CHEMISTRY STUDY MATERIALS FOR CLASS 12
(NCERT Based Questions - Answers)
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Electrochemistry

Question 48. Silver is uniformly electro-deposited on a metallic vessel of surface area of 900 cm^2 by passing a current of 0.5 ampere for 2 hours. Calculate the thickness of silver deposited.

[Given: the density of silver is 10.5 g cm^{-3} and atomic mass of Ag = 108 amu.]

Answer: Quantity of electricity passed = 0.5 × 2 × 60 × 60 = 3600 c

$$Ag^+ + e^- \rightarrow Ag$$

96500 c deposits Ag = 107.92 g

$$\therefore \quad \text{3600 c will deposit Ag} = \frac{107.92}{96500} \times 3600$$
$$= 4.026 \text{ g}$$
$$\text{Volume deposited} = \frac{\text{Mass}}{\text{Density}} = \frac{4.026}{10.5}$$
$$= 0.3947 \text{ cc}$$
$$\therefore \quad \text{Thickness deposited} = \frac{\text{Volume}}{\text{Area}} = \frac{0.3947}{900}$$
$$= \mathbf{4.38 \times 10^{-4}\ cm}$$

Question 49.

(a) Calculate $\Delta_r G^0$ for the reaction

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

Given : E^0_{cell} = + 2.71 V, 1 F = 96500 C mol^{-1}

(b) Name the type of cell which was used in Apollo space programme for providing electrical power.

Answer:

(a) $\Delta_r G^0 = -\,nFE^0$

$= -2 \times 96500 \times 2.71$ ($\because$ n = 2)

$= -523{,}030$ J mol^{-1} = -523.03 KJ mol^{-1}

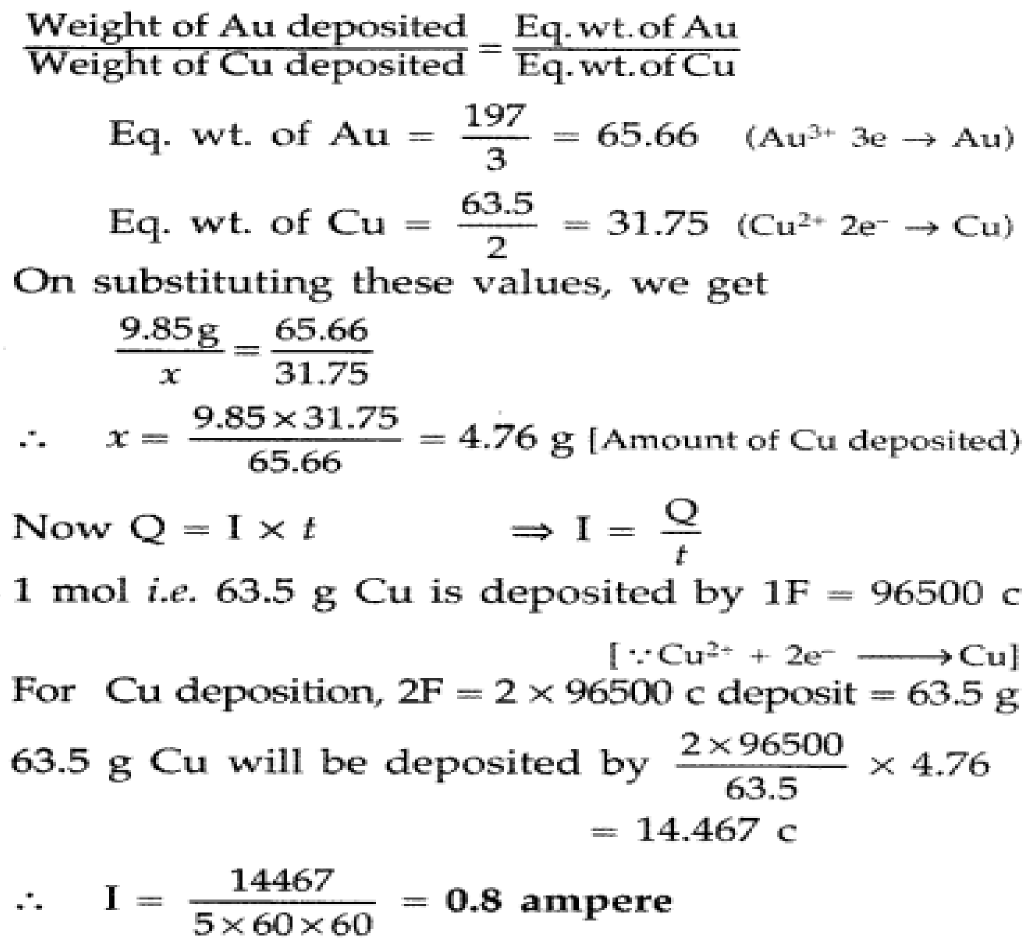
(b) Fuel cell was used in Apollo space programme for providing electrical power.

Question 49.

A current was passed for 5 hours through two electrolytic cells connected in series. The first cell contains $AuCl_3$ and second cell $CuSO_4$ solution. If 9.85 g of gold was deposited in the first cell, what amount of copper gets deposited in the second cell? Also calculate magnitude of current in ampere.

Given: Atomic mass of Au = 197 amu and Cu = 63.5 amu.

Answer:



$$\frac{\text{Weight of Au deposited}}{\text{Weight of Cu deposited}} = \frac{\text{Eq. wt. of Au}}{\text{Eq. wt. of Cu}}$$

$$\text{Eq. wt. of Au} = \frac{197}{3} = 65.66 \quad (Au^{3+}\ 3e \rightarrow Au)$$

$$\text{Eq. wt. of Cu} = \frac{63.5}{2} = 31.75 \quad (Cu^{2+}\ 2e^- \rightarrow Cu)$$

On substituting these values, we get

$$\frac{9.85\,g}{x} = \frac{65.66}{31.75}$$

$$\therefore \quad x = \frac{9.85 \times 31.75}{65.66} = 4.76\ g \text{ [Amount of Cu deposited)}$$

$$\text{Now } Q = I \times t \qquad \Rightarrow I = \frac{Q}{t}$$

1 mol *i.e.* 63.5 g Cu is deposited by 1F = 96500 c

$$[\because Cu^{2+} + 2e^- \longrightarrow Cu]$$

For Cu deposition, 2F = 2 × 96500 c deposit = 63.5 g

$$\text{63.5 g Cu will be deposited by } \frac{2 \times 96500}{63.5} \times 4.76$$

$$= 14.467\ c$$

$$\therefore \quad I = \frac{14467}{5 \times 60 \times 60} = \mathbf{0.8\ ampere}$$

Question 51.

The resistance of 0.01 M NaCl solution at 25° C is 200 Ω. The cell constant of the conductivity cell used is unity. Calculate the molar conductivity of the solution.

Answer:

For 0.01 M NaCl solution,

R = 200 Ω, cell constant is unity.

∴ Conductivity (K) = Cell constant Resistance = 1200 = 0.005 Sm^{-1}

Concentration of solution = 0.01 M = 0.01 mol L^{-1}

$= 0.01 \times 10^3$ mol m^{-3} = 10 mol m^{-3}

Molar conductivity = KCm=0.005105 × 10^{-4} Sm^2 mol^{-1}

Question 52.

Calculate emf of the following cell at 25°C :

Fe | Fe^{2+} (0.001 M) || H^+ (0.01 M) | H_2(g) (1 bar) | Pt(s)

$E^0(Fe^{2+} | Fe) = -0.44$ V $E^0(H^+ | H_2) = 0.00$V (Delhi 2015)

Answer:

Fe | Fe^{2+} (0.001 M) || H^+ (0.01 M) | H_2(g) (1 bar) | Pt(s)

$$\begin{aligned} E^0_{cell} &= E^0_c - E^0_a \\ &= [0 - (-0.44)]\ V = 0.44\ V \end{aligned}$$

$$E_{cell} = E^0_{cell} - \frac{0.0591}{2}\log\frac{[Fe^{+2}]}{[H^+]^2}$$

$$\Rightarrow \quad E_{cell} = 0.44\ V - \frac{0.0591}{2}\log\frac{0.001}{0.01 \times 0.01}$$

$$= 0.44\ V - \frac{0.0591}{2}\ \log 10$$

$$= 0.44\ V - \frac{0.0591}{2} = (0.44 - 0.0295)\ \text{Volts}$$

= 0.4105 Volts

Question 53. Conductivity of 2.5×10^{-4} M methanoic acid is 5.25×10^{-5} S cm^{-1}.

Calculate its molar conductivity and degree of dissociation.

Given : $\lambda^0(H^+) = 349.5$ Scm^2 $mol^{-1}$and $\lambda^0(HCOO^-) = 50.5$ Scm^2 mol^{-1}.

Answer:

Concentration is 2.5×10^{-4} M

$K = 5.25 \times 10^{-5}$ Scm^{-1}.

$\Lambda cm = K \times 1000$ Concentration

$$= \frac{5.25 \times 10^{-5} \times 1000}{2.5 \times 10^{-4}}\ Scm^2\ mol^{-1}$$

$$= \frac{5.25 \times 10^2}{2.5} = \frac{525}{2.5} = 210\ Scm^2\ mol^{-1}$$

Molar conductivity at infinite dilution,

$$\begin{aligned} \Lambda^0_m &= \lambda^0 H^+ + \lambda^0 HCOO^- = (349.5 + 50.5) \\ &= 400\ Scm^2\ mol^{-1} \end{aligned}$$

Degree of dissociation,

$$\alpha = \frac{\Lambda^c_m}{\Lambda^0_m} = \frac{210}{400} = \frac{21}{40} = 0.525$$

Question 54.

Calculate e.m.f. of the following cell at 298 K: $2Cr(s) + 3Fe^{2+}$ (0.1 M) →

$2Cr^{3+}$ (0.01 M) + 3 Fe(s)

Given: $E^0(Cr^{3+}| Cr) = -0.74$ V $E^0(Fe^{2+} | Fe) = -0.44$ V (Delhi 2016)

Answer:

Cell reaction: $2Cr(s) + 3Fe^{2+}$ (0.1 M) → $2Cr^{3+}$ (0.01 M) + 3Fe(s)

Given: $E^0(Cr^{3+}/Cr) = -0.74$

$E^0(Fe^{2+}/Fe) = -0.44$ V

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\begin{aligned}
E_{cell} &= ? \\
E^0_{cell} &= E^0_{cathode} - E^0_{anode} \\
&= (-0.44) - (-0.74) \\
&= -0.44 + 0.74 = +0.30 \text{ V}
\end{aligned}
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Using Nernst equation:

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\begin{aligned}
E_{cell} &= E^0_{cell} - \frac{0.0591}{n} \log \frac{[Cr^{3+}]^2}{[Fe^{2+}]^3} \\
&= 0.30 - \frac{0.0591}{6} \log \frac{(0.01)^2}{(0.1)^3} \\
&= 0.30 - 0.00985 \log \frac{0.0001}{0.001} \\
&= 0.30 - 0.00985 \log \left(\frac{1}{10}\right) \\
&= 0.30 - 0.00985 (\log 1 - \log 10) \\
&= 0.30 - 0.00985 (0 - 1) = 0.30 + 0.00985
\end{aligned}
$$

$$\therefore \quad E_{cell} = \mathbf{0.3098 \text{ V}}$$

Question 55.(i) Calculate the mass of Ag deposited at cathode when a current of 2 amperes was passed through a solution of $AgNO_3$ for 15 minutes.

[Given: Molar mass of Ag = 108 g mol^{-1} 1F = 96,500 C mol^{-1})

(ii) Define fuel cell.

Answer:

(i) Q = I × t …(Charge = Current ∝ Time). = 2 × 15 × 60 = 1800 C

∵ 96500 C deposit Ag = 108 g

∴ 1800 C deposit Ag = 10896500 × 1800 = 2.0145 g

(ii) Cells that convert the energy of combustion of fuels like hydrogen, methanol, methane, etc directly into electrical energy are called fuel cells.

Question 56.(a) The cell in which the following reaction occurs:

$$2Fe^{3+} (aq) + 2I^-(aq) \rightarrow 2Fe^{2+} (aq) + I_2(s)$$

has Eo Cell = 0.236 V at 298 K. Calculate the standard Gibbs energy of the cell reaction. (Given: 1F = 96,500 C mol^{-1})

(b) How many electrons flow through a metallic wire if a current of 0.5 A is passed for 2 hours? (Given: 1F = 96,500 C mol^{-1})

Answer:

(a) $2Fe^{3+}$ (aq) + $2I^{-}$ (aq) → $2Fe^{2+}$ (aq) + I_2 (s)

For the given reaction, $n = 2$, $E° = 0.236$ V

Using formula, $\Delta G° = -nF\,E°_{cell} = -2 \times 96{,}500$ C $mol^{-1} \times 0.236$ V

$\therefore \Delta G° = -45.55$ kj mol^{-1}

(b) Given:

$I = 0.5$ A

$t = 2$ hrs. $= 2 \times 60 \times 60$ s $= 7{,}200$ s

$Q = I \times t = 0.5 \times 7200 = 3{,}600$ C

96,500 C electricity flows to produce = 6.022×10^{23} electrons

∴ 1 C electricity flows to produce

$$= \frac{6.022 \times 10^{23}}{96,500}$$

∴ 3,600 C electricity flows to produce

$$= \frac{6.022 \times 10^{23}}{96,500} \times 3,600$$

$$= \mathbf{2.246 \times 10^{22}} \textbf{ electrons}$$

Question 57. Why electrochemical cells stop working after some time? The reduction potential of an electrode depends upon the concentration of solution with which it is in contact.

Answer:

As the cell works, the concentration of reactants decrease. Then according to Le Chatelier's principle it will shift the equilibrium in backward direction. On the other hand if the concentration is more on the reactant side then it will shift the equilibrium in forward direction. When cell works concentration in anodic compartment in cathodic compartment decreases and hence E° cathode will decrease. Now EMF of cell is

$E^0_{cell} = E^0_{cathode} - E^0_{anode}$

A decrease in E° cathode and a corresponding increase in E° anode will mean that EMF of the cell will decrease and will ultimately become zero i.e., cell slops working after some time.

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