Short Answer Questions-II

Short Answer Questions-II (PYQ)

Q.1.Answer the following questions

[CBSE (AI) 2017]

Q. The cell in which the following reaction occurs:

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

has $E^0_{\text{cell}} = 0.236$ **V at 298 K. Calculate the standard Gibbs energy of the cell reaction.**

(Given : 1 F = 96,500 C mol^{-1})

Ans.

Given, n = 2, $E^0_{\text{cell}} = 0.236$ V, F = 96,500 C mol^{-1}

$$\therefore \quad \Delta G^\circ = -\mathrm{nFE}^o_{\text{cell}}$$

$$= -2 \times 96500\ \mathrm{C\ mol^{-1}} \times 0.236\ \mathrm{V} = -45548\ \mathrm{J\ mol^{-1}} = \mathbf{-45.548\ kJ\ mol^{-1}}$$

Q. How many electrons flow through a metallic wire if a current of 0.5A is passed for 2 hours?

(Given : 1 F = 96,500 C mol^{-1})

[*CBSE (AI) 2017*]

Ans. $Q = I \times t = 0.5$ A × 2 × 60 × 60 s = 3600 C

Number of electrons flowing through the wire on passing a charge of 96500 C = 6.022 × 10^{23}

∴ Number of electrons flowing through the wire on passing a charge of 3600 C

$$= \frac{6.022 \times 10^{23} \times 3600\ C}{96500\ C} = 2.246 \times 10^{22}\ \text{electrons}$$

Q.2.

i. Write two advantages of $H_2 - O_2$ fuel cell over ordinary cell.

ii. Equilibrium constant (K_c) for the given cell reaction is 10. Calculate E^o_{cell}.

$$A(s) + B^{2+}(aq) \rightleftharpoons A^{2+}(aq) + B(s)$$

[*CBSE (F) 2014*]

Ans. (i) The two main advantages of H_2–O_2 fuel cell over ordinary cell are as follows:

• It has high efficiency of 60%–70%.

• It does not cause any pollution.

ii. $A(s) + B^{2+}(aq) \rightleftharpoons A^{2+}(aq) + B(s)$

Here, $n = 2,\ K_c = 10$

$$E^{o}{}_{cell} = \frac{0.059}{n} \times \log K_c$$

$$E^{o}{}_{cell} = \frac{0.059}{n} \times \log 10 = \times 1 = 0.0295 \text{ V}$$

Q.3. 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: $\lambda^o(H^+)$ = 349.5 S cm² mol⁻¹ and $\lambda^o(HCOO^-)$ = 50.5 S cm² mol⁻¹.

[*CBSE Allahabad 2015*]

Ans.

$$= \Lambda^c_m \frac{k \times 1000}{M}$$

$$\Lambda^c_m = \frac{5.25 \times 10^{-5}\ S\ cm^{-1} \times 1000\ cm^3\ L^{-1}}{2.5 \times 10^{-4}\ mol\ L^{-1}}$$

$$\Lambda^c_m = 210 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\Lambda_m^c\ (HCOOH) = \Lambda^o_{HCOO^-} + \Lambda^o_{H^+}$$

$$= 50.5\ S\ cm^2\ mol^{-1} + 349.5\ S\ cm^2\ mol^{-1}$$

$$= 400\ S\ cm^2\ mol^{-1}$$

Degree of dissociation, $\alpha = \frac{\Lambda_m^c}{\Lambda_m^o} = \frac{210\ S\ cm^2\ mol^{-1}}{400\ S\ cm^2\ mol^{-1}} = 0.525$

or α = **52.5%**

Q.4. Answer the following questions

[CBSE (AI) 2014]

Calculate $\Delta_r G^\circ$ for the reaction

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

Given: E^o_{cell} = +2.71 V, 1 F = 95600 C mol^{-1}

Ans.

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

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

Here, $n = 2$, $F = 96500\ C\ mol^{-1}$ and $E^o_{cell} = 2.71\ V$

$\therefore$

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

$$= -\ 523030\ J\ mol^{-1} = \mathbf{-\ 523.03\ kJ\ mol^{-1}}$$

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

Ans. H_2—O_2 fuel cell

Q.5. Conductivity of 0.00241 M acetic acid solution is 7.896 × 10^{-5} S cm^{-1}. Calculate its molar conductivity in this solution. If Λ_m^o for acetic acid is 390.5 S cm^2 mol^{-1}, what would be its dissociation constant?

[*CBSE Delhi 2008*]

Ans.

$c = 0.00241\ M,\ k = 7.896 \times 10^{-5}\ S\ \ cm^{-1}\ ,\ \Lambda_m^o = 390.5\ S\ \ cm^2\ \ mol^{-1}$

$\Lambda_m = \frac{k \times 1000}{c}$

Substituting the values

$\Lambda_m = \frac{7.896 \times 10^{-5} \times 1000}{0.002} = 32.76\ S\ \ cm^2\ \ mol^{-1}$

$\alpha = \frac{\Lambda_m^c}{\Lambda_m^o} = \frac{32.76}{390.5} = 0.084$

α = 8.4%

$$\begin{array}{lccc} CH_3\,COOH & \rightleftharpoons & CH_3\,COO^- & + \quad H^+ \\ c & & 0 & 0 \\ c(1-\alpha) & & c\alpha & c\alpha \end{array}$$

$K_a = \frac{[CH_3 - COO^-]\ [H^+]}{[CH_3 - COOH]} = \frac{c\alpha\ .\ c\alpha}{c(1-\alpha)} = \frac{c\alpha^2}{1-\alpha}$

$K_a = \frac{0.0241\ \ (0.084)^2}{(1-0.084)} = 1.86\ \times\ 10^{-5}$

Q.6.

alculate $\Delta r\ G^\circ$ and log K_c for the following reaction at 298 K:

$2Cr(s) + 3Fe^{2+}(aq)\ \ \rightarrow\ \ 2Cr^{3+}(aq) + 3Fe(s)$

Given: E^o_{cell} = 0.30 V

[*CBSE (F) 2016*]

Ans.

$$
\begin{array}{lll}
Cr\,(s) & \rightarrow & Cr^{3+}\,(aq) + 3e^{-}\,]\times 2 \\
Fe^{2+}\,(aq) + 2e^{-} & \rightarrow & Fe\,(s) \qquad]\times 3 \\
\hline
2\,Cr\,(s) + 3\,Fe^{2+}\,(aq) & \rightarrow & 2\,Cr^{3+}\,(aq) + 3\,Fe\,(s), \quad n = 6 \\
\hline
\end{array}
$$

Here, $n = 6$, $E^{o}_{cell} = 0.30$ V

Substituting the values in the expression, $\log K_c = \frac{n}{0.059} E^{o}_{cell}$, we get

$\log K_c = \frac{6}{0.059} \times 0.30$ or $\log K_c = 30.5084$

$K_c = \text{Antilog}\,(30.5084)$ or $K_c = 3.224 \times 10^{30}$

$\Delta G^{\circ} = -\,nFE^{o}_{cell}$

$= -\,6 \times 96500 \times 0.30 = -\,173700\ \text{J mol}^{-1}$

$\Delta G^{\circ} = \mathbf{-\,173.7\ kJ\ mol^{-1}}$

Q.7. One half-cell in a voltaic cell is constructed from a silver wire dipped in silver nitrate solution of unknown concentration. The other half-cell consists of a zinc electrode in a 0.10 M solution of $Zn(NO_3)_2$. A voltage of 1.48 V is measured for this cell. Use this information to calculate the concentration of silver ions in the solution.

$\left(\textbf{Given: } = E^{o}_{Zn^{2+}/Zn}\ -0.763\ \text{V},\ E^{o}_{Ag^{+}/Ag} = +\,0.80\ \text{V}\right)$

[*CBSE (F) 2010*]

Ans.

Electrochemical cell

$$Zn(s) \mid Zn^{2+}\ (0.10\ M) \mid\mid Ag^{+}\ (conc.) \mid Ag(s)$$

$$E^{o}_{cell} = E^{o}_{R} - E^{o}_{L} = E^{o}_{Ag^{+}/Ag} - E^{o}_{Zn^{2+}/Zn}$$

$$= 0.80\ V - (-\ 0.763)V = 1.563\ V$$

We know that,

$$E_{cell} = E^{o}_{cell} - \frac{0.0591}{n}\log\frac{[Zn^{2+}]}{[Ag^{+}]^{2}}$$

$$1.48 = 1.563 - \frac{0.0591}{2}\log\frac{[0.10]}{[Ag^{+}]^{2}}$$

$$\log = \frac{[0.10]}{[Ag^{+}]^{2}} = \frac{0.083}{0.02955}\ 2.8087$$

Or $$\frac{[0.10]}{[Ag^{+}]^{2}} = \text{antilog } 2.8087 = 643.7$$

$$[Ag^{+}]^{2} = \frac{0.10}{643.7} = 1.553 \times 10^{-4}$$

$$[Ag^{+}] = \mathbf{1.247 \times 10^{-2}\ M}$$

Q.8. A strip of nickel metal is placed in a 1-molar solution of $Ni(NO_3)_2$ and a strip of silver metal is placed in a 1-molar solution of $AgNO_3$. An electrochemical cell is created when the two solutions are connected by a salt bridge and the two strips are connected by wires to a voltmeter.

i. **Write the balanced equation for the overall reaction occurring in the cell and calculate the cell potential.**
ii. **Calculate the cell potential, *E*, at 25°C for the cell if the initial concentration of $Ni(NO_3)_2$ is 0.100 molar and the initial concentration of $AgNO_3$ is 1.00 molar.**

$$[\ E^o_{Ni^{2+}/Ni} = -0.25\ V;\ E^o_{Ag^+/Ag} = 0.80\ V,\ \log 10^{-1} = -1]$$

[CBSE (F) 2012]

Ans.

i. At anode: $Ni \rightarrow Ni^{2+} + 2e^-$

At cathode: $[Ag^+ + e^- \rightarrow Ag] \times 2$

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

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

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

$E^o{}_{cell} = \mathbf{1.05\ V}$

ii. $E_{cell} = E^o_{cell} - \frac{0.059}{n} \log \frac{[Ni^{2+}]}{[Ag^+]^2}$

Here, $n = 2$, $E^o{}_{cell} = 1.05\ V$, $[Ni^{2+}] = 0.1\ M$, $[Ag^+] = 1.0\ M$

$E_{cell} = 1.05\ V - \frac{0.059}{2} \log \frac{(0.1)}{(1)^2}$

$E_{cell} = 1.05V - 0.0295 \log 10^{-1}$

$= 1.05 + 0.0295\ V = \mathbf{1.0795\ V}$

Q.9. The electrical resistance of a column of 0.05 M NaOH solution of diameter 1 cm and length 50 cm is 5.55 × 10^3ohm. Calculate its resistivity, conductivity and molar conductivity.

[*CBSE (AI) 2012*]

Ans.

$$A = \pi r^2 = 3.14 \times \left(\tfrac{1}{2}\text{cm}\right)^2 = 0.785\ \text{cm}^2\ ;\ l = 50\ \text{cm}$$

$$\text{Resistivity, } \rho = \frac{R \times A}{l} = \frac{5.55 \times 10^3 \text{ohm} \times 0.785\ \text{cm}^2}{50\ \text{cm}} = 87.135\ \text{ohm cm}$$

$$\text{Conductivity, } k = \frac{1}{\rho} = \frac{1}{87.135\ \text{ohm cm}} = 0.01148\ S\ \text{cm}^{-1}$$

$$\text{Molar Conductivity, } \Lambda_m = \frac{k \times 1000}{M} = \frac{0.01148\ S\ \text{cm}^{-1} \times 1000\ \text{cm}^3\ L^{-1}}{0.05\ \text{mol}\ L^{-1}}$$

$$= 229.6\ S\ \text{cm}^2\ \text{mol}^{-1}$$

Q.10. A voltaic cell is set up at 25°C with the following half cells:

Al/Al^{3+} (0.001 M) and Ni/Ni^{2+} (0.50 M)

Write an equation for the reaction that occurs when the cell generates an electric current and determine the cell potential.

$$E^o_{Ni^{2+}/Ni} = -0.25\text{ V};\ E^o_{Al^{3+}/Al} = -1.66\text{ V}\ (\log 8 \times 10^{-6} = -0.54)$$

[*CBSE (AI) 2012*]

Ans.

At anode: $[Al \rightarrow Al^{3+} + 3e^-] \times 2$

At cathode: $[Ni^{2+} + 2e^- \rightarrow Ni] \times 3$

Cell reaction: $2Al + 3Ni^{2+} \rightarrow 2Al^{3+} + 3Ni$

$$E^o_{cell} = E^o_{cathode} - E^o_{anode} = E^o_{Ni^{2+}/Ni} - E^o_{Al^{3+}/Al}$$

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

$$[Al^{3+}] = 1 \times 10^{-3}\text{ M};\ [Ni^{2+}] = 0.5\text{ M};\ n = 6$$

Substituting the values in the Nernst equation,

$$E_{cell} = E^o_{Ni^{2+}/Ni} - E^o_{Al^{3+}/Al}$$

$$E_{cell} = 1.41\text{ V} - \frac{0.059}{6}\log\frac{(10^{-3})^2}{(0.5)^3}$$

$$= 1.41\text{ V} - \frac{0.059}{6}\log(8 \times 10^{-6}) = 1.41\text{ V} - \frac{0.059}{6}(-0.54)$$

$$E_{cell} = 1.41\text{ V} + 0.0053\text{ V} = \mathbf{1.4153\text{ V}}$$

Short Answer Questions-II (OIQ)

Q.1. Chromium metal can be plated out from an acidic solution containing CrO_3 according to the following equation:

$CrO_3(aq) + 6H^+(aq) + 6e^- \rightarrow Cr(s) + 3H_2O$

Calculate (*i*) how many grams of chromium will be plated out by 24,000 coulombs and (*ii*) how long will it take to plate out 1.5 g of chromium by using 12.5 A current?

(At. mass of Cr = 52).

Ans.

i. 6×96500 coulomb deposit Cr = 1 mole = 52 g

$\therefore$ 24,000 coulomb deposit Cr $= \frac{52 \times 24000}{6 \times 96500}$ g $= \mathbf{2.1554\ g}$

ii. 52 g of Cr is deposited by electricity $= 6 \times 96500$ C

$\therefore$ 1.5 g require electricity $= \times 1.5$ C $= 16071.9$ C

$\therefore$ Time for which the current is required to be passed $= \frac{16071.9\ \text{C}}{12.5\ \text{A}} = \mathbf{1336\ s}$.

Q.2. Answer the following questions

Q. A current of 1.50 A was passed through an electrolytic cell containing $AgNO_3$ solution with inert electrodes. The weight of Ag deposited was 1.50 g. How long did the current flow?

Ans.

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

Quantity of charge required to deposit 108 g of silver = 96500 C

$\therefore$ Quantity of charge required to deposit 1.50 g of silver $= \frac{96500}{108} \times 1.50 = 1340.28$ C

$$t = \frac{Q}{I}$$

$\therefore$ Time taken $= \frac{1340.28}{1.50} = \mathbf{893.52\ s}$

Q. Write the reactions taking place at the anode and cathode in the above cell.

Ans.

Inert electrodes

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

Cathode: $Ag^+(aq) + e^- \rightarrow Ag(s)$

Q. Give reactions taking place at the two electrodes if these are made up of Ag.

Ans.

Ag electrodes

Anode: $Ag(s) \rightarrow Ag^+(aq) + e^-$

Cathode: $Ag^+(aq) + e^- \rightarrow Ag(s)$

Q.3. A zinc rod is dipped in 0.1 M solution of $ZnSO_4$. The salt is 95% dissociated at this dilution at 298 K. Calculate the electrode potential $\left(E^o_{Zn^{2+}/Zn} = -0.76\ V\right)$.

Ans.

The electrode reaction written as reduction reaction is

$$Zn^{2+} + 2e^- \rightarrow Zn \ (n = 2)$$

Applying Nernst equation, we get

$$E_{Zn^{2+}/Zn} = E^o_{Zn^{2+}/Zn} - \frac{0.0591}{2}\log\frac{1}{[Zn^{2+}]}$$

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

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

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

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

$$= -0.76 - 0.02955(3 - 1.9777) = -0.76 - 0.03021 = \mathbf{-0.79021\ V}$$

Q.4. The conductivity of 0.001028 mol L^{-1} acetic acid is 4.95×10^{-5} S cm^{-1}. Calculate its dissociation constant Λ^o_m if for acetic acid is 390.5 S cm^2 mol^{-1}.

[*HOTS*]

Ans.

$$\Lambda_m = \frac{k \times 1000}{c} = \frac{4.95 \times 10^{-5}\ S\ cm^{-1}}{0.001028\ mol\ L^{-1}} \times \frac{1000\ cm^3}{L} = 48.15 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\alpha = \frac{\Lambda_m}{\Lambda^o_m} = \frac{48.15\ S\ cm^2\ mol^{-1}}{390.5\ S\ cm^2\ mol^{-1}} = 0.1233$$

$$K = \frac{c\alpha^2}{(1-\alpha)} = \frac{0.001028\ mol\ L^{-1} \times (0.1233)^2}{1 - 0.1233} = \mathbf{1.78 \times 10^{-5}\ mol\ L^{-1}}$$

Q.5. Estimate the minimum potential difference needed to reduce Al_2O_3 at 500°C. The free energy change for the decomposition reaction

$\frac{2}{3}$ Al_2O_3 $\quad \frac{4}{3}$ Al + O2 is 960 kJ (F = 96,500 C mol−1). [HOTS]

Ans.

$$Al_2O_3(2Al^{3+} + 3O^{2-}) \rightarrow 2Al + \frac{3}{2}O_2,\ n = 6e^-$$

$$\therefore \quad \frac{2}{3}Al_2O_3 \rightarrow \frac{4}{3}Al + O_2,\ n = \frac{2}{3} \times 6e^- = 4e^-$$

$$\Delta_r G = 960 \times 1000 = 960000 \text{ J}$$

Now, $\Delta_r G = -\, nFE_{cell}$

$\Rightarrow$ $E_{cell} = -\dfrac{\Delta_r G}{nF} = \dfrac{-960000}{4 \times 96500}$

$\Rightarrow$ $E_{cell} = \mathbf{-\ 2.487\ V}$

$\therefore$ Minimum potential difference needed to reduce Al_2O_3 is **– 2.487 V.**

Q.6.

The emf of a cell corresponding to the reaction.

$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(0.1\ M) + H_2\ (g,\ 1\ atm)$ is 0.28 volt at 25°C.

Write the half-cell reactions and calculate the pH of the solution at the hydrogen electrode.

$E^o_{Zn^{2+}/Zn} = -\ 0.76\ V,\ E^o_{H^+/H_2} = 0\ V$ [*HOTS*]

Ans.

Half-cell reactions:

At anode: $Zn \rightarrow Zn^{2+} + 2e^-$

At cathode: $2H^+ + 2e^- \rightarrow H_2$

Cell reaction: $\overline{\underline{Zn + 2H^+ \rightarrow Zn^{2+} + H_2}}$

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

$$= \left(E^o_{H^+/H_2} - E^o_{Zn^{2+}/Zn}\right) - \frac{0.0591}{2}\log\frac{0.1}{[H^+]^2}$$

$$= [0 - (-\ 0.76)] - 0.02955\ [\log 10^{-1} - 2\log(H^+)]$$

$0.28 = 0.76 - 0.02955\ (-1 + 2\ pH)$ $\quad [\because pH = -\log(H^+)]$

$2\ pH - 1 = 16.244$

$pH = 8.62$

Q.7.

From the following molar conductivities at infinite dilution, calculate Λ_m^0 for NH_4OH.

Λ_m^0 for $Ba(OH)_2 = 457.6\ \Omega^{-1}\ cm^2\ mol^{-1}$

Λ_m^0 for $BaCl_2 = 240.6\ \Omega^{-1}\ cm^2\ mol^{-1}$

Λ_m^0 for $NH_4Cl = 129.8\ \Omega^{-1}\ cm^2\ mol^{-1}$

Ans.

$$\Lambda^o_{m(NH_4OH)} = \lambda^o_{NH_4^+} + \lambda^o_{OH^-}$$

$$= (\lambda^o_{NH_4^+} + \lambda^o_{Cl^-}) + \frac{1}{2}(\lambda^o_{Ba^{2+}} + 2\lambda^o_{OH^-}) - \frac{1}{2}(\lambda^o_{Ba^{2+}} + 2\lambda^o_{Cl^-})$$

$$= \Lambda^o_{m(NH_4Cl)} + \frac{1}{2}[\Lambda^o_{m(Ba(OH)_2)}] - \frac{1}{2}[\Lambda^o_{m(BaCl_2)}]$$

$$129.8 + \frac{1}{2} \times 457.6 - \frac{1}{2} \times 240.6 = 238.3\ ohm^{-1}\ cm^2\ mol^{-1}$$