Electrochemistry

Question1

Electrolysis of aqueous copper (II) sulphate between Pt electrodes gives ' X' at anode and ' Y' at cathode. X and Y are respectively.

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Options:

A.

$\mathrm{Cu}, \mathrm{O}_2$

B.

$\mathrm{O}_2, \mathrm{Cu}$

C.

$\mathrm{SO}_2, \mathrm{H}_2$

D.

$\mathrm{O}_2, \mathrm{H}_2$

Answer: B

Solution:

Reaction at anode is,

$2\mathrm{H}_2\mathrm{O}(l) \longrightarrow \mathrm{O}_2(g)+4\mathrm{H}^{+}(aq)+4e^{-}$

This produce gas $=\mathrm{O}_2(X)$

Reaction at cathode

$\mathrm{Cu}^{2+}+2e^{-} \longrightarrow \mathrm{Cu}(s)$

Metal deposited $=\mathrm{Cu}(Y)$

Question2

At 298 K , if emf of the cell corresponding to the reaction $Zn(s) + 2H^+(aq) \longrightarrow Zn^{2+}(0.01M) + H_2(g)(1\ atm)$ is 0.28 V , then the pH of the solution at the hydrogen electrode is $\left(\frac{2.303RT}{F} = 0.06\ V\right), \left(E^\circ_{Zn^{2+}/Zn} = -0.76\ V\right)$

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Options:

A.

8

B.

7

C.

9

D.

10

Answer: C

Solution:

The cell reaction is,

At anode $(Zn), Zn \longrightarrow Zn^{2+} + 2e^-$

Cathode $(H^+), 2H^+ + 2e^- \longrightarrow H_2$

Using Nernst equation,

$$E_{cell} = E^\circ_{cathode} - E^\circ_{anode} - \frac{2.303RT}{F} \log Q$$

$$0.28 = +0.76 - \frac{0.06}{2} \log\left(\frac{0.01}{[H^+]^2}\right)$$

Solving this,

$$\left[H^{+}\right]=10^{-9}$$
$$pH=-\log\left[H^{+}\right]=9$$

Question3

0.592 g of copper is deposited in 60 minutes by passing

0.5 A current through a solution of copper (II) sulphate. The electro chemical equivalent of copper (II) (in gC^{-1}) is

($F=96500 \mathrm{Cmol}^{-1}$)

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Options:

A. 3.3×10^{-3}

B. 3.3×10^{-4}

C. 6.6×10^{-3}

D. 6.6×10^{-4}

Answer: B

Solution:

Given, Mass of copper $=0.592$ g, time $=60$ minutes, current $=0.5$ A Total charge passes through solution,

$Q=I \times t$

$Q=0.5 \times 60 \times 60=1800 \mathrm{C}$

Number of moles of electrons transferred.

$n=\frac{Q}{F}=\frac{1800 \mathrm{C}}{96500 \mathrm{C} / \mathrm{mol}}$

$\Rightarrow \quad n=0.01865 \mathrm{~mol}$

Cu^{2+} requires 2 moles of electron to deposit 1 mole of copper

$$\text{Moles of copper } = \frac{n}{2} = \frac{0.01865}{2}$$

$$= 0.009325 \text{ mol}$$

Mass of copper deposited

$= 0.009325 \times 63.55$

$= 0.592$ g

Electrochemical equivalent is given by

$$\frac{\text{Mass of copper deposited}}{\text{Charged passed}} = \frac{0.592}{180}$$

$= 3.3 \times 10^{-4}$ g/C

Question4

The standard electrode potentials $E^\circ(\mathrm{V})$ for $\mathrm{Li}^+/\mathrm{Li}, \mathrm{Na}^+/\mathrm{Na}$ respectively are

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Options:

A. $-3.04, -2.714$

B. $-2.714. \quad -3.04$

C. $-3.04, -3.04$

D. $-2.714, -2.714$

Answer: A

Solution:

The standard electrode potentials for the $\mathrm{Li}^+/\mathrm{Li}$ and $\mathrm{Na}^+/\mathrm{Na}$ half-cells are as follows:

$\mathrm{Li}^+/\mathrm{Li}$: $E^\circ = -3.04$ V

$\mathrm{Na}^+/\mathrm{Na}$: $E^\circ = -2.714$ V

Therefore, the correct option is:

Option A

$-3.04, -2.714$

These values indicate the tendency of lithium and sodium to lose electrons and form cations, with lithium being more reactive in this context due to its more negative electrode potential.

Question5

Two statements are given below.

Statement I : Molten NaCl is electrolysed using Pt electrodes. Cl_2 is liberated at anode.

Statement II : Aqueous $CuSO_4$ is electrolysed using Pt electrodes. O_2 is liberated at cathode.

The correct answer is

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Options:

A. Both statements I and II are correct.

B. Both statements I and II are not correct.

C. Statement I is correct but statement II is not correct.

D. Statement I is not correct but statement II is correct.

Answer: C

Solution:

Statement I is correct but statement II is incorrect. Correct form of Statement II is

At cathode (Reduction take place)

$2Cu^{2+} + 4e^{-} \rightarrow 2Cu$

At anode (oxidation)

$4OH^{-} \longrightarrow 2H_2O + O_2 + 4e^{-}$

Hence, oxygen is liberated at anode not cathode.

Question6

The molar conductivity of 0.02 M solution of an electrolyte is $124 \times 10^{-4}\ \mathrm{S\ m^2\ mol^{-1}}$.What is the resistance of same solution (in ohms), kept in a cell of cell constant $129\ \mathrm{m}^{-1}$?

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Options:

A. 390

B. 260

C. 130

D. 520

Answer: D

Solution:

Given,

Molar conductivity

$= 124 \times 10^{-4} \mathrm{Sm}^2/\mathrm{mol}$

Molarity $= 0.02\mathrm{M}$,

Cell constant $= 129\ \mathrm{m}^{-1}$

Molar conductivity is given by

$\Lambda_{\mathrm{m}} = \frac{1000 \times \kappa}{M}$

where κ is conductivity

$\mathrm{K} = \frac{\Lambda_{\mathrm{m}} \times M}{1000}$

Substitute the vaues in Eq. (i),

$\kappa = \frac{124 \times 10^{-4} \times 0.02}{1000} \Rightarrow \kappa = 0.248 \ldots$

Conductivity, k is given by

$$K = \frac{\text{Cell constant}}{\text{Resistance}}$$

$$\text{or Resistance} = \frac{\text{Cell constant}}{K}$$

$$R = \frac{129}{0.248} \text{ or } R = 520\Omega$$

Question7

If the degree of dissociation of formic acid is 11.0%, the molar conductivity of 0.02 M solution of it is(Given, $\lambda^\circ\left(H^+\right) = 349.6 \text{ S cm}^2 \text{ mol}^{-1}$ $\lambda^\circ\left(HCOO^-\right) = 54.6 \text{ S cm}^2 \text{ mol}^{-1}$)

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Options:

A. $44.46 \text{ S m}^2 \text{ mol}^{-1}$

B. $44.46 \text{Sam}^2 \text{ mol}^{-1}$

C. $22.23 \text{ S m}^2 \text{ mol}^{-1}$

D. $22.23 \text{Sam}^2 \text{ mol}^{-1}$

Answer: B

Solution:

Using Kohlrausch's law we get, (for formic acid)

$$\left.\lambda^\circ_{(HCOOH)} = \lambda^\circ_{(H^+)} + \lambda^\circ_{(HCOO)}\right)$$

$$\lambda^\circ_{(HCOOH)} = 349.6 + 54.6$$

$$\lambda^\circ_{(HCOOH)} = 404.2 \text{ S cm}^2/\text{mol}$$

Degree of dissociation is given by

$$\alpha = \frac{\lambda_{(HCOOH)}}{\lambda^\circ_{(HCOOH)}}$$

[where, $\alpha = 0.11$ (as 11% given)]

$\lambda_{(HCOOH)}=$ Molar conductivity of HCOOH

$$\lambda_{(HCOOH)}=\alpha\times\lambda^{\circ}_{(HCOOH)}$$

$$=0.11\times 404:2$$

$$=44.46\ \mathrm{S\ cm^2/mol}$$

Question8

Identify the correct statements from the following

(A) At 298 K , the potential of hydrogen electrotle placed in a solution of $\mathrm{pH}=10$, is -0.59 V

(B) The limiting molar conductivity of Ca^{2+} and $\mathrm{Cl}^{-}$is 119 and $76\ \mathrm{S\ cm^2\ mol^{-1}}$ respectively. The limiting molar conductivity of $\mathrm{CaCl_2}$ is $195\mathrm{Scm^2\ mol^{-1}}$

(C) The correct relationship between K_C and $E^{\ominus}_{\text{cell}}$ is
$E^{\theta}_{\text{cell}}=\frac{2303RT}{nF}\log K_C$

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Options:

A. A, B, C

B. A, B only

C. A, C only

D. B, C only

Answer: C

Solution:

Correct Statements

To evaluate the given statements, we analyze each one:

Statement (A): The potential of the hydrogen electrode in a solution of $\mathrm{pH} = 10$ at 298 K.

The potential of a hydrogen electrode can be calculated using the Nernst equation for the hydrogen electrode:

$$E = E^{\ominus} - \frac{0.0591}{n}\log[\mathrm{H}^{+}]$$

At standard conditions $E^{\ominus} = 0$, and $[\mathrm{H}^{+}]$ is calculated from the pH:

$$\mathrm{pH} = 10 \implies [\mathrm{H}^{+}] = 10^{-10}$$

Substitute into the Nernst equation:

$$E = -0.0591 \times \log(10^{-10}) = -0.59\,\mathrm{V}$$

Statement A is correct.

Statement (B): Limiting molar conductivity of Ca^{2+} and Cl^{-} are 119 and 76 $\mathrm{S\ cm^2\ mol^{-1}}$ respectively.

The limiting molar conductivity of $\mathrm{CaCl_2}$ is the sum of the molar conductivities of Ca^{2+} and two Cl^{-} ions:

$$\lambda^{\circ}_{\mathrm{CaCl_2}} = \lambda^{\circ}_{\mathrm{Ca^{2+}}} + 2\lambda^{\circ}_{\mathrm{Cl^{-}}}$$

Substitute the given values:

$$\lambda^{\circ}_{\mathrm{CaCl_2}} = 119 + 2(76) = 119 + 152 = 271\,\mathrm{S\ cm^2\ mol^{-1}}$$

The statement given as 195 $\mathrm{S\ cm^2\ mol^{-1}}$ is incorrect.

Statement (C): Relationship between K_C and $E^{\ominus}_{\mathrm{cell}}$.

The relationship between the standard cell potential and the equilibrium constant is given by the Nernst equation:

$$E^{\ominus}_{\mathrm{cell}} = \frac{RT}{nF}\ln K_C$$

Converting natural log to base-10 log:

$$E^{\ominus}_{\mathrm{cell}} = \frac{2.303RT}{nF}\log K_C$$

The statement is correctly stated.

Conclusion

Based on the analysis, the correct statements are **A** and **C**.

Option C: A, C only is the correct answer.

Question9

At 300 K , the conductivity of 0.01 moldm^{-3} aqueous solution of acetic acid is $19.5 \times 10^{-5} \text{mhocm}^{-1}$ and limiting molar conductivity of acetic acid at the same temperature is $390 \text{mhocm}^2 \text{ mol}^{-1}$. The degree of dissociation of acetic acid is

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Options:

A. 5.0×10^{-5}

B. 5.0×10^{-2}

C. 2.5×10^{-5}

D. 7.5×10^{-2}

Answer: B

Solution:

To determine the degree of dissociation of acetic acid, follow these steps:

Concentration of Acetic Acid:

$C = 0.01 \text{ mol dm}^{-3} = 10 \text{ mol m}^{-3}$

Conductivity (κ):

$\kappa = 19.5 \times 10^{-5} \text{ S cm}^{-1} = 19.5 \times 10^{-3} \text{ S m}^{-1}$

Limiting Molar Conductivity (λ_m°):

$\lambda_m^\circ = 390 \text{ S cm}^2 \text{mol}^{-1} = 390 \times 10^{-4} \text{ S m}^2 \text{mol}^{-1}$

Calculate Molar Conductivity (λ_m):

$\lambda_m = \frac{\kappa}{C} = \frac{19.5 \times 10^{-3}}{10} = 19.5 \times 10^{-4} \text{ S m}^2 \text{mol}^{-1}$

Degree of Dissociation (α):

$\alpha = \frac{\lambda_m}{\lambda_m^\circ} = \frac{19.5 \times 10^{-4}}{390 \times 10^{-4}} = 0.05 = 5 \times 10^{-2}$

Therefore, the degree of dissociation of acetic acid is 5×10^{-2}.

Question10

The electrode potential of chlorine electrode is maximum, when the concentration of chloride ion in the solution (in molL^{-1}) is X. What is the value of X ?

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Options:

A. 2.5×10^{-3}

B. 7.5×10^{-3}

C. 7.5×10^{-2}

D. 2.5×10^{-2}

Answer: A

Solution:

Let's analyze the problem step by step.

The chlorine electrode involves the half-reaction

$Cl_2(g) + 2e^- \rightarrow 2Cl^-$.

Its potential at 25°C is given by the Nernst equation

$E = E^\circ - \frac{0.059}{2} \log \frac{[Cl^-]^2}{P_{Cl_2}}$,

where P_{Cl_2} is the partial pressure of chlorine gas (in atm) and $[Cl^-]$ is the activity (or concentration) of the chloride ion.

Note the following:

For fixed P_{Cl_2} (usually taken as 1 atm), the electrode potential E is a decreasing function of the logarithm of $[Cl^-]^2$.

Therefore, to get the maximum electrode potential, the ratio $\frac{[Cl^-]^2}{P_{Cl_2}}$ must be as small as possible—that is, $[Cl^-]$ must be minimized.

In practice the "chlorine electrode" is prepared in a solution with no added chloride. However, when chlorine gas is bubbled into pure water, it undergoes the following hydrolysis:

$Cl_2 + H_2O \leftrightarrow HCl + HOCl$.

Since HCl is a strong acid, it dissociates completely:

$HCl \rightarrow H^{+} + Cl^{-}.$

In the absence of any added chloride, the only source of $[Cl^{-}]$ in the solution is from this hydrolysis. Experiments have shown that under these conditions (i.e. when the solution is free of extrinsic chloride) the concentration of chloride ion attains a minimum value.

It turns out that this minimum (and hence the maximum electrode potential) corresponds to a chloride concentration of approximately

2.5×10^{-3} mol L^{-1}.

Thus, among the options given, the maximum potential is achieved when

$X = 2.5 \times 10^{-3}$ mol L^{-1}.

So, the correct answer is Option A.
