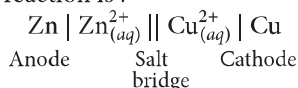


ELECTROCHEMISTRY

Detailed Solutions

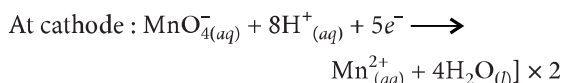
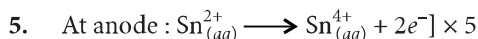
1. The device which converts the chemical energy liberated during the chemical reaction to electrical energy is called electrochemical cell. If external potential applied becomes greater than E°_{cell} of electrochemical cell then the cell behaves as an electrolytic cell and the direction of flow of current is reversed.

2. Representation of the galvanic cell for the given reaction is :

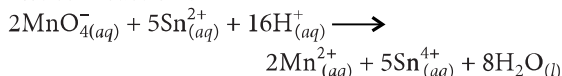


3. The salt bridge allows the movement of ions from one solution to the other without mixing of the two solutions. Moreover, it helps to maintain the electrical neutrality of the solutions in the two half cells.

4. Because standard electrode potential of $\text{Ti}^{4+}/\text{Ti}^{3+}$ is less than that of $\text{Fe}^{3+}/\text{Fe}^{2+}$ so, it cannot oxidise Fe^{II} to Fe^{III} .



Net cell reaction :

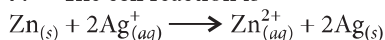


$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = 1.51 \text{ V} - 0.15 \text{ V} = 1.36 \text{ V}$$

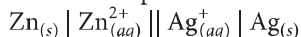
Since, cell potential is positive therefore the reaction is product favoured.

6. The reducing power increases with decreasing value of electrode potential. Hence, the order is $\text{Ag} < \text{Cu} < \text{Fe} < \text{Cr} < \text{Mg} < \text{K}$.

7. The cell reaction is

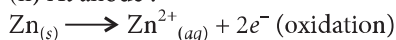


The cell is represented as



(i) Anode i.e., zinc electrode will be negatively charged.

(ii) At anode :

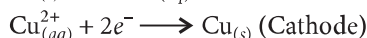
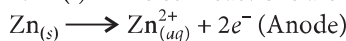


At cathode :

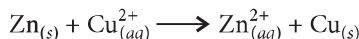


(iii) Ions are the carriers of current within the cell.

8. (i) The cell reactions are :



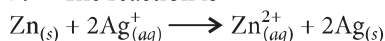
Net reaction :



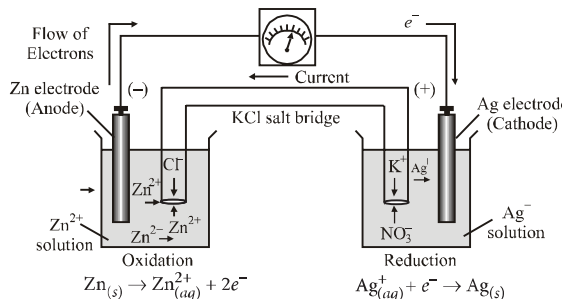
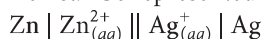
(ii) $E^\circ_{\text{cell}} = E^\circ_{\text{right}} - E^\circ_{\text{left}} = 0.34 \text{ V} - (-0.76 \text{ V}) = 1.10 \text{ V}$

(iii) Copper electrode will be positive on which reduction takes place.

9. The reaction is



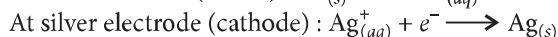
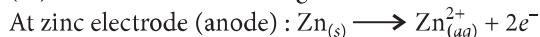
Cell can be represented as



(i) The zinc electrode is negatively charged (anode) as it pushes the electrons into the external circuit.

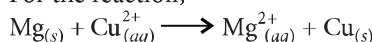
(ii) Ions are the current carriers within the cell.

(iii) The reactions occurring at two electrodes are :



10. (a) Given : $E^\circ_{\text{cell}} = 2.71 \text{ V}$

For the reaction,

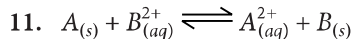


$n = 2, \Delta_r G^\circ = ?$

Using formula, $\Delta_r G^\circ = -nFE^\circ_{\text{cell}}$

$$\Delta_r G^\circ = -2 \times 96500 \text{ C mol}^{-1} \times 2.71 \text{ V}$$

$$\text{or } \Delta_r G^\circ = 523.03 \text{ kJ mol}^{-1}$$



Here, $n = 2$

using formula,

$$E^\circ_{\text{cell}} = \frac{0.059}{n} \log K_c$$

$$E^\circ_{\text{cell}} = \frac{0.059}{2} \log 10$$

$$E^\circ_{\text{cell}} = 0.0295 \text{ V}$$

12. Here $n = 2$, $E^\circ_{\text{cell}} = 1.1 \text{ V}$

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

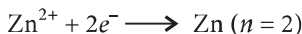
$$\Delta_r G^\circ = -nFE^\circ_{\text{cell}}$$

$$\Delta_r G^\circ = -2 \times 1.1 \times 96500 = -212300 \text{ J mol}^{-1}$$

$$= -212.3 \text{ kJ mol}^{-1}$$

13. Refer to answer 12.

14. The electrode reaction written as reduction reaction is



Applying Nernst equation, we get,

$$\left[E_{\text{Zn}^{2+}/\text{Zn}} = E^\circ_{\text{Zn}^{2+}/\text{Zn}} - \frac{0.0591}{2} \log \frac{1}{[\text{Zn}^{2+}]} \right]$$

As 0.1 M ZnSO_4 solution is 95% dissociated, this means that in the solution,

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

$$\therefore E_{\text{Zn}^{2+}/\text{Zn}} = -0.76 - \frac{0.0591}{2} \log \frac{1}{0.095}$$

$$= -0.76 - 0.02955 (\log 1000 - \log 95)$$

$$= -0.76 - 0.0295 (3 - 1.9777) = -0.79021 \text{ V}$$

15. $\text{Ni}_{(s)} + 2\text{Ag}^+_{(aq)} \rightarrow \text{Ni}^{2+}_{(aq)} + 2\text{Ag}_{(s)}$, $E^\circ = 1.05 \text{ V}$

Here, $n = 2$

$$\text{Using formula, } \log K_c = \frac{nE^\circ_{\text{cell}}}{0.059}$$

$$\text{or } \log K_c = \frac{2 \times 1.05}{0.059} = 35.5932$$

$$K_c = \text{antilog } 35.5932 \text{ or } K_c = 3.92 \times 10^{35}$$

Again, $\Delta_r G^\circ = -nFE^\circ_{\text{cell}}$

$$\Delta_r G^\circ = -2 \times 96500 \times 1.05 = -202650 \text{ J}$$

$$\Delta_r G^\circ = -202.65 \text{ kJ}$$

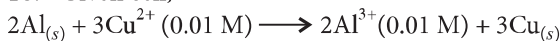
16. $E^\circ_{\text{cell}} = E^\circ_{\text{right}} - E^\circ_{\text{left}} = -0.44 - (-0.74) = 0.30 \text{ V}$

$$E_{\text{cell}} = 0.30 - \frac{0.0591}{6} \log \frac{(0.1)^2}{(0.01)^3}$$

$$= 0.30 - 0.0394 = 0.26 \text{ V}$$

17. Refer to answer 16.

18. Given cell,



$$E_{\text{cell}} = 1.98 \text{ V}, E^\circ_{\text{cell}} = ?$$

Using Nernst equation at 298 K

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{6} \log \frac{[\text{Al}^{3+}]^2}{[\text{Cu}^{2+}]^3}$$

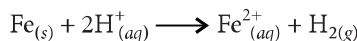
$$1.98 \text{ V} = E^\circ_{\text{cell}} - \frac{0.0591}{6} \log \frac{[10^{-2}]^2}{[10^{-2}]^3}$$

$$1.98 \text{ V} = E^\circ_{\text{cell}} - \frac{0.0591}{6} \log 10^2$$

$$1.98 \text{ V} = E^\circ_{\text{cell}} - \frac{0.0591}{6} \times 2$$

$$\therefore E^\circ_{\text{cell}} = 1.98 + \frac{0.0591}{6} = 1.99 \text{ V}$$

19. The cell reaction is



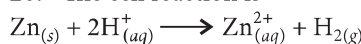
$$E^\circ_{\text{cell}} = 0.00 - (-0.44) = 0.44 \text{ V}$$

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

$$= 0.44 - \frac{0.0591}{2} \log \frac{0.001}{(0.01)^2}$$

$$= 0.44 - 0.02955 = 0.41045 \text{ V}$$

20. The cell reaction is



$$E^\circ_{\text{cell}} = 0.00 - (-0.76) = 0.76 \text{ V}$$

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

$$= 0.76 - \frac{0.0591}{2} \log \frac{(0.001)}{(0.01)^2}$$

$$= 0.76 - 0.02955 = 0.730 \text{ V}$$

21. At anode : $\text{Ni} \longrightarrow \text{Ni}^{2+} + 2e^-$

At cathode : $[\text{Ag}^+ + e^- \longrightarrow \text{Ag}] \times 2$

Cell reaction : $\text{Ni} + 2\text{Ag}^+ \longrightarrow \text{Ni}^{2+} + 2\text{Ag}$

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

$$= E^\circ_{\text{Ag}^+/\text{Ag}} - E^\circ_{\text{Ni}^{2+}/\text{Ni}} = 0.80 \text{ V} - (-0.25)$$

$$E^\circ_{\text{cell}} = 1.05 \text{ V}$$

$$E_{\text{cell}} = \frac{0.0591}{n} \log K_c$$

$$\log K_c = \frac{E^\circ_{\text{cell}} \times n}{0.0591} = \frac{1.05 \times 2}{0.0591}$$

$$\log k_c = 35.53$$

$$K_c = \text{antilog } 35.53 = 3.38 \times 10^{35}$$

22. $E^\circ_{\text{cell}} = +0.80 \text{ V} - 0.77 \text{ V} = +0.03 \text{ V}$

$$\Delta_r G^\circ = -nFE^\circ_{\text{cell}} = -1 \times 96500 \times 0.03$$

$$= -2895 \text{ J mol}^{-1} = -2.895 \text{ kJ mol}^{-1}$$

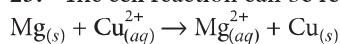
$$\Delta_r G^\circ = -2.303 RT \log K_c$$

$$-2895 = -2.303 \times 8.314 \times 298 \times \log K$$

$$\text{or } \log K_c = 0.5074$$

$$K_c = \text{Antilog } (0.5074) = 3.22$$

23. The cell reaction can be represented as :

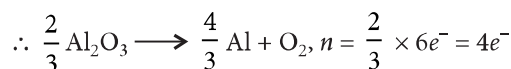
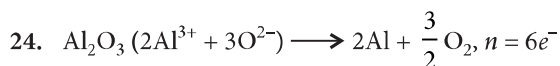


Given: $E_{\text{cell}}^{\circ} = +2.71 \text{ V}$, $T = 298 \text{ K}$

According to the Nernst equation :

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log \frac{[\text{Mg}_{(aq)}^{2+}]}{[\text{Cu}_{(aq)}^{2+}]}$$

$$= 2.71 - \frac{0.0591}{2} \log \frac{0.1}{0.01} = 2.6805 \text{ V}$$



$$\Delta G = 960 \times 1000 = 960000$$

Now, $\Delta G = -nFE_{\text{cell}}^{\circ}$

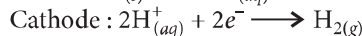
$$E_{\text{cell}}^{\circ} = -\frac{\Delta G}{nF} = \frac{-960000}{4 \times 96500}$$

$$E_{\text{cell}}^{\circ} = -2.487 \text{ V}$$

Minimum potential difference needed to reduce Al_2O_3 is -2.487 V .



Reactions :



$n = 2$.

Using Nernst equation at 298 K

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

For the given cell,

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = E_{\text{H}^{+}/\text{H}_2}^{\circ} - E_{\text{Fe}^{2+}/\text{Fe}}^{\circ}$$

$$= 0 - (-0.44) = +0.44 \text{ V}$$

Given $[\text{Fe}^{2+}] = 0.001 \text{ M}$; $[\text{H}^{+}] = 1 \text{ M}$; $p_{\text{H}_2} = 1 \text{ bar}$

Putting in Nernst equation

$$E_{\text{cell}} = 0.44 - 0.0295 \log \frac{0.001 \times 1}{1^2}$$

$$= 0.44 - 0.0295 \log 10^{-3}$$

$$= 0.44 - [(0.0295) \times (-3)]$$

$$= 0.44 + 0.0885 = 0.53 \text{ V}$$

26. The cell may be represented as



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

$$E_{\text{cell}} = 0.46 \text{ V} - 0.0295 \log \frac{(10^{-3})^2}{10^{-1}}$$

$$= 0.46 - 0.0295 \log 10^{-5} = 0.46 - 0.0295 (-5)$$

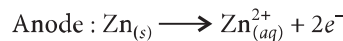
$$= 0.46 + 0.0295 \times 5 = 0.6075 \text{ V}$$

27. The cell reaction in button cell :

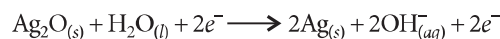


(i) Calculation of E_{cell}°

Reactions :



Cathode :



$n = 2$

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = E_{\text{Ag}_2\text{O}/\text{Ag}}^{\circ} - E_{\text{Zn}^{2+}/\text{Zn}}^{\circ}$$

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

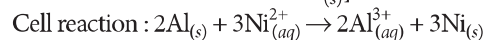
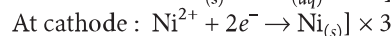
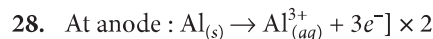
(ii) Calculation of $\Delta_r G^{\circ}$

$$\Delta_r G^{\circ} = -nFE_{\text{cell}}^{\circ}$$

$$= -2 \times 96500 \text{ C mol}^{-1} \times 1.56 \text{ V}$$

$$= -301080 \text{ C V mol}^{-1}$$

$$= -301080 \text{ J mol}^{-1} = -301 \text{ kJ mol}^{-1}$$



Applying Nernst equation to the above cell reaction,

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

$$\text{Now, } E_{\text{cell}}^{\circ} = E_{\text{Ni}^{2+}/\text{Ni}}^{\circ} - E_{\text{Al}^{3+}/\text{Al}}^{\circ}$$

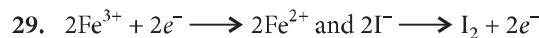
$$= -0.25 \text{ V} - (-1.66) = 1.41 \text{ V}$$

$$\therefore E_{\text{cell}} = 1.41 \text{ V} - \frac{0.0591}{6} \log \frac{(10^{-3})^2}{(0.5)^3}$$

$$= 1.41 \text{ V} - \frac{0.0591}{6} \log (8 \times 10^{-6})$$

$$= 1.41 \text{ V} - \frac{0.0591}{6} (5.09)$$

$$= 1.41 \text{ V} + 0.050 \text{ V} = 1.46 \text{ V}$$



Hence, for the given cell reaction, $n = 2$

$$\Delta_r G^{\circ} = -nFE_{\text{cell}}^{\circ} = -2 \times 96500 \times 0.236 = -45.55 \text{ kJ mol}^{-1}$$

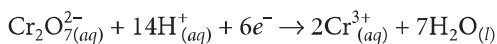
$$\Delta_r G^{\circ} = -2.303 RT \log K_c$$

$$\text{or } \log K_c = -\frac{\Delta_r G^{\circ}}{2.303 RT}$$

$$= -\frac{-45.55 \text{ kJ mol}^{-1}}{2.303 \times 8.314 \times 10^{-3} \text{ kJ K}^{-1} \text{ mol}^{-1} \times 298 \text{ K}}$$

$$\therefore K_c = \text{Antilog} (7.983) = 9.616 \times 10^7$$

30. For half cell reaction,

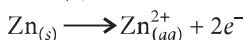


$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log \frac{[\text{Cr}^{3+}]^2}{[\text{Cr}_2\text{O}_7^{2-}][\text{H}^+]^{14}}$$

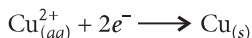
Given, $E_{\text{cell}}^{\circ} = 1.33 \text{ V}$, $n = 6$, $[\text{Cr}^{3+}] = 0.2 \text{ M}$
 $[\text{Cr}_2\text{O}_7^{2-}] = 0.1 \text{ M}$, $[\text{H}^+] = 1 \times 10^{-4} \text{ M}$

$$\begin{aligned} \Rightarrow E_{\text{cell}} &= 1.33 \text{ V} - \frac{0.0591}{6} \log \frac{(0.20)^2}{(0.1)(10^{-4})^{14}} \\ &= 1.33 \text{ V} - \frac{0.0591}{6} \log (4 \times 10^{55}) \\ &= 1.33 \text{ V} - \frac{0.0591}{6} [\log 4 + \log 10^{55}] \\ &= 1.33 \text{ V} - \frac{0.0591}{6} [\log 4 + 55 \log 10] \\ &= 1.33 \text{ V} - \frac{0.0591}{6} [0.602 + 55] \\ &= 1.33 \text{ V} - 0.548 \text{ V} = 0.782 \text{ V} \end{aligned}$$

31. (a) Oxidation half reaction :



Reduction half reaction :



(b) $E_{\text{cell}}^{\circ} = 0.34 - (-0.76) = 1.10 \text{ V}$

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

$$\begin{aligned} \Rightarrow E_{\text{cell}} &= 1.10 - \frac{0.0591}{2} \log \frac{2}{0.5} \\ &= 1.10 - \frac{0.0591}{2} \log \frac{2}{0.5} \\ &= 1.10 - \frac{0.059}{2} \times 0.6021 \\ &= 1.10 - 0.0177 = 1.0823 \text{ V} \end{aligned}$$

32. $\Delta G^{\circ} = -RT \ln K_c = -2.303 RT \log K_c$

$$\Rightarrow -212300 = -2.303 \times 8.314 \times 298 \times \log K_c$$

$$\text{or } \log K_c = \frac{212300}{2.303 \times 8.314 \times 298} = 37.2074$$

$$K_c = \text{Antilog } 37.2074 = 1.6 \times 10^{37}$$

33. The given cell may be represented as



$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = 0.80 \text{ V} - 0.34 \text{ V} = 0.46 \text{ V}$$

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

$$\text{or, } 0.422 \text{ V} = 0.46 \text{ V} - \frac{0.0591}{2} \log \frac{0.1}{[\text{Ag}^+]^2}$$

$$-0.038 \text{ V} = -0.0295 \log \frac{0.1}{[\text{Ag}^+]^2}$$

$$\text{or, } \log \frac{0.1}{[\text{Ag}^+]^2} = \frac{-0.038}{-0.0295} = 1.288$$

$$\text{or, } \frac{0.1}{[\text{Ag}^+]^2} = \text{antilog } 1.288 = 19.41$$

$$\therefore [\text{Ag}^+]^2 = \frac{0.1}{19.41} = 5.1519 \times 10^{-3}$$

$$[\text{Ag}^+] = 7.1 \times 10^{-2} \text{ M}$$

34. Refer to answer 26.

35. Refer to answer 28.

36. $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = -0.40 \text{ V} - (-0.44 \text{ V}) = 0.04 \text{ V}$

$$\text{Using formula, } \log K_c = \frac{nE^{\circ}}{0.0591} \text{ at } 298 \text{ K}$$

$$\text{or, } K_c = \text{antilog } \frac{2 \times 0.04 \text{ V}}{0.0591 \text{ V}}$$

$$\text{or, } K_c = \text{antilog } 1.356 = 22.38$$

37. The cell may be represented as



$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = 0.80 \text{ V} - (-0.76 \text{ V}) = 1.56 \text{ V}$$

Using formula,

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

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

$$\text{or, } \log [\text{Ag}^+] = -1.354$$

$$\text{or, } [\text{Ag}^+] = \text{antilog } (-1.354)$$

$$\text{or, } [\text{Ag}^+] = 4.426 \times 10^{-2} \text{ M}$$

38. $T = 273 + 25^{\circ}\text{C} = 298 \text{ K}$ and $n = 6$

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = -0.40 - (-0.74) = 0.34 \text{ V}$$

$$\Delta_r G^{\circ} = -nFE_{\text{cell}}^{\circ} = -6 \times 96500 \times 0.34 = -196860 \text{ J mol}^{-1}$$

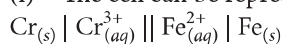
$$\text{Again } \Delta_r G^{\circ} = -2.303 RT \log K$$

$$\Rightarrow -196860 = -2.303 \times 8.314 \times 298 \times \log K$$

$$\Rightarrow \log K = 34.5014$$

$$\Rightarrow K = \text{antilog } 34.5014 = 3.172 \times 10^{34}$$

39. (i) The cell can be represented as

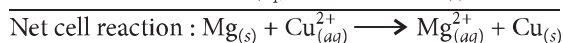
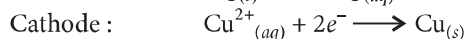
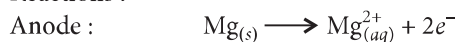


$$(ii) E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} \\ = -0.44 - (-0.74) = -0.44 + 0.74 = 0.30 \text{ V}$$

(iii) Refer to answer 16.



Reactions :



$$\therefore n = 2$$

Using Nernst equation :

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{2.303RT}{nF} \log \frac{[\text{Mg}^{2+}]}{[\text{Cu}^{2+}]}$$

For the given cell

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = E_{\text{Cu}^{2+}/\text{Cu}}^{\circ} - E_{\text{Mg}^{2+}/\text{Mg}}^{\circ} \\ = 0.34 \text{ V} - (-2.37 \text{ V}) = 2.71 \text{ V}$$

Given $[\text{Mg}^{2+}] = 0.001 \text{ M}$, $[\text{Cu}^{2+}] = 0.0001 \text{ M}$

Putting in Nernst equation at 298 K

$$E_{\text{cell}} = 2.71 \text{ V} - \frac{0.059}{2} \log \frac{0.001}{0.0001}$$

$$E_{\text{cell}} = 2.71 - 0.0295 \log 10 = 2.71 - 0.03 = 2.68 \text{ V}$$

$$\Delta_r G = -nFE_{\text{cell}} \\ = -2 \times 96500 \text{ C mol}^{-1} \times 2.68 \\ = -517,240 \text{ J mol}^{-1} = -517.24 \text{ kJ mol}^{-1}$$

41. We have

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059}{n} \log \frac{[\text{Ni}_{(aq)}^{2+}]}{[\text{Cu}_{(aq)}^{2+}]}$$

$$0.059 = E_{\text{cell}}^{\circ} - \frac{0.059}{2} \log \left(\frac{0.01}{0.1} \right) \quad [\text{Here } n = 2]$$

$$0.059 = E_{\text{cell}}^{\circ} - \frac{0.059}{2} \log \left(\frac{1}{10} \right)$$

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

$$0.059 = E_{\text{cell}}^{\circ} + \frac{0.059}{2}$$

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

$$\therefore E_{\text{cell}}^{\circ} = \frac{0.059}{2} = 0.0295 \approx 0.03$$

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

$$0.03 = 0.34 - E_{\text{anode}}^{\circ}$$

$$E_{\text{anode}}^{\circ} = 0.34 - 0.03 = 0.31 \text{ V}$$

Hence, $E_{\text{Ni}^{2+}/\text{Ni}}^{\circ} = +0.31 \text{ V}$

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

$$= -0.403 - (-0.763) = -0.403 + 0.763 = 0.360 \text{ V}$$

$$\therefore E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{2.303RT}{nF} \log \frac{[\text{Zn}_{(aq)}^{2+}]}{[\text{Cd}_{(aq)}^{2+}]}$$

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059}{n} \log \frac{[\text{Zn}_{(aq)}^{2+}]}{[\text{Cd}_{(aq)}^{2+}]}$$

$$\left[\because \frac{2.303RT}{F} = 0.059 \right]$$

$$= 0.36 - \frac{0.059}{2} \log \left(\frac{0.1}{0.01} \right) = 0.36 - \frac{0.059}{2} \log 10$$

$$= 0.36 - 0.0295 = 0.3305 = 0.33 \text{ V}$$

$$\therefore \Delta_r G^{\circ} = -nFE_{\text{cell}}^{\circ}$$

$$= -2 \times 96500 \times 0.36 = -69480 \text{ J mol}^{-1}$$

$$= -69.48 \text{ kJ mol}^{-1}$$

43. The limiting molar conductivity of an electrolyte is defined as its molar conductivity when the concentration of the electrolyte in the solution approaches zero.

Conductivity of an electrolyte decreases with dilution because the number of current carrying particles *i.e.*, ions present per cm^3 of the solution becomes less and less on dilution.

44. **Kohlrausch's law of independent migration of ions :** It states that limiting molar conductivity of an electrolyte can be represented as the sum of the individual contributions of the anion and cation of the electrolyte.

If $\lambda_{\text{Na}^+}^{\circ}$ and $\lambda_{\text{Cl}^-}^{\circ}$ are limiting molar conductivities of the sodium and chloride ions respectively then the limiting molar conductivity for sodium chloride is given by

$$\Lambda_m^{\circ}(\text{NaCl}) = \lambda_{\text{Na}^+}^{\circ} + \lambda_{\text{Cl}^-}^{\circ}$$

Kohlrausch's law helps in the calculation of degree of dissociation of weak electrolyte like acetic acid.

The degree of dissociation α is given by

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

where Λ_m is the molar conductivity and Λ_m° is the limiting molar conductivity.

45. Molar Conductivity : Molar conductivity of a solution at a dilution V is the conductance of all the ions produced from one mole of the electrolyte dissolved in $V \text{ cm}^3$ of the solution when the electrodes are one cm apart and the area of the electrodes is so large that the whole solution is contained between them.

$$\Lambda_m = \kappa V$$

Its units is $\text{S cm}^2 \text{mol}^{-1}$

46. Refer to answer 44.

47. Refer to answer 43.

48. Refer to answer 44.

$$49. \Lambda_m = \frac{\kappa \times 1000}{M} \text{ in CGS units}$$

$$\Lambda_m = \frac{\kappa \times 10^{-3}}{M} \text{ in SI units}$$

where κ is the conductivity, M is the molar concentration and Λ_m is molar conductivity.

$$50. \kappa = \frac{1}{R} \times \frac{l}{A}$$

where κ is the conductivity R is resistance and l/A is the cell constant.

51. Refer to answer 49.

52. The fraction of the total number of molecules present in solution as ions is known as degree of dissociation.

$$\text{Molar conductivity } (\lambda_m) = \alpha \lambda_m^\circ$$

where λ_m° is the molar conductivity at infinite dilution.

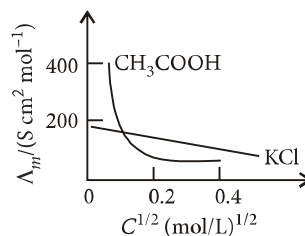
53. The reciprocal of resistivity is known as specific conductance or simply conductivity. It is denoted by κ (kappa).

$$\kappa = \frac{1}{\rho} \text{ or } \kappa = G \times \frac{l}{a}$$

Hence, conductivity of a solution is defined as the conductance of a conductor of 1 cm length and having 1 sq. cm as the area of cross section. Alternatively, it may be defined as conductance of one centimetre cube of the solution of the electrolyte. Molar conductivity of a solution at a dilution V is the conductance of all the ions produced from 1 mole of the electrolyte dissolved in $V \text{ cm}^3$ of the solution. It is represented by Λ_m .

$$\Lambda_m = \kappa V$$

Variation of conductivity and molar conductivity with concentration : Conductivity always decreases with decrease in concentration, for both weak and strong electrolytes. Because the number of ions per unit volume that carry the current in a solution decreases on dilution.



Molar conductivity vs $C^{1/2}$ for acetic acid (weak electrolyte) and potassium chloride (strong electrolyte) in aqueous solutions

Molar conductivity increases with decrease in concentration. Because that total volume, V , of solution containing one mole of electrolyte also increases. It has been found that decrease in κ on dilution of a solution is more than compensated by increase in its volume.

54. Refer to answers 43 and 44.

55. Refer to answer 53.

Variation of conductivity and molar conductivity with temperature : Both increase with increase in temperature as degree of ionisation increases.

$$56. \text{Conductivity } (\kappa) = \frac{1}{R} \times G = \frac{1}{200} \times 1 \\ = 5 \times 10^{-3} \Omega^{-1} \text{ cm}^{-1}$$

Molar conductivity (Λ_m)

$$= \frac{\kappa \times 1000}{M} = \frac{5 \times 10^{-3} \times 1000}{0.01} = 500 \Omega^{-1} \text{ cm}^2 \text{ mol}^{-1}$$

57. Refer to answer 55.

58. Here, $\kappa = 0.025 \text{ S cm}^{-1}$, Molarity = 0.20 M

$$\text{Molar conductivity } \Lambda_m^c = \frac{\kappa \times 1000}{\text{Molarity}}$$

$$= \frac{0.025 \times 1000}{0.20} = 125 \text{ S cm}^2 \text{ mol}^{-1}$$

59. $C = 0.001 \text{ M}$, $\kappa = 4 \times 10^{-5} \text{ S cm}^{-1}$,

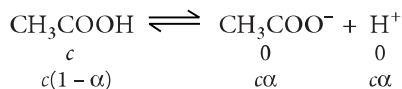
$$\Lambda_m^c = 390 \text{ S cm}^2/\text{mol}$$

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

Substituting the values,

$$\Lambda_m^c = \frac{4 \times 10^{-5} \times 1000}{0.001} = 40 \text{ S cm}^2/\text{mol}$$

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



$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]} = \frac{c\alpha \cdot c\alpha}{c(1-\alpha)} = \frac{c\alpha^2}{1-\alpha}$$

$$K_a = \frac{0.001(0.103)^2}{(1-0.103)} = \frac{1.061 \times 10^{-5}}{0.897} = 1.18 \times 10^{-5}$$

60. Refer to answers 49 and 50.

$$61. \Lambda_m^c = \frac{\kappa \times 1000}{M}$$

$$\Rightarrow \kappa = \frac{\Lambda_m^c \times M}{1000} = \frac{138.9 \times 1.5}{1000} = 0.20835 \text{ S cm}^{-1}$$

62. Here, conductivity (κ) = $0.146 \times 10^{-3} \text{ S cm}^{-1}$, resistance (R) = 1500Ω

$$\text{Cell constant} = \frac{\text{Conductivity}}{\text{Conductance}}$$

$$= \text{Conductivity} \times \text{Resistance}$$

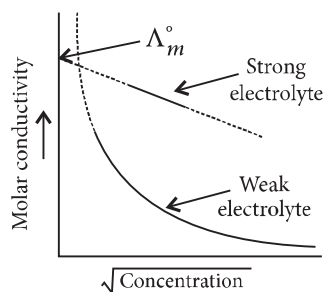
$$= \kappa \times R \quad \left[\because \text{conductance} = \frac{1}{\text{resistance}} \right]$$

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

63. **Strong electrolyte**: The molar conductivity of strong electrolyte decreases slightly with the

increase in concentration. This increase is due to increase in attraction as a result of greater number of ions per unit volume. With dilution the ions are far apart, interionic attractions become weaker and conductance increases.

Weak electrolyte: When the concentration of weak electrolyte becomes very low, its degree of ionisation rises sharply. There is sharp increase in the number of ions in the solution. Hence the molar conductivity of a weak electrolyte rises steeply at low concentration.



64. Refer to answers 45 and 49.

65. Refer to answer 44.

66. **Weak electrolytes**: The electrolytes which are not completely dissociated into ions in solution are called weak electrolytes e.g., CH_3COOH , NH_4OH , HCN , etc.

Strong electrolytes: The electrolytes which are completely dissociated into ions in solution are called strong electrolytes. e.g., HCl , KCl , NaOH , NaCl , etc.

$$67. \text{ Using formula, } \Lambda_m^c = \frac{\kappa \times 1000}{C}$$

$$\text{Given } \kappa = 3.905 \times 10^{-5} \text{ S cm}^{-1}$$

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

$$\therefore \Lambda_m^c = \frac{3.905 \times 10^{-5} \times 1000}{0.001} = 39.05 \text{ S cm}^2 \text{ mol}^{-1}$$

The degree of dissociation,

$$\alpha = \frac{\Lambda_m^c}{\Lambda_m^\circ} = \frac{39.05}{390.5} = 0.1$$

$$[\because \Lambda_m^\circ = 349.6 + 40.9 = 390.5 \text{ S cm}^2 \text{ mol}^{-1}]$$

68. Given: Conductivity, $\kappa = 0.0248 \text{ S cm}^{-1}$

Molarity, $C = 0.20 \text{ M} = 0.20 \text{ mol L}^{-1}$

$$\text{Using formula, } \Lambda_m = \frac{1000 \times \kappa}{C}$$

$$\Lambda_m^c = \frac{(1000 \text{ cm}^3 \text{ L}^{-1}) \times (0.0248 \text{ S cm}^{-1})}{(0.20 \text{ mol L}^{-1})}$$

$$= 124 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\alpha = \frac{\Lambda_m^c}{\Lambda_m^\circ} = \frac{124}{73.5 + 76.5} = 0.82$$

69. Resistance of 0.1 M KCl solution $R = 100 \Omega$

Conductivity $\kappa = 1.29 \text{ S m}^{-1}$

Cell constant $G^* = \kappa \times R = 1.29 \times 100 = 129 \text{ m}^{-1}$

Resistance of 0.02 M KCl solution, $R = 520 \Omega$

$$\text{Conductivity, } \kappa = \frac{\text{cell constant}}{R} = \frac{129 \text{ m}^{-1}}{520 \Omega} = 0.248 \text{ S m}^{-1}$$

Concentration, $C = 0.02 \text{ mol L}^{-1}$

$$= 1000 \times 0.02 \text{ mol m}^{-3} = 20 \text{ mol m}^{-3}$$

$$\text{Molar conductivity, } \Lambda_m = \frac{\kappa}{C} = \frac{0.248 \text{ S m}^{-1}}{20 \text{ mol m}^{-3}} = 0.0124 \text{ S m}^2 \text{ mol}^{-1}$$

$$70. \Lambda_m^\circ \text{Al}_2(\text{SO}_4)_3 = 2\lambda_m^\circ \text{Al}^{3+} + 3\lambda_m^\circ \text{SO}_4^{2-}$$

$$\Rightarrow 858 = 2\lambda_m^\circ \text{Al}^{3+} + 3 \times 160$$

$$\Rightarrow \lambda_m^\circ \text{Al}^{3+} = \frac{858 - 480}{2} = 189 \text{ S cm}^2 \text{ mol}^{-1}$$

71. Given : Diameter = 1 cm, length = 50 cm

$$R = 5.5 \times 10^3 \text{ ohm}, M = 0.05 \text{ M}$$

$$\rho = ? \quad \kappa = ? \quad \Lambda_m = ?$$

Area of the column,

$$a = \pi r^2 = 3.14 \times \left(\frac{1}{2} \text{ cm}\right)^2 = \frac{3.14}{4} \text{ cm}^2$$

Resistivity,

$$\rho = R \cdot \frac{a}{l} = 5.5 \times 10^3 \text{ ohm} \times \frac{3.14 \text{ cm}^2}{4 \times 50 \text{ cm}}$$

$$= 86.35 \text{ ohm cm}$$

$$\text{Again, conductivity, } \kappa = \frac{1}{\rho}$$

$$= \frac{1}{86.35} = 1.158 \times 10^{-2} \text{ ohm}^{-1} \text{ cm}^{-1}$$

$$\text{and molar conductivity, } \Lambda_m = \kappa \cdot \frac{10^3}{M}$$

$$= 1.158 \times 10^{-2} \text{ ohm}^{-1} \text{ cm}^{-1} \times \frac{10^3}{5 \times 10^{-2}}$$

$$= 231.6 \text{ ohm}^{-1} \text{ cm}^2 \text{ mol}^{-1}$$

$$72. \kappa = 1.29 \times 10^{-2} \text{ ohm}^{-1} \text{ cm}^{-1}$$

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

$$\Rightarrow \text{Cell constant} = \kappa \times R = 1.29 \text{ S m}^{-1} \times 85 = 109.65 \text{ m}^{-1}$$

For second solution,

$$\kappa = \frac{1}{R} \times \text{Cell constant} = \frac{1}{96 \Omega} \times 109.65 \text{ m}^{-1}$$

$$= 1.142 \Omega^{-1} \text{ m}^{-1}$$

$$\Lambda_m = \frac{\kappa \times 1000}{M} = \frac{1.142 \Omega^{-1} \text{ m}^{-1} \times 1000 \text{ cm}^3}{0.052}$$

$$\Lambda_m = \frac{1.142 \Omega^{-1} \text{ cm}^{-1} \times 10^{-2} \times 1000 \text{ cm}^3}{0.052 \text{ mol}}$$

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

73. According to Kohlrausch's law,

$$\Lambda^\circ \text{CH}_3\text{COOH} = \lambda^\circ \text{CH}_3\text{COO}^- + \lambda^\circ \text{H}^+$$

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

$$= \frac{11.7 \text{ S cm}^2 \text{ mol}^{-1}}{(49.9 + 349.1) \text{ S cm}^2 \text{ mol}^{-1}} = \frac{11.7}{390} = 3 \times 10^{-2}$$

74. For electrolyte X :

$$\text{Molarity} = 0.05 \text{ M}$$

$$\text{resistance} = 100 \text{ ohms}$$

$$\text{conductivity} = 1.0 \times 10^{-4} \text{ S cm}^{-1}$$

For electrolyte Y :

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

$$\text{resistance} = 50 \text{ ohms}$$

$$\text{conductivity} = ?$$

(i) Cell constant, G^*

$$= \text{conductivity} (\kappa) \times \text{resistance} (R)$$

$$= 100 \times 1 \times 10^{-4} = 10^{-2} \text{ cm}^{-1}$$

\(\therefore\) Conductivity of solution Y is

$$\kappa = \frac{G^*}{R} = \frac{10^{-2}}{50} = 0.02 \times 10^{-2} = 2 \times 10^{-4} \text{ S cm}^{-1}$$

$$\text{Molar conductance, } \Lambda_m = \frac{\kappa \times 1000}{M}$$

$$= \frac{2 \times 10^{-4} \times 1000}{0.01} = 20 \text{ ohm}^{-1} \text{ cm}^2 \text{ mol}^{-1}$$

75. (a) Refer to answers 45 and 63.

(b) Refer to answer 62.

76. (a) Refer to answer 44.

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

$$(b) \Lambda_m^\circ (\text{HCl}) = \lambda_{\text{H}^+}^\circ + \lambda_{\text{Cl}^-}^\circ$$

$$\Lambda_m^\circ (\text{NaCl}) = \lambda_{\text{Na}^+}^\circ + \lambda_{\text{Cl}^-}^\circ$$

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

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

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

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

$$= 426 + 91 - 126 = 391 \text{ S cm}^2 \text{ mol}^{-1}$$

77. Conductivity of acetic acid,

$$\kappa = 7.896 \times 10^{-5} \text{ S cm}^{-1},$$

$$\Lambda_m^\circ \text{ for acetic acid} = 390.5 \text{ S cm}^2 \text{ mol}^{-1}.$$

$$\text{Molar conductivity, } \Lambda_m^c = \frac{\kappa \times 1000}{\text{Molarity}}$$

$$= \frac{7.896 \times 10^{-5} \times 1000}{0.00241} = 32.76 \text{ S cm}^2 \text{ mol}^{-1}$$

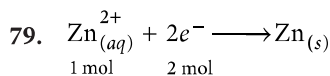
Degree of dissociation,

$$\alpha = \frac{\Lambda_m^c}{\Lambda_m^\circ} = \frac{32.76}{390.5} = 8.4 \times 10^{-2}$$

Dissociation constant of acetic acid,

$$K_a = \frac{C\alpha^2}{1-\alpha} = \frac{(0.00241) \times (8.4 \times 10^{-2})^2}{1-0.084} = 1.86 \times 10^{-5}$$

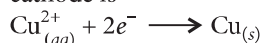
78. The species that get reduced at cathode is the one having higher value of standard reduction potential. Hence, the reaction that will occur at cathode is $\text{Ag}_{(aq)}^+ + e^- \longrightarrow \text{Ag}_{(s)}$.



One mole of Zn^{2+} requires 2 moles of electrons for reduction *i.e.*

$$Q = 2 \times F = 2 \times 96500 = 193000 \text{ C}$$

80. The species that get reduced at cathode is the one which have higher value of standard reduction potential. Hence, the reaction that will occur at cathode is



81. **Faraday's first law of electrolysis** : During electrolysis the amount of any substance deposited or liberated at an electrode is directly proportional to the quantity of electricity passed through the electrolyte *i.e.*,

$$w \propto Q \quad \text{or} \quad w \propto I \times t \quad [\because Q = I \times t]$$

$$w = Z \times I \times t$$

where, Z is a constant of proportionality known as electrochemical equivalent of the substance deposited.

82. The electrode reaction is $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$

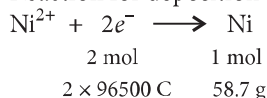
\therefore Quantity of charge required for reduction of 1 mole of $\text{Cu}^{2+} = 2F = 2 \times 96500 = 193000 \text{ C}$

83. Refer to answers 81 and 82.

84. Given : Current $I = 5 \text{ A}$; $t = 20 \times 60 \text{ s}$, $w = ?$

$$Q = I \times t = 5 \times 20 \times 60 = 6000 \text{ C}$$

Reaction for deposition of Ni,

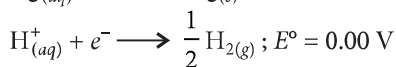


Thus, $2 \times 96500 \text{ C}$ of electricity produces 58.7 g Ni

\therefore 6000 C of electricity would produce

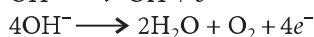
$$= \frac{58.7 \times 6000}{2 \times 96500} = 1.825 \text{ g}$$

85. (i) **At cathode** : The following reduction reactions compete to take place at the cathode.

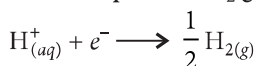


The reaction with a higher value of E° takes place at the cathode. Therefore, the deposition of silver will take place at the cathode.

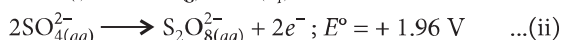
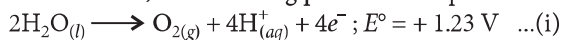
Since, Pt electrodes are inert, the anode is not attacked by NO_3^- ions. Therefore, OH^- or NO_3^- ions can be oxidized at the anode. But OH^- ions having a lower discharge potential get preference and decompose to liberate O_2 .



(ii) At the cathode, the following reduction reaction occurs to produce H_2 gas.

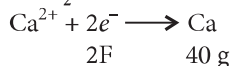
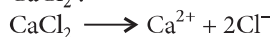


At the anode, the following processes are possible :



For dilute sulphuric acid, reaction (i) is preferred to produce O_2 gas but for concentrated sulphuric acid, (ii) occurs.

86. Reaction for production of Ca from molten CaCl_2 :



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

\therefore Electricity required to produce $20 \text{ g} = 0.5 \times 2F = 1F$

87. Calculation of mass of Ag deposited

The electrode reaction is $\text{Ag}^+ + e^- \longrightarrow \text{Ag}$

The quantity of electricity passed = Current \times Time
 $= 0.5 \text{ (amp)} \times 2 \times 60 \times 60 \text{ (sec)} = 3600 \text{ C}$

From the electrode reaction, it is clear that 96500 C of electricity deposit $\text{Ag} = 108 \text{ g}$

3600 C of electricity will deposit Ag

$$= \frac{108}{96500} \times 3600 = 4.03 \text{ g}$$

Calculation of thickness :

Let the thickness of deposit be $x \text{ cm}$

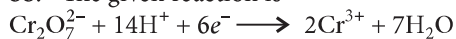
Mass = volume \times density = Area \times thickness \times density

[\because volume = area \times thickness]

$$\therefore 4.03 \text{ g} = 900 \text{ (cm}^2\text{)} \times x \text{ (cm)} \times 10.5 \text{ (g cm}^{-3}\text{)}$$

$$\therefore x = \frac{4.03}{900 \times 10.5} \text{ cm} = 4.26 \times 10^{-4} \text{ cm.}$$

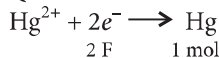
88. The given reaction is



one mole $\text{Cr}_2\text{O}_7^{2-}$ requires 6 mol of electrons for reduction. Hence, quantity of electricity required
 $= 6 \text{ mol} \times 96500 \text{ C mol}^{-1} = 5.79 \times 10^5 \text{ coulomb}$

89. Quantity of electricity passed

$$Q = I \times t = 2.0 \text{ A} \times 3 \times 60 \times 60 \text{ s} = 21600 \text{ C}$$



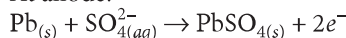
$2 \times 96500 \text{ C}$ electricity produces 1 mole Hg

$$\therefore 21600 \text{ C will produce} = \frac{21600}{2 \times 96500}$$

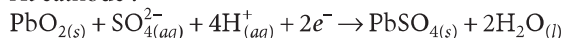
$$= 0.112 \text{ mole of Hg}$$

102. The lead storage battery is most important secondary cell. The cell reactions when the battery is in use :

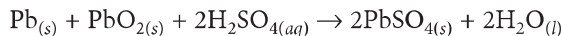
At anode:



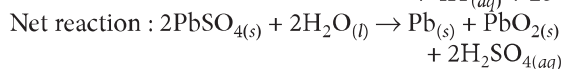
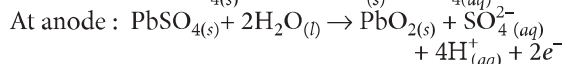
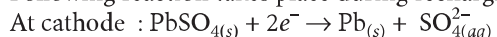
At cathode :



The overall cell reaction is



Following reaction takes place during recharging

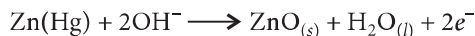


103. Refer to answer 101.

104. Mercury Cell : It is a miniature cell which finds a frequent use these days to supply energy for watches, video cameras, hearing aids and other compact devices. In mercury cell the anode is zinc-mercury amalgam, and the cathode is a paste of mercury (II) oxide and carbon, electrolyte is a moist paste of KOH - ZnO.

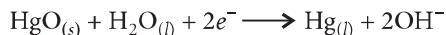
The cell reactions are as follows :

Anode :

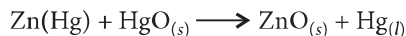


(Amalgam)

Cathode :



Net reaction :



The cell potential remains constant during its life as the overall reaction does not involve any ion in solution whose concentration can change during its use.

105. Refer to answer 101.

106. Nickel cadmium cell is a secondary battery which consists of a cadmium anode, nickel hydroxide as cathode and sodium or potassium hydroxide acts as electrolyte.

Merit : It has longer life than lead storage battery.

Demerit : It is more expensive than lead storage battery.

The following reaction takes place during discharging :



107. Those galvanic cells which give us direct electrical energy by the combustion of fuels like hydrogen, methane, methanol etc. are called fuel cells.

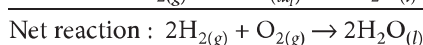
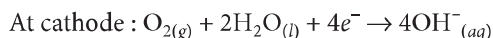
108. $\text{H}_2 - \text{O}_2$ fuel cell was used in Appollo space programme.

109. (i) It is pollution free.

(ii) It has high efficiency of 70 - 75% and its rate can be controlled.

110. Fuel cells : Refer to answer 107.

The reactions taking place in hydrogen - oxygen fuel cell :

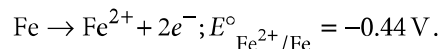


111. Refer to answer 111.

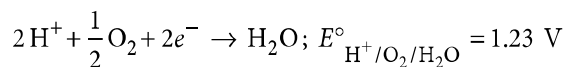
112. Metals of lower electrode potential value when connected with iron protect it from oxidation and prevent corrosion. Hence, coating of metal A having lower electrode potential will be better than B which has higher $E^{\circ}_{(B^{2+}/B)} = -0.14$ V.

113. According to electrochemical theory of rusting the impure iron surface behaves like small electrochemical cell. In this any point of iron acts as anode and other iron surface acts as cathode. Moisture having dissolved CO_2 or O_2 acts as an electrolyte. The reactions are given below.

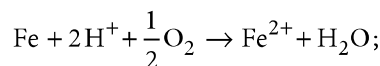
At anode :



At cathode :

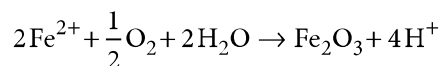


Overall reaction :



$$E^{\circ}_{\text{cell}} = 1.67 \text{ V}$$

The Fe^{2+} ions are further oxidised by atmospheric oxygen to Fe^{3+} ions, which form hydrated ferric oxide (rust).



114. Refer to answer 113.

115. Corrosion is the process of slowly eating away of the metal due to attack of the atmospheric gases on the surface of the metal resulting into the formation of compounds such as oxides, sulphides, carbonates, etc. The corrosion of iron is called rusting.

Electrochemical theory of rusting of iron : Refer to answer 113.

