation and find  $E_{\text{cell}}$ .

(i) Refer to solution 9 (iii)



Net reaction



$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{2} \log \frac{[\text{Sn}^{2+}]}{[\text{H}^+]^2} \quad (1)$$

$$= (E^\circ_{\text{H}^+/\text{H}_2} - E^\circ_{\text{Sn}^{2+}/\text{Sn}}) - \frac{0.0591}{2} \log \frac{0.04}{(0.02)^2}$$

$$= [0\text{V} - (-0.14\text{V})] - \frac{0.0591}{2} \log \frac{4}{100} \times \frac{100}{2} \times \frac{100}{2}$$

$$= +0.14\text{V} - 0.0591\text{V} = 0.0809 \text{ V} \quad (1)$$

## ☑ Explanations

### 1. Conductivity,

$$\kappa = \frac{1}{R} \times \frac{l}{A} \quad \left( \text{As } \kappa = \frac{1}{\rho} \text{ and } \rho = \frac{RA}{l} \right)$$

where,  $\kappa$  = conductivity,  $R$  = resistance

and  $\frac{l}{A}$  = cell constant

(1)

**2. Limiting molar conductivity** The molar conductivity of a solution at infinite dilution is called limiting molar conductivity and is represented by the symbol  $\Lambda_m^\circ$ .

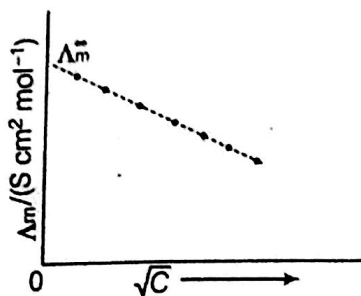
(1)

**3. Strong electrolytes** are completely ionised in all concentrations. For strong electrolytes,  $\Lambda_m$  increases slowly with dilution. It is found to vary with concentration according to the Debye-Huckel-Onsager equation.

$$\Lambda_m = \Lambda_m^\circ - [A + B\Lambda_m^\circ] \sqrt{C} \quad (1)$$

where,  $A$  and  $B$  are called Debye-Huckel constants.

If solution is diluted (concentration is decreased), there is decrease in ionic attractions hence, molar conductivity increases with decrease in concentration.



4. To find out the degree of dissociation of acetic acid, we first need to calculate the  $\Lambda_m^\circ(\text{CH}_3\text{COOH})$  with the help of Kohlrausch law.

According to the question

$$\text{Given, } \Lambda_m = 39.05 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\lambda^\circ(\text{H}^+) = 349.6 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\lambda^\circ(\text{CH}_3\text{COO}^-) = 40.9 \text{ S cm}^2 \text{ mol}^{-1} \quad (1)$$

$$\begin{aligned} \Lambda_m^\circ(\text{CH}_3\text{COOH}) &= \lambda^\circ(\text{H}^+) + \lambda^\circ(\text{CH}_3\text{COO}^-) \\ &= 349.6 \text{ S cm}^2 \text{ mol}^{-1} + 40.9 \text{ S cm}^2 \text{ mol}^{-1} \\ &= 390.5 \text{ S cm}^2 \text{ mol}^{-1} \end{aligned}$$

$$\alpha = \frac{\Lambda_m}{\Lambda_m^\circ} = \frac{39.05 \text{ S cm}^2 \text{ mol}^{-1}}{390.5 \text{ S cm}^2 \text{ mol}^{-1}}$$

$$\Rightarrow \alpha = 0.1 \quad (1)$$

5. Kohlrausch's law of independent migration of ions states that the limiting molar conductivity of an electrolyte can be expressed as the sum of the individual contributions of the anion and the cation of the electrolyte, e.g.

$$\Lambda_m^\circ(\text{CH}_3\text{COOH}) = \lambda_{\text{CH}_3\text{COO}^-}^\circ + \lambda_{\text{H}^+}^\circ \quad (1)$$

The conductivity of a solution is related with the number of ions present per unit volume of the solution. When the solution is diluted; the number of ions per unit volume decreases. Hence, conductivity or specific conductance of the solution decreases. (1)

6. (i) First we find out molar conductivity at given concentration using the formula,

$$\Lambda_m^c = \frac{\kappa \times 1000}{M}$$

- (ii) Then, we find out the degree of dissociation ( $\alpha$ )

$$\text{using the formula, } \alpha = \frac{\Lambda_m^c}{\Lambda_m^\circ}$$

- (iii) After dissociation of acetic acid into ions, we will find out the dissociation constant using the

$$\text{formula, } K_a = \frac{C\alpha^2}{1-\alpha}$$

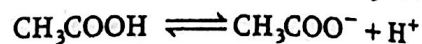
Molar conductivity at 0.001 M,

$$\begin{aligned} \Lambda_m^c &= \frac{\kappa \times 1000}{M} \\ &= \frac{4 \times 10^{-5} \times 1000}{0.001} \text{ S cm}^2 / \text{mol} \\ &= 40 \text{ S cm}^2 / \text{mol} \end{aligned}$$

Degree of dissociation,

$$\alpha = \frac{\Lambda_m^c}{\Lambda_m^\circ} = \frac{40}{390} = 0.10256$$

For the dissociation of acetic acid,  $\text{CH}_3\text{COOH}$



Initial (conc.) $C \text{ mol}^{-1}$	0	0
After time, $t (C - C\alpha)$	$C\alpha$	$C\alpha$

$$\text{Dissociation constant, } K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]}$$

$$= \frac{C\alpha \cdot C\alpha}{C - C\alpha} = \frac{C^2\alpha^2}{C(1-\alpha)} = \frac{C\alpha^2}{(1-\alpha)}$$

$$K_a = \frac{0.001 \times (0.103)^2}{(1 - 0.103)}$$

$$= \frac{1.061 \times 10^{-5}}{0.897}$$

$$= 1.18 \times 10^{-5}$$

7. The resistance,  $R$  of a conductor varies directly with length ( $l$ ) and inversely with area ( $A$ ) of cross-section, i.e.

$$R \propto \frac{l}{A} \quad \text{or} \quad R = \rho \frac{l}{A}$$

where,  $\rho$  = resistivity ;  $\rho = \frac{RA}{l}$

Conductivity ( $\kappa$ ) is the reciprocal of resistivity ( $\rho$ ). So,

$$\kappa = \frac{1}{\rho} = \frac{1}{RA/l} = \frac{l}{RA} = \frac{1}{R} \cdot \left(\frac{l}{A}\right)$$

$$\therefore \kappa = \frac{1}{R} \cdot \left(\frac{l}{A}\right) = \frac{1}{R} \cdot G^*$$

where,  $G^* = \frac{l}{A} = \text{cell constant}$

The molar conductivity of a solution is related to its conductivity by the following relation

$$\Lambda_m = \frac{\kappa \times 1000}{M} \text{ S cm}^2 \text{ mol}^{-1}$$

where,  $M$  = molarity

8. Use formula  $\Lambda_m = \frac{\kappa \times 1000}{M}$  to find the value of  $\kappa$ .

Given  $\Lambda_m = 138.9 \text{ S cm}^2 \text{ mol}^{-1}$

Molarity,  $M = 1.5 \text{ M} = 1.5 \text{ mol/L}$

We know that,  $\Lambda_m = \frac{\kappa \times 1000}{M}$  (1)

$$138.9 \text{ S cm}^2 \text{ mol}^{-1} = \frac{\kappa \times 1000}{1.5 \text{ mol cm}^{-3}}$$

$$\kappa = \frac{138.9 \text{ S cm}^2 \text{ mol}^{-1} \times 1.5 \text{ mol cm}^{-3}}{1000}$$

$$= 138.9 \times 1.5 \times 10^{-3} \text{ S cm}^{-1}$$

$$= 0.208 \text{ S cm}^{-1} \quad (1)$$

9. Refer to solution 5. (2)

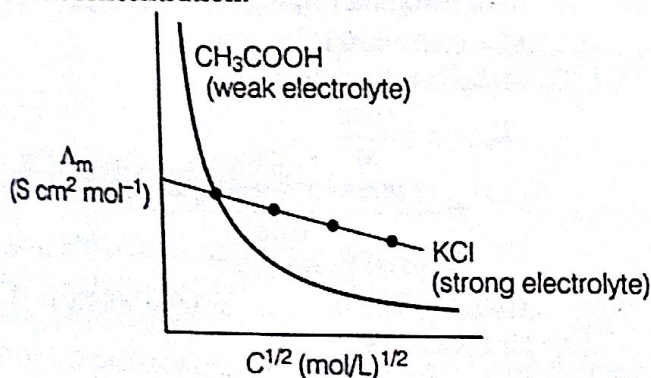
10. **Conductivity** The inverse of resistivity is called conductivity. It is denoted by  $\kappa$  (kappa). It is also known as specific conductance.  $\kappa = 1/\rho$

SI unit of conductivity is  $\text{S m}^{-1}$  or  $\text{ohm}^{-1} \text{ m}^{-1}$ . (1/2)

**Molar conductivity** It is defined as the conductance of the solution which contains one mole of the electrolyte such that entire solution is in between the two electrodes kept one centimetre apart, and large enough to contain all the electrolytes.

Molar conductivity,  $\Lambda_m = \frac{\kappa}{C}$  (1/2)

Variation of conductivity and molar conductivity with concentration.



Conductivity and molar conductivity change with change in concentration of electrolyte.

Conductivity always decreases with decrease in concentration for both weak as well as strong electrolytes. (1/2)

But molar conductivity increases with decrease in concentration. For strong electrolytes,  $\Lambda_m$  increases slowly with dilution but for weak electrolyte,  $\Lambda_m$  increases steeply on dilution, especially near lower concentrations. (1/2)

11. (i) First, we calculate  $\Lambda_m^C$  using the formula

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

(ii) Then, degree of dissociation is calculated using the relation given below.

$$\alpha = \frac{\Lambda_m^C}{\Lambda_m^\infty}$$

Molar conductivity of  $2.5 \times 10^{-4} \text{ M}$  solution is

$$\Lambda_m^C = \frac{\kappa \times 1000}{M} = \frac{5.25 \times 10^{-5} \times 1000}{2.5 \times 10^{-4}}$$

$$\Lambda_m^C = 210 \text{ S cm}^2 \text{ mol}^{-1} \quad (1)$$

$$\Lambda_m^\infty (\text{HCOOH}) = \lambda_{(\text{HCOO}^-)}^\circ + \lambda_{(\text{H}^+)}^\circ$$

$$\Lambda_m^\infty = 50.5 + 349.5$$

$$\Lambda_m^\infty = 400$$

$$\alpha = \frac{\Lambda_m^C}{\Lambda_m^\infty} = \frac{210}{400}$$

$$\alpha = 0.525 \quad (1)$$

12. (i) First find conductivity using the formula conductivity,  $\kappa = 1/R \times l/A$

(ii) Then, find molar conductivity of solution by using the formula,

$$\Lambda_m = \frac{\kappa \times 1000}{\text{Molarity}}$$

Given, Resistance ( $R$ ) =  $200 \Omega$

Molarity of NaCl solution =  $0.01 \text{ M}$

$$\text{Cell constant} = \frac{l}{A} = 1 \text{ cm}^{-1} \quad (1)$$

$$\text{Conductivity } (\kappa) = \frac{1}{R} \times \frac{l}{A} = \frac{1}{200 \Omega} \times 1 \text{ cm}^{-1}$$

$$= 5.0 \times 10^{-3} \Omega^{-1} \text{ cm}^{-1} \quad (1)$$

Molar conductivity ( $\Lambda_m$ )

$$= \frac{\kappa \times 1000}{\text{Molarity}} = \frac{(5.0 \times 10^{-3} \text{ S cm}^{-1}) (1000 \text{ cm}^3 \text{ L}^{-1})}{0.01 \text{ mol L}^{-1}}$$

$$= 500 \text{ S cm}^2 \text{ mol}^{-1} \quad (1)$$

13. Given,  $\Lambda_m^\circ$  of  $\text{Al}_2(\text{SO}_4)_3 = 858 \text{ S cm}^2 \text{ mol}^{-1}$

$$\lambda_{(\text{SO}_4^{2-})}^\circ = 160 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\Lambda_m^\circ [\text{Al}_2(\text{SO}_4)_3] = 2\lambda_{(\text{Al}^{3+})}^\circ + 3\lambda_{(\text{SO}_4^{2-})}^\circ \quad (1)$$

$$858 = 2\lambda_{(\text{Al}^{3+})}^\circ + 3 \times 160$$

$$858 = 2\lambda_{(\text{Al}^{3+})}^\circ + 480$$

$$\lambda_{(\text{Al}^{3+})}^\circ = \frac{858 - 480}{2} = 189 \text{ S cm}^2 \text{ mol}^{-1} \quad (2)$$



14. Given, conductance of cell when electrolyte of 0.1 M KCl is filled = 85 ohm's  
 Conductance of cell when unknown electrolyte of 0.052 M is filled = 96 ohm's  
 Specific conductance ( $\kappa$ ) of 0.1 M KCl =  $1.29 \times 10^{-2} \Omega^{-1} \text{cm}^{-1}$   
 $\kappa = \frac{1}{R} \times \text{cell constant}$   
 Cell constant =  $\kappa \times R$   
 $= 0.0129 \text{ S cm}^{-1} \times 85 = 1.0965 \text{ cm}^{-1}$  (1)  
 $= 109.65 \text{ m}^{-1}$

In case of second solution,

$$\kappa = \frac{1}{R} \times \text{cell constant}$$

$$= \frac{1}{96} \times 109.65 \text{ m}^{-1} = 0.01142 \Omega^{-1} \text{cm}^{-1}$$
 (1)
$$\Lambda_m = \frac{\kappa \times 1000}{M}$$

$$= \frac{0.01142 \Omega^{-1} \text{cm}^{-1} \times 1000 \text{ cm}^3 \text{L}^{-1}}{0.052 \text{ mol L}^{-1}}$$

$$= 219.61 \text{ S cm}^2 \text{mol}^{-1}$$
 (1)

15. (i) Find area from radius by using,  $A = \pi r^2$   
 (ii) Find resistivity from the formula,  $\rho = \frac{RA}{l}$   
 (iii) Find conductivity ( $\kappa$ ) from resistivity as  $\kappa = 1/\rho$   
 (iv) Find molar conductivity,  $\Lambda_m$  by using the formula,  $\Lambda_m = \frac{\kappa \times 1000}{M}$

Given, concentration of NaOH = 0.05 M,  
 diameter = 1 cm, length = 50 cm,  
 resistance =  $5.55 \times 10^3 \Omega$

$$\text{Area, } A = \pi r^2 = 3.14 \times \left(\frac{1}{2}\right)^2 \text{ cm}^2 = 0.785 \text{ cm}^2$$

$$\text{Resistivity, } \rho = \frac{RA}{l}$$

$$= \frac{5.55 \times 10^3 \Omega \times 0.785 \text{ cm}^2}{50 \text{ cm}} = 87.135 \Omega \text{ cm}$$
 (1/2)

Conductivity,

$$\kappa = \frac{1}{\rho} = \frac{1}{87.135} \Omega^{-1} \text{cm}^{-1}$$
 (1/2)
$$= 0.01148 = 1.148 \times 10^{-2} \text{ S cm}^{-1}$$
 (1/2)

Molar conductivity,

$$\Lambda_m = \frac{\kappa \times 1000}{M}$$
 (1/2)
$$= \frac{1.148 \times 10^{-2} \times 1000}{0.05} = 229.6 \text{ S cm}^2 \text{mol}^{-1}$$
 (1/2)

16. Given,  $\Lambda_m(\text{CH}_3\text{COOH}) = 11.7 \text{ S cm}^2 \text{mol}^{-1}$   
 $\Lambda_m^\circ(\text{CH}_3\text{COO}^-) = 40.9 \text{ S cm}^2 \text{mol}^{-1}$   
 $\Lambda_m^\circ(\text{H}^+) = 349.1 \text{ S cm}^2 \text{mol}^{-1}$

$$\text{Degree of dissociation, } \alpha = \frac{\Lambda_m}{\Lambda_m^\circ}$$

$$\Lambda_m^\circ(\text{CH}_3\text{COOH}) = \Lambda_m^\circ(\text{CH}_3\text{COO}^-) + \Lambda_m^\circ(\text{H}^+)$$

$$\Lambda_m^\circ(\text{CH}_3\text{COOH}) = (40.9 + 349.1) = 390 \text{ S cm}^2 \text{mol}^{-1}$$
 (1)
$$\alpha = \frac{\Lambda_m}{\Lambda_m^\circ} = \frac{11.7}{390} = 3.0 \times 10^{-2}$$
 (1)

17. Recall Kohlrausch's law to find the conductance of weak electrolyte.

Given,  $\Lambda_m^\circ(\text{HCl}) = 426 \text{ S cm}^2 \text{mol}^{-1}$ ;

$$\Lambda_m^\circ(\text{NaCl}) = 126 \text{ S cm}^2 \text{mol}^{-1},$$

$$\Lambda_m^\circ(\text{CH}_3\text{COONa}) = 91 \text{ S cm}^2 \text{mol}^{-1}$$

$$\Lambda_m^\circ(\text{CH}_3\text{COOH}) = \Lambda_m^\circ(\text{H}^+) + \Lambda_m^\circ(\text{CH}_3\text{COO}^-)$$
 (1/2)

$$= \Lambda_m^\circ(\text{H}^+) + \Lambda_m^\circ(\text{Cl}^-) + \Lambda_m^\circ(\text{CH}_3\text{COO}^-) + \Lambda_m^\circ(\text{Na}^+) - \Lambda_m^\circ(\text{Cl}^-) - \Lambda_m^\circ(\text{Na}^+)$$

$$= \Lambda_m^\circ(\text{HCl}) + \Lambda_m^\circ(\text{CH}_3\text{COONa}) - \Lambda_m^\circ(\text{NaCl})$$
 (1)

$$= (426 + 91 - 126) \text{ S cm}^2 \text{mol}^{-1}$$
 (1/2)

$$= 391 \text{ S cm}^2 \text{mol}^{-1}$$
 (1)

18. (i) Given: Conductivity of  $\text{CH}_3\text{COOH}$  solution,  
 $\kappa = 3.905 \times 10^{-5} \text{ S cm}^{-1}$

Concentration of  $\text{CH}_3\text{COOH}$  solution,

$$M = 0.001 \text{ mol L}^{-1} = 0.001 \text{ M}$$

Molar conductivity,

$$\lambda_m = \kappa \times \frac{1000}{M}$$

$$= \frac{(3.905 \times 10^{-5} \text{ S cm}^{-1}) \times (1000)}{0.001 \text{ M}}$$

$$= 39.05 \text{ S cm}^2 \text{mol}^{-1}$$
 (1)

Molar conductivity at infinite dilution ( $\Lambda_m^\circ$ ) for  $\text{CH}_3\text{COOH}$ ,

$$\Lambda_m^\circ(\text{CH}_3\text{COOH}) = \Lambda_m^\circ(\text{CH}_3\text{COO}^-) + \Lambda_m^\circ(\text{H}^+)$$

$$= 40.9 + 349.6 \text{ S cm}^2 \text{mol}^{-1}$$

$$= 390.5 \text{ S cm}^2 \text{mol}^{-1}$$

$$\text{Degree of dissociation, } \alpha = \frac{\lambda_m}{\Lambda_m^\circ}$$

$$= \frac{39.05 \text{ S cm}^2 \text{mol}^{-1}}{390.5 \text{ S cm}^2 \text{mol}^{-1}} = 0.1$$
 (2)

- (ii) A device used to convert the chemical energy produced in a spontaneous redox reaction into electrical energy is called an **electrochemical cell**.

When external potential applied becomes greater than  $E_{\text{cell}}^{\circ}$  of electrochemical cell, electrons flow from cathode to anode, i.e. electrochemical cell behaves like an electrolytic cell.

(2)

19. (i) **Degree of dissociation** The ratio of molar conductivity ( $\Lambda_m$ ) at a specific concentration  $C$  to the molar conductivity at infinite dilution, i.e. limiting molar conductivity,  $\Lambda_m^{\circ}$ , is known as degree of dissociation.

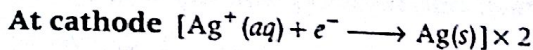
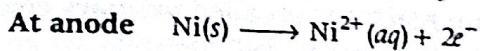
(1)

At any concentration  $C$ , if  $\alpha$  is the degree of dissociation, then for a weak electrolyte,

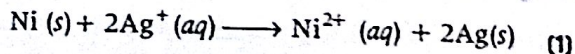
$$\alpha = \frac{\Lambda_m}{\Lambda_m^{\circ}}$$

(1)

- (ii) Given,  $E_{\text{Ni}^{2+}/\text{Ni}}^{\circ} = 0.25 \text{ V}$  and  $E_{\text{Ag}^+/\text{Ag}}^{\circ} = 0.80 \text{ V}$



**Overall reaction**



$$E_{\text{cell}}^{\circ} = E_{\text{Ag}^+/\text{Ag}}^{\circ} - E_{\text{Ni}^{2+}/\text{Ni}}^{\circ} = 0.80 - 0.25$$

$$E_{\text{cell}}^{\circ} = 0.55$$

Equilibrium constant ( $K$ ) =

$$\text{Antilog} \left( \frac{nE^{\circ}}{0.0591} \right) = \text{antilog} \left( 2 \times \frac{0.55}{0.0591} \right)$$

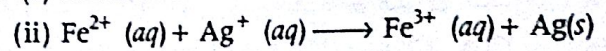
$$K = 4 \times 10^{18} \quad (1)$$

Maximum work ( $\Delta G^{\circ}$ ) =  $-nFE^{\circ}$

$$= -2 \times 96500 \times 0.55 = -106150 \text{ J mol}^{-1} \quad (1)$$

20. (i) Refer to solution 10.

(2)



$$E_{\text{cell}}^{\circ} = E_{\text{Ag}^+/\text{Ag}}^{\circ} - (E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^{\circ})$$

$$= 0.80 \text{ V} - 0.77 \text{ V} = 0.03 \text{ V} \quad (1)$$

$$\Delta_r G^{\circ} = -nFE^{\circ} = -1 \times 96500 \times 0.03 \quad (1)$$

$$= -2895 \text{ J mol}^{-1}$$

Equilibrium constant ( $K$ ) =  $\text{antilog} \left( \frac{nE^{\circ}}{0.0591} \right)$

$$K = 3.21 \quad (1)$$

21. (i) Refer to solution 10.

(2)

- (ii) Given, conductivity,

$$\kappa = 0.146 \times 10^{-3} \text{ S cm}^{-1}$$

Resistance,  $R = 1500 \Omega$

$\therefore$  Cell constant,  $G^* = \kappa \times R$

$$= 0.146 \times 10^{-3} \times 1500 = 0.219 \text{ cm}^{-1} \quad (3)$$

## ☑ Explanations

1. Given,  $w = 1.50 \text{ g}$ ;  $I = 1.50 \text{ A}$

Molar mass of Ag =  $108 \text{ g mol}^{-1}$

$$Z = \frac{\text{Equivalent weight}}{F}$$

where,  $F = 96500 \text{ C}$

As we know,  $w = Z \times I \times t$

$$\begin{aligned} \text{or } t &= \frac{w}{Z \times I} = \frac{1.50 \times 96500}{108 \times 1.50} \\ &= 89351 \text{ s} \end{aligned}$$

2. Given, conductivity of  $\text{CH}_3\text{COOH}$  solution,  
 $\kappa = 1.65 \times 10^{-4} \text{ S cm}^{-1}$

Concentration of  $\text{CH}_3\text{COOH}$  solution,

$$M = 0.01 \text{ M}$$

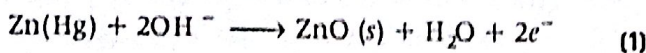
(2)

$$\begin{aligned} \text{Molar conductivity, } \Lambda_m &= \frac{\kappa \times 1000}{M} \\ &= \frac{(1.65 \times 10^{-4}) \times 1000}{0.01 \text{ M}} \\ &= 16.5 \text{ S cm}^2 \text{ mol}^{-1} \end{aligned} \quad (2)$$

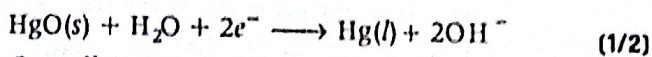
3. Mercury cell is suitable for low current devices like hearing aids, watches, etc.

The electrode reactions of a mercury cell can be written as

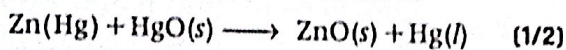
**At anode**



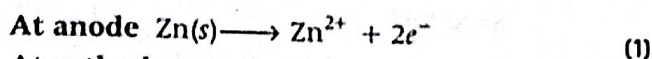
**At cathode**



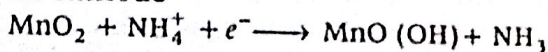
Overall reaction of the cell is written as



4. Leclanche cell is commonly used in transistors. The electrode reactions of Leclanche cell can be written as



**At cathode**



$\text{NH}_3$  formed in reaction at cathode combines with  $\text{Zn}^{2+}$  to form the complex like  $[\text{Zn}(\text{NH}_3)_4]^{2+}$  (1)

5. (i) Mercury cell (1/2)  
 (ii) Fuel cell (1/2)  
 (iii) Lead storage cell (1/2)  
 (iv) Dry cell (1/2)

6. Given,  $w = 1.27 \text{ g}$ ,  $I = 2 \text{ A}$

Molar mass of  $\text{Cu} = 63.5 \text{ g mol}^{-1}$ ,

$$Z = \frac{\text{Equivalent weight}}{F}$$

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

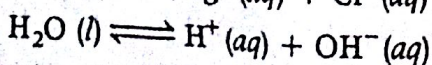
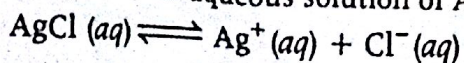
$$w = Z \times I \times t \quad (1)$$

$$t = \frac{w}{Z \times I} = \frac{1.27 \times 2 \times 96500}{63.5 \times 2}$$

$$t = 1930 \text{ s}$$

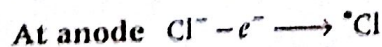
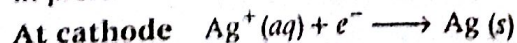
$$= 32.1 \text{ min} \quad (1)$$

7. (i) Electrolysis of an aqueous solution of  $\text{AgCl}$ .



On passing electricity,  $\text{Ag}^+(\text{aq})$  and  $\text{H}^+(\text{aq})$  move towards the cathode. Since, the discharge potential of  $\text{Ag}^+$  ion is lower than

that of  $\text{H}^+$  ions. In the other words,  $\text{Ag}^+$  ions have more value of standard reduction potential. Therefore,  $\text{Ag}^+$  ions are discharged in preference of  $\text{H}^+$  ions.

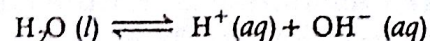
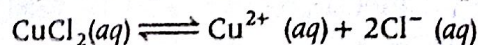


Hence, reduction of  $\text{Ag}^+$  (ion) is feasible at cathode.

- (ii) **Limiting molar conductivity** Refer to solution 2 of topic 2.

Conductivity of an electrolyte solution decreases with the decrease in concentration due to the fact that the number of ions per unit volume that carry the current in a solution decreases on dilution.

8. (i) Electrolysis of aqueous solution of  $\text{CuCl}_2$



On passing electricity,  $\text{Cu}^{2+}(\text{aq})$  and  $\text{H}^+(\text{aq})$  move towards the cathode. Since, the discharge potential of  $\text{Cu}^{2+}$  ion is lower than that of  $\text{H}^+$  ions, or  $\text{Cu}^{2+}$  ion have more value of standard reduction potential.

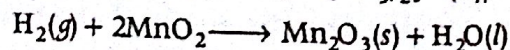
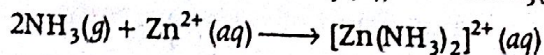
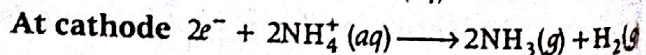
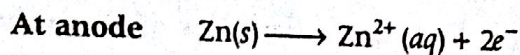
Therefore,  $\text{Cu}^{2+}$  ions are discharged in the presence of  $\text{H}^+$  ions.

- (ii) **Kohlrausch's law of independent migration of ions** Refer to solution 5 of topic 2

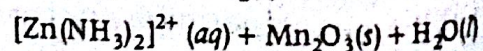
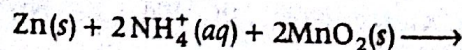
**Application** Using this law, it is possible to calculate  $\Lambda_m^\circ$  for weak electrolyte from the  $\lambda^\circ$  of individual ions at infinite dilution.

9. **Dry cell** (Leclanche cell) consists of a zinc container which acts as anode and cathode is a carbon (graphite) rod surrounded by powdered manganese dioxide and carbon. The space between the electrodes is filled by a moist paste of  $\text{NH}_4\text{Cl}$  and  $\text{ZnCl}_2$ .

The reactions are:



**Overall reaction**



By using Nernst equation,

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{2.303 RT}{nF} \ln \frac{[\text{Zn}(\text{NH}_3)_2]^{2+}}{[\text{NH}_4^+]^2}$$

Due to the presence of ions in the overall reaction, its voltage decreases with time. (1)

10. (i) First find the amount of charge by  $Q = I \times t$   
(ii) Then, mass of copper deposited using the formula,  $m = ZIt$   

$$= \frac{\text{Equivalent weight}}{96500} \times I \times t$$

Given, Current,  $I = 5.0 \text{ A}$

time ( $t$ ) =  $20 \times 60 = 1200 \text{ s}$

$$Q = I \times t \Rightarrow Q = 5.0 \times 1200 = 6000 \text{ C}$$

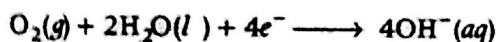
$$m = Z \times I \times t = \frac{\text{Equivalent weight}}{96500} \times Q$$

$$= \frac{58.7}{2 \times 96500} \times 6000 = 1.82 \text{ g} \quad (1)$$

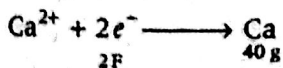
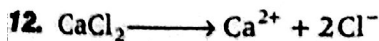
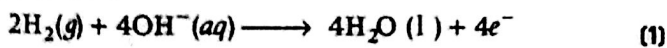
11. Those galvanic cells in which chemical energy of combustion of fuels like hydrogen, methane, etc. is converted into electrical energy are called fuel cells. (1)

The electrode reactions involved in the working of  $\text{H}_2$ - $\text{O}_2$  fuel cell are as follows:

**At cathode,**



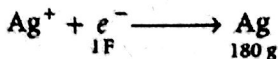
**At anode,**



The electricity required to produce  $40 \text{ g} = 2 \text{ F}$

Therefore, electricity required to produce  $20 \text{ g} = (2 \times 20) / 40 = 1 \text{ F}$  (2)

13. Electrode reaction,



$\therefore 96500 \text{ C (1F)}$  of electricity deposit  $\text{Ag} = 108 \text{ g}$

Quantity of current passed = Current  $\times$  Time

$$= 0.5 \times 60 \times 60 \times 2 = 3600 \text{ C}$$

$3600 \text{ C}$  of electricity will deposit  $\text{Ag}$

$$= \frac{108}{96500} \times 3600 = 4.03 \text{ g} \quad (1)$$

Now, let the thickness of deposition be  $a \text{ cm}$

Mass = volume  $\times$  Density or Area  $\times$  Thickness  $\times$  density

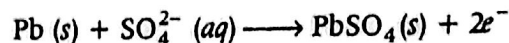
$$4.03 \text{ g} = 900 \text{ cm}^2 \times a \text{ cm} \times 10.5 \text{ g cm}^{-3}$$

$$a = 4.26 \times 10^{-4} \text{ cm} \quad (1)$$

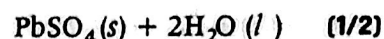
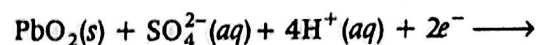
14. Lead storage battery is a secondary cell. Thus, it can be recharged by passing direct current through it. Therefore, it can be reused. It is used in automobiles. (1/2)

In a lead storage cell, the anode is made of spongy lead and the cathode is a grid of lead packed with lead dioxide. The electrolyte used is  $\text{H}_2\text{SO}_4$  (38% by mass). (1/2)

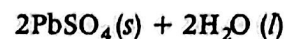
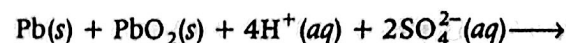
**At anode**



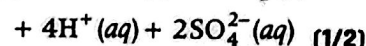
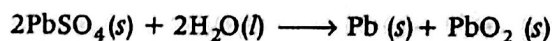
**At cathode**



**The overall cell reaction**

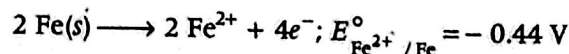


On recharging the battery, the reaction is reversed.

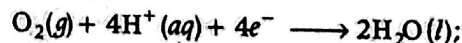


15. The chemistry of corrosion of iron is essentially an electrochemical phenomenon. In which a metal oxide or other salt of metal forms a coating on metal oxide. The reactions occurring during the corrosion of iron in the atmosphere are as follows:

**At anode,**

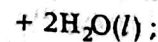
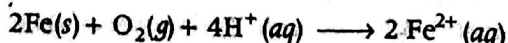


**At cathode**



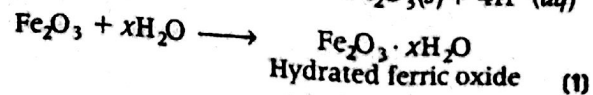
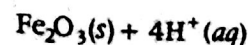
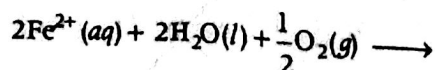
$$E_{\text{H}^+/\text{O}_2/\text{H}_2\text{O}}^{\circ} = 1.23 \text{ V}$$

**Overall reaction**



$$E_{\text{cell}}^{\circ} = 1.67 \text{ V} \quad (1)$$

The ferrous ions are further oxidised to ferric ions by atmospheric oxygen. Reaction is as follows:



16. Given, current = 2A,

$$\text{Time} = 3 \text{ h} = 3 \times 60 \times 60 \text{ s}$$

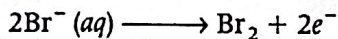
$$m = ZIt = \frac{E}{96500} It \quad (1/2)$$

$$= \frac{200.6 \times 2 \times 3 \times 60 \times 60}{2 \times 96500} \quad (1/2)$$

$$= 22.45 \text{ g} \quad (1/2)$$

$$\text{Number of moles} = \frac{22.45}{200.6} = 0.112 \text{ mol} \quad (1/2)$$

17. In the electrolysis of NaBr,



occurs at anode because it has lower reduction potential, and liberation of  $\text{O}_2$  requires over voltage. (2)

18. Mercury cell

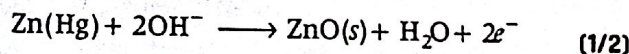
**Anode** Zinc-mercury amalgam

**Cathode** Paste of HgO and carbon

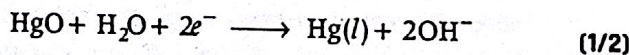
**Electrolyte** Paste of KOH and ZnO

Electrode reactions are as follows

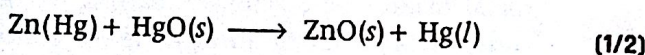
**At anode**



**At cathode**



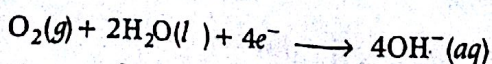
**Overall reaction**



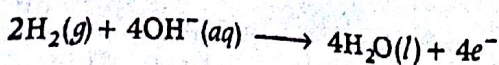
The voltage of a mercury cell remains constant during its life as the overall reaction does not involve any ion in solution whose concentration can change during its life-time. (1/2)

19. **Fuel cells** Galvanic cells that are designed to convert the chemical energy from combustion of fuels like hydrogen, methane, methanol directly into electrical energy are called fuel cells. The efficiency of fuel cell is higher than that of galvanic cell, e.g. hydrogen-oxygen fuel cell, in which hydrogen and oxygen are bubbled through porous carbon electrodes into conc. KOH solution. Finely divided Pt or Pd metals are incorporated into the electrodes for increasing the rate of electrode reactions. (1)

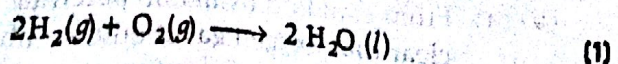
**At cathode**



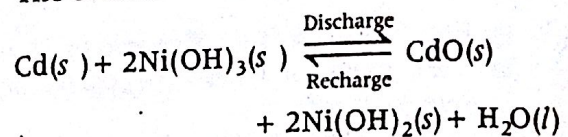
**At anode**



**Overall reaction**



20. Nickel-cadmium cell is a secondary cell. It has longer life than the lead storage cell but more expensive to manufacture. (1/2)  
The overall reaction during discharging is (1/2)



21. Refer to solution 18. (2)

22. Refer to solution 20. (2)

23. Given, Charge ( $Q$ ) = 24000 C and

$$\text{current } (i) = 12.5 \text{ A.}$$

$$\text{Mass of chromium (Cr) deposited} = 1.5 \text{ g}$$

$$\text{Molar mass of (Cr)} = 52 \text{ g mol}^{-1}$$

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

$$\therefore 6 \times 96500 \text{ C charge deposit Cr} = 52 \text{ g}$$

$$\therefore 24000 \text{ C charge deposit Cr} = \frac{52 \times 24000}{6 \times 96500}$$

$$= \frac{1248000}{6 \times 96500}$$

$$\text{Mass of chromium deposited} = 2.16 \text{ g} \quad (2)$$

Also,

$$\therefore \text{Charge } (Q) = i \times t \text{ (time)}$$

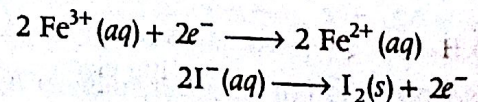
$$\therefore t = \frac{Q}{i} = \frac{24000}{12.5} = 1920 \text{ s}$$

$$\text{Duration of current} = 32 \text{ min}$$

24. (i) Standard Gibbs Free energy is given by ...  
 $\Delta G^\circ = -nFE^\circ_{\text{cell}}$   
where,  $n$  = number of moles of electrons transferred,  $F$  = Faraday's constant = 96500 C mol<sup>-1</sup>

$$E^\circ_{\text{cell}} = \text{Cell potential}$$

Two half-reactions for the given redox reaction can be written as



2 moles of electrons are involved in the reaction, hence  $n = 2$

Therefore, by substituting all the values in Eq. (i) we get

$$\Delta_r G^\circ = - (2 \text{ mol}) \times (96500 \text{ C mol}^{-1}) \times (0.236 \text{ V})$$
$$= - 45548 \text{ J}$$

$$\Delta_r G^\circ = - 45.55 \text{ kJ} \quad (1/2)$$

(ii) Given, current ( $i$ ) = 0.5 A, time ( $t$ ) = 2h  
Quantity of charge ( $Q$ ) passed = Current ( $i$ )  $\times$  time ( $t$ ) = (0.5 A)  $\times$  (2  $\times$  60  $\times$  60 s) = 3600 C (1/2)

Again,  $Q = ne^-$

where,  $n$  = number of electrons

$e^-$  = charge on electron

$$\therefore n = \frac{Q}{e^-} = \frac{3600\text{C}}{1.6 \times 10^{-19}\text{C}} = 2250 \times 10^{19}$$

Thus, number of electrons =  $2.250 \times 10^{22}$  (1½)

15. (i)

To find out the mass of Ag deposited at cathode, we first need to find the quantity of electricity passed ( $Q$ ) using faraday's law. The value of  $Q$  helps in determining the mass of 108 g of Ag.

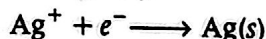
Given, current ( $I$ ) = 2A

time ( $t$ ) = 15 min

Quantity of electricity passed will be

$$Q = It = 2 \times 15 \times 60 = 1800\text{C}$$

Electrolysis of  $\text{AgNO}_3$



(Atomic mass of Ag = 108 g mol<sup>-1</sup>)

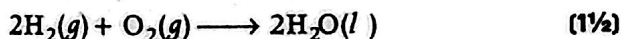
As, 96500 C deposit 108 g of Ag

$$\therefore 1800\text{C will deposit} = \frac{108 \times 1800}{96500} \text{ g of Ag}$$

$$= 2.014 \text{ g of Ag} \quad (1\frac{1}{2})$$

(ii) **A Fuel cell** Refer to solution 11.

The overall reaction of the cell is



16. (i) For the reaction,



$n = 2$

$$E_{\text{cell}}^\circ = +2.71\text{V}, 1\text{F} = 96500\text{C mol}^{-1} \quad (\text{given})$$

$$\Delta G^\circ = -nFE_{\text{cell}}^\circ \quad (1)$$

$$= -2 \times 96500 \times 2.71 = -523030\text{J mol}^{-1}$$

$$= -523.030\text{kJ mol}^{-1} \quad (1)$$

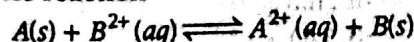
(ii)  $\text{H}_2$ - $\text{O}_2$  fuel cell was used in Apollo space programme for providing electrical power. (1)

17. (i) Advantages of fuel cells

(a) Because of the continuous supply of fuels, these cells never become dead. These cells are usually operated at a temperature of 70-140°C and gives a potential of about 0.9 V.

(b) Fuel cells do not cause any pollution unlike thermal plant which burn fossil fuels like coal gas, oil, etc. (1)

(ii) The reaction



By using Nernst equation,

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.059}{n} \log \frac{[\text{A}^{2+}]}{[\text{B}^{2+}]} \quad (1)$$

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

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

$$K_C = 10 \text{ and } n = 2$$

$$E_{\text{cell}}^\circ = \frac{0.059}{2} \log 10 \quad (\because \log 10 = 1)$$

$$E_{\text{cell}}^\circ = \frac{0.059}{2} = 0.0295 \quad (1)$$

28. Refer to solution 14. (3)

29.

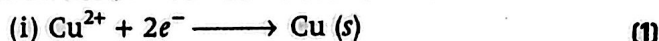
(i) Cathode is the site of reduction.

(ii) Use  $Q = It$  to find the amount of charge passed and  $m = ZIt$

$$= \frac{\text{Equivalent weight} \times It}{96500}$$

to find the mass of copper deposited.

Given that,  $t = 50 \times 60 = 3000\text{ s}$ ,  $I = 0.1287\text{A}$



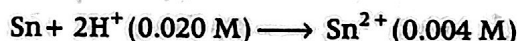
$$(b) m = Z \times I \times t = \frac{\text{Equivalent weight}}{96500} \times Q$$

$$= \frac{63.5}{2 \times 96500} \times 386.1 = 0.127\text{g} \quad (1)$$

30. (i)  $\text{Sn}(s) \mid \text{Sn}^{2+}(0.004\text{M}) \parallel \text{H}^+(0.020\text{M})$

$\mid \text{H}_2(g) (1\text{bar}) \mid \text{Pt}(s)$

Cell reaction is



$+ \text{H}_2(1\text{bar})$

$$E_{\text{cell}}^\circ = E_{\text{H}^+, \frac{1}{2}\text{H}_2}^\circ - E_{\text{Sn}^{2+}/\text{Sn}}^\circ$$

$$= 0 - (-0.14\text{V}) = 0 + 0.14\text{V} = 0.14\text{V}$$

For calculation of emf, apply Nernst equation;

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0591}{n} \log \frac{[\text{Sn}^{2+}] \times p_{\text{H}_2}}{[\text{H}^+]^2}$$

$$= 0.14 - \frac{0.0591}{2} \log \frac{[0.004] \times 1}{[0.020]^2}$$

$$= 0.14 - \frac{0.0591}{2} \log \frac{0.004}{0.0004}$$

$$= 0.14 - \frac{0.0591}{2} \log 10$$

$$= 0.14 - \frac{0.0591}{2} = 0.110\text{V}$$

(2½)

(ii) (a) From standard oxidation potential it is clear that oxygen gas should be liberated at anode, but its rate of production is very

low. In order to increase that we increase the voltage of external battery. Because of which chloride ions get oxidised easily and  $\text{Cl}_2$  gas is liberated at anode.

- (b) Conductivity of solution is conductance of ions present in unit volume of solution. On dilution number of  $\text{CH}_3\text{COOH}$  ions per unit volume decreases. Hence conductivity decreases. (2½)

31. (i) Gibbs energy of reaction is

$$\Delta G^\circ = -nFE^\circ_{\text{cell}} \quad \dots(i)$$

Given,  $\Delta G^\circ = -43600 \text{ J}$ ,  $n=2$ ,  $E^\circ_{\text{cell}} = ?$ ,

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

On putting the above values in Eq. (i), we get

$$-43600 \text{ J} = -2 \times 96500 \times E^\circ$$

$$E^\circ = \left( \frac{43600}{2 \times 96500} \right) \text{ V}$$

$$E^\circ = 0.22 \text{ V}$$

Now using Nernst equation,

$$E = E^\circ - \frac{0.0591}{n} \log \frac{[\text{Product}]}{[\text{Reactant}]}$$

$$E = E^\circ - \frac{0.0591}{n} \log \left[ \frac{[\text{H}^+]^2 [\text{Cl}^-]^2}{\text{H}_2(g)} \right] \quad (n=2)$$

Putting values, we get

$$E = 0.22 - \frac{0.0591}{2} \log \left[ \frac{(0.1)^2 (0.1)^2}{1} \right]$$

$$E = 0.22 - \frac{0.0591}{2} \log (10^{-4})$$

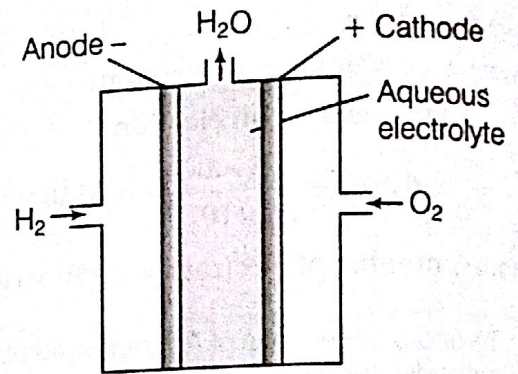
$$(\log 10^{-x} = -x)$$

$$E = 0.22 - \frac{0.0591}{2} (-4) = 0.3382 \text{ V}$$

$$\therefore E = 0.3382 \text{ V} \quad (2\frac{1}{2})$$

- (ii) It is a galvanic cell in which chemical energy from combustion of fuels is converted into electrical energy. In this type of galvanic cells, reactants are continuously feed to the electrodes, which react to produce electricity and product thus formed are continuously removed. Fuel cells that are designed to convert the energy of combustion of fuels like hydrogen, methane, methanol, etc.

directly into electrical energy are highly efficient (70%) when compared to efficiency thermal power plants (40%).



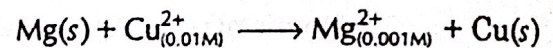
Fuel cell using and produces electricity

Their two advantages are :

- fuel cells are pollution free.
- They never become dead due to continuous supply of fuel.

There have been tremendous progress in the development of new electrode materials, better catalyst and electrolytes for increasing the efficiency of fuel cells.

32. For the reaction,



Given,  $E^\circ_{\text{cell}} = 2.71 \text{ V}$

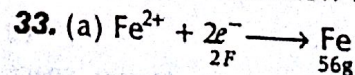
$$\therefore E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{n} \log \frac{[\text{Mg}^{2+}]}{[\text{Cu}^{2+}]}$$

$$E_{\text{cell}} = 2.71 - \frac{0.0591}{2} \log \frac{[0.001]}{[0.01]}$$

$$E^\circ_{\text{cell}} = 2.71 + \frac{0.0591}{2} \log 10$$

$$E^\circ_{\text{cell}} = 2.74 \text{ V}$$

- When  $E_{\text{ext}} < 2.71 \text{ V}$ ; the current flows from Cu to Mg as the electron flows from Mg to Cu.
- When  $E_{\text{ext}} > 2.71 \text{ V}$ , the reactions are reversed and so, the polarity of the electrodes. Thus, current will flow from Mg to Cu.



$\therefore 56 \text{ g}$  of Fe is deposited by  $2F$  or  $2 \times 96500 \text{ C}$  charge

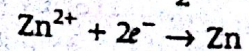
$\therefore 2.8 \text{ g}$  of Fe will be deposited by

$$\frac{96500 \times 2 \times 2.8}{56} \text{ C charge} = 9650 \text{ C}$$

Now,

$$Q = It$$

$$\therefore \text{Time, } t = \frac{9650}{2} = 4825 \text{ s}$$



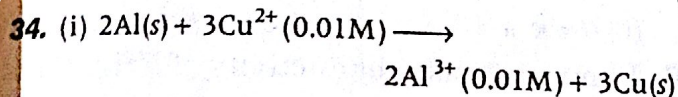
$\therefore 2 \times 96500 \text{ C}$  charge deposits Zn = 653g

$\therefore 9650 \text{ C}$  charge will deposit

$$\text{Zn} = \frac{65.3 \times 9650}{96500 \times 2} = 3.27 \text{ g}$$



- (b) (i) 'A' is a strong electrolyte since there is a small increase in  $\Lambda_m$  with dilution while 'B' is a weak electrolyte since  $\Lambda_m$  increases sharply near high dilution. (1)
- (ii) • For strong electrolyte (A) the plot becomes linear near high dilution and we can determine the value of  $\Lambda_m^\circ$  on extrapolation of  $\Lambda_m$  to  $C \rightarrow 0$ .
- For weak electrolyte,  $\Lambda_m$  increases sharply near high dilution. Hence, an extrapolation of  $\Lambda_m$  to  $C \rightarrow 0$  is not possible. (2)



$$E_{\text{cell}} = 1.98 \text{ V}$$

Applying Nernst equation for the given cell reaction,

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0591}{n} \log \frac{[Al^{3+}]^2}{[Cu^{2+}]^3}$$

$$E_{\text{cell}}^\circ = E_{\text{cell}} + \frac{0.0591}{n} \log \frac{[Al^{3+}]^2}{[Cu^{2+}]^3} \quad (1)$$

Here,  $n = 6$  ( $n =$  number of electrons transferred)

$$= 1.98 + \frac{0.0591}{6} \log \frac{(0.01)^2}{(0.01)^3}$$

$$= 1.98 + \frac{0.0591}{6} \log 100 = 1.98 + \frac{0.0591}{6} \times 2$$

[ $\log 100 = \log 10^2 = 2 \log 10 = 2$ , as  $\log 10 = 1$ ]

$$= 1.9997 \text{ V} \approx 2.0 \text{ V} \quad (2)$$

- (ii) Lower the reduction potential, more reactive is the metal. Coating the surface of iron with a more active metal prevents the iron from losing electrons and hence from corrosion.

As reduction potential of A is lower than that of iron, therefore A would be used for coating the surface of iron to prevent it from corrosion. This is because lower the reduction potential of an element, more reactive is the element. Therefore, A would be more reactive than iron as well as B and hence, A protects the iron metal from corrosion. (2)

35. (i) (a) Refer to solution 10 of Topic '2'. (1)
- (b) **Secondary batteries** After use, they can be recharged by passing current through it in the opposite direction so that it can be used again. Secondary cell undergo a large number of discharging and charging cycles. The most important secondary cell

is the lead storage battery. Other example includes nickel-cadmium cell. (1)

- (c) Refer to solution 11. (1)
- (ii) (a) **Faraday's first law of electrolysis** The amount of substance deposited during electrolysis is directly proportional to quantity of electricity passed.

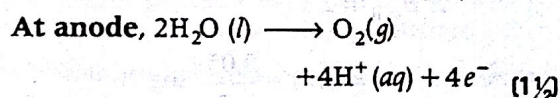
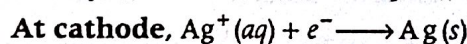
$$m \propto Q, \quad m \propto It, \quad m = ZIt,$$

where,  $Z$  is electrochemical equivalent,  $I$  is current in ampere,  $t$  is time in seconds and  $Q$  is charge in coulomb.

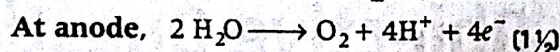
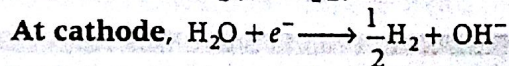
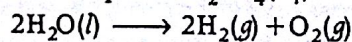
**Electrochemical equivalent ( $Z$ )** It is amount of substance deposited in grams when 1 ampere current is passed for 1 second, i.e. when 1 coulomb of charge is passed. Its unit is  $g C^{-1}$ . (1)

- (b) **Kohlrausch's law of independent migration of ions** Refer to solution 5 of Topic 2. (1)

36. (i) (a) Electrolysis of  $AgNO_3(aq)$  using Pt electrodes



- (b) Electrolysis of  $H_2SO_4(aq)$  using Pt electrodes.



- (ii) Reaction involved,  $\frac{2}{3}Al_2O_3 \longrightarrow \frac{4}{3}Al + O_2$

Here,  $n = 4$

$$\text{Given, } \Delta G = -960 \text{ kJ} \Rightarrow \Delta G = -nFE^\circ$$

$$-960000 \text{ J} = -4 \times 96500 \times E^\circ$$

$$E^\circ = \frac{960000 \text{ J}}{4 \times 96500 \text{ J}} = 2.48 \text{ V} \approx 2.5 \text{ V} \quad (2)$$

37. (i) (a) **Limiting molar conductivity** The molar conductivity of a solution at infinite dilution is called limiting molar conductivity and is represented by the symbol  $\Lambda_m^\circ$ . (1)

- (b) **Fuel cell** Refer to solution 19. (1)

- (ii) (a) First calculate the cell constant using relation, //  $A$  (cell constant) =  $R \times \kappa$  then, find conductivity ( $\kappa$ ).
- (b) Find molar conductivity of solution by the formula,  $\Lambda_m = \frac{\kappa \times 1000}{\text{Molarity}}$

(a) Cell constant, (1)

$$G^* = \text{conductivity} \times \text{resistance} \\ = 1.29 \text{ Sm}^{-1} \times 100 \Omega = 129 \text{ m}^{-1} = 1.29 \text{ cm}^{-1}$$

(b) Conductivity of 0.02 M KCl solution,

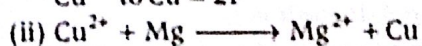
$$\kappa = \frac{\text{Cell constant}}{\text{Resistance}} = \frac{1.29 \text{ cm}^{-1}}{520 \Omega} \\ = 2.48 \times 10^{-3} \text{ S cm}^{-1} \quad (1)$$

(c) Molar conductivity,  $\Lambda_m = \frac{\kappa \times 1000}{M}$

$$= \frac{2.48 \times 10^{-3} \times 1000}{0.02} = 124 \text{ S cm}^2 \text{ mol}^{-1} \quad (1)$$

**38.** (i) Faraday's first law of electrolysis Refer to solution 35 (ii) (a)

Charge required for the reduction of 1 mole of  $\text{Cu}^{2+}$  to Cu = 2F (1)



Given,  $E^\circ_{\text{cell}} = +2.71 \text{ V}$

By using Nernst equation,

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.059}{n} \log \frac{[\text{Mg}^{2+}]}{[\text{Cu}^{2+}]} \quad (1)$$

Here,  $n = 2$  and  $E^\circ_{\text{cell}} = +2.71 \text{ V}$

$$\therefore E_{\text{cell}} = 2.71 - \frac{0.059}{2} \log \frac{(0.1)}{(0.01)} \quad (1)$$

$$= 2.71 - \frac{0.059}{2} \log 10$$

$$= 2.71 - 0.0295 = 2.68 \text{ V} \quad [:\log 10 = 1]$$