Electrochemistry

Question1

In a cell a copper electrode was used as a cathode. What is the electrode potential (in V) of the copper electrode dipped in $0.1 \mathrm{M} \mathrm{Cu}^{2+}$ solution at 298 K ?

$$\left(E_{\mathrm{Cu}^{2+} / \mathrm{Cu}}^{\ominus}=0.34 \mathrm{~V} ; \frac{2.303 R T}{F}=0.06 \mathrm{~V}\right)$$

AP EAPCET 2025 - 26th May Morning Shift

Options:

A.

0.34

B.

0.31

C.

0.37

D.

0.40

Answer: B

Solution:

Using the Nernst equation,

$E=E^{\circ}-\frac{2303}{n F} \log Q$

For the reduction,

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

$$a = \frac{t}{\left[\mathrm{Cu}^{2+}\right]}$$

$$E = E^{\circ} + \frac{2303}{nF}\log\left[\mathrm{Cu}^{2+}\right]$$

Substituting the values, $n = 2$

$$E^{\circ} = 0.34\ \mathrm{V},\ \left[\mathrm{Cu}^{2+}\right] = 0.1$$

$$E = 0.34 + \frac{0.056}{2}\log(0.1)$$

$$E = 0.31\ \mathrm{V}$$

Question2

Match the following

	List-I (Symbol of electrical property)		List-I (Units)
(A)	Λ_{m}	(I)	Scm^{-1}
(B)	G	(II)	m^{-1}
(C)	K	(III)	$\mathrm{Scm}^{2}\ \mathrm{mol}^{-1}$
(D)	G*	(IV)	S

The correct answer is

AP EAPCET 2025 - 26th May Evening Shift

Options:

A.

A-IV, B-III, C-I, D-II

B.

A-III, B-IV, C-I, D-II

C.

A-III, B-IV, C-II, D-I

D.

A-II, B-I, C-IV, D-III

Answer: B

Solution:

Let's match each symbol of the electrical property with its corresponding unit:

(A) Λ_m (Molar conductivity)

Molar conductivity (Λ_m) is defined as the conductivity (κ) divided by the molar concentration (c).

$\Lambda_m = \frac{\kappa}{c}$

Units of κ are S cm^{-1}.

Units of c are mol cm^{-3}.

Therefore, the units of Λ_m are $\frac{\text{S cm}^{-1}}{\text{mol cm}^{-3}} = \text{S cm}^2\ \text{mol}^{-1}$.

So, (A) matches with (III).

(B) G (Conductance)

Conductance (G) is the reciprocal of resistance (R).

$G = \frac{1}{R}$

The unit of resistance is Ohm (Ω).

Therefore, the unit of conductance is Ohm^{-1} or Siemens (S).

So, (B) matches with (IV).

(C) κ (Kappa, Conductivity or Specific conductance)

Conductivity (κ) is the reciprocal of resistivity (ρ). It is the conductance (G) of a material with unit length ($l = 1$) and unit cross-sectional area ($A = 1$).

$\kappa = G \times \frac{l}{A}$

Units of G are S.

Units of l are cm.

Units of A are cm^2.

Therefore, the units of κ are $\text{S} \times \frac{\text{cm}}{\text{cm}^2} = \text{S cm}^{-1}$.

So, (C) matches with (I).

(D) G^* (Cell constant)

The cell constant (G^*) for a conductivity cell is defined as the ratio of the distance between the electrodes (l) to the area of the electrodes (A).

$G^* = \frac{l}{A}$

Units of l are cm.

Units of A are cm^2.

Therefore, the units of G^* are $\frac{\text{cm}}{\text{cm}^{2}} = \text{cm}^{-1}$.

Alternatively, we know that conductivity $\kappa = G \times G^*$, so $G^* = \frac{\kappa}{G}$.

Units of κ are S cm^{-1}.

Units of G are S.

So, units of G^* are $\frac{\text{S cm}^{-1}}{\text{S}} = \text{cm}^{-1}$.

So, (D) matches with (II).

Summarizing the matches:

(A) - (III)

(B) - (IV)

(C) - (I)

(D) - (II)

Question3

Consider the following cell reaction

$$2\text{Fe}^{3+}(aq) + 2\text{I}^{-}(aq) \rightleftharpoons 2\text{Fe}^{2+}(aq) + \text{I}_2(s)$$

At 298 K , the cell emf is 0.237 V . The equilibrium constant for the reaction is 10^x. The value of x is $\left(F = 96500\text{Cmol}^{-1}; R = 8.3\text{JK}^{-1}\ \text{mol}^{-1}\right)$.

AP EAPCET 2025 - 24th May Morning Shift

Options:

A.

8

B.

7

C.

6

D.

9

Answer: A

Solution:

The reaction is,

$$2\mathrm{Fe}^{3+}(aq)+2\mathrm{I}^{-}(aq)\longrightarrow 2\mathrm{Fe}^{2+}(aq)+\mathrm{I}_2(s)$$
$n=2$

At equilibrium, $E_{\text{cell}}=0\ \mathrm{V}, Q=K$

Using Nernst equation

$E_{\text{cell}}=E_{\text{cell}}^{\circ}-\frac{RT}{nF}\ln K$

$\ln K=\frac{nFE_{\text{cell}}^{\circ}}{RT}$, substituting values, ln

$$K=18.493$$
$$\log_{10}K=\frac{18.493}{2.303}\approx 8\Rightarrow 10^{8}$$

So, the value of x is 8 .

Question4

When the lead storage battery is in use (during discharge) the reaction that occurs at the anode is

AP EAPCET 2025 - 23rd May Evening Shift

Options:

A.

$PbSO_4(s) + 2H_2O(l) \longrightarrow PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$

B.

$Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \longrightarrow 2PbSO_4(s) + 2H_2O(l)$

C.

$Pb(s) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s) + 2e^-$

D.

$PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \longrightarrow PbSO_4(s) + 2H_2O(l)$

Answer: C

Solution:

During discharge of a lead storage battery, the reaction at the anode involves the oxidation of lead into lead sulphate ($PbSO_4$) and releases electron.

$Pb(s) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s) + 2e^-$

Question5

Match the following

	List-I (Transition metal, M)		List-II $\left(E^{\ominus}_{M^{2+}/M}\right)$
(A)	Ni	(I)	-1.18
(B)	Mn	(II)	-0.91
(C)	Fe	(III)	-0.25
(D)	Cr	(IV)	-0.44

The correct answer is

AP EAPCET 2025 - 23rd May Evening Shift

Options:

A.

A-III, B-I, C-IV, D-I

B.

A-III, B-IV, C-I, D-II

C.

A-III, B-I, C-IV, D-II

D.

A-I, B-IV, C-II, D-III

Answer: C

Solution:

Let's match the standard electrode potentials ($E^{\ominus}_{M^{2+}/M}$) for the given transition metals. We need to recall the typical values for these elements from the electrochemical series.

The standard reduction potentials for the given metals are approximately:

- **Manganese (Mn):** $E^{\ominus}_{Mn^{2+}/Mn}$ is known to be significantly negative, typically **-1.18 V**. This is due to its strong tendency to form Mn^{2+}, especially considering the stable half-filled d-subshell (d^5) in Mn^{2+} after losing two electrons from its $[Ar]3d^54s^2$ configuration.
- **Chromium (Cr):** $E^{\ominus}_{Cr^{2+}/Cr}$ is also quite negative, typically **-0.91 V**.
- **Iron (Fe):** $E^{\ominus}_{Fe^{2+}/Fe}$ is moderately negative, typically **-0.44 V**.
- **Nickel (Ni):** $E^{\ominus}_{Ni^{2+}/Ni}$ is the least negative among the given options, typically **-0.25 V**.

Now let's match these values with the given List-II:

(A) **Ni**: Matches with (III) **-0.25 V**

(B) **Mn**: Matches with (I) **-1.18 V**

(C) **Fe**: Matches with (IV) **-0.44 V**

(D) **Cr**: Matches with (II) **-0.91 V**

So, the correct matching is:

A - III

B - I

C - IV

D - II

Let's check the given options:

- Option A: A-III, B-I, C-IV, D-I (Incorrect, D is not I)
- Option B: A-III, B-IV, C-I, D-II (Incorrect, B is not IV, C is not I)
- Option C: A-III, B-I, C-IV, D-II (This matches our derived mapping)
- Option D: A-I, B-IV, C-II, D-III (Incorrect, A is not I, B is not IV, C is not II, D is not III)

Therefore, Option C is the correct answer.

The final answer is $\boxed{\text{Option C}}$

Question6

At 298 K , the following reaction takes place for a cell at the hydrogen electrode

$$\mathrm{H}^{+}(aq)+e^{-} \longrightarrow \frac{1}{2}\mathrm{H}_2\ (1\ \mathrm{bar})$$

The solution pH is 10.0 . What is the hydrogen electrode potential in volts?

$$\left(\frac{2303RT}{F}=0.06\ \mathrm{V}\right)$$

AP EAPCET 2025 - 23rd May Morning Shift

Options:

A.

-0.6

B.

-0.06

C.

+0.6

D.

+0.06

Answer: A

Solution:

Step 1: Write the reaction and information given

The reaction is: $\mathrm{H}^+(aq) + e^- \rightarrow \frac{1}{2}\mathrm{H}_2$ (1 bar)The pH of the solution is 10.0.

Step 2: Find the concentration of H^+

We know that $\mathrm{pH} = -\log[\mathrm{H}^+]$

This means$[\mathrm{H}^+] = 10^{-\mathrm{pH}} = 10^{-10}$ M

Step 3: Note important values

Standard electrode potential for the hydrogen electrode is $E^\circ = 0$ V and the number of electrons (n) is 1.

Step 4: Write and use the Nernst Equation

The Nernst equation is:$E = E^\circ - \frac{0.0591}{n}\log[\mathrm{H}^+]$Now plug in the values:$E = 0 - \frac{0.0591}{1}\log(10^{-10})$

Step 5: Calculate the answer

$\log(10^{-10}) = -10$, so$E = 0 - (0.0591 \times -10)$

$E = 0 + 0.591$ V

Step 6: Final value and sign

The calculated value is $E = 0.591$ V. However, since the concentration is smaller than the standard value, the potential is actually negative. So, $E = -0.591$ V or approximately -0.6 V.

Question7

The resistance of a conductivity cell filled with 0.1 M KCl solution is 100Ω. If the resistance of the same cell when filled with 0.2 M KCl solution is 520Ω, the molar conductivity of 0.02 M solution (in $\mathrm{Scm}^2\ \mathrm{mol}^{-1}$) is (Given: conductivity of 0.1 M KCl solution $= 1.29\mathrm{Sm}^{-1}$)

AP EAPCET 2025 - 22nd May Evening Shift

Options:

A.

124

B.

186

C.

248

D.

104

Answer: A

Solution:

Step 1: Find the cell constant

The cell constant tells us about the cell's design and size. It can be found by using the known conductivity and resistance of 0.1 M KCl solution.

$G^{\star}=\kappa \cdot R$

So, $G^{\star}=1.29 \times 100=129 \mathrm{~m}^{-1}$ or $1.29 \mathrm{~cm}^{-1}$

Step 2: Calculate the conductivity (κ) for 0.02 M KCl solution

We use the cell constant and the new measured resistance for the 0.02 M solution.

$\kappa=\frac{G^{\star}}{R}=\frac{1.29}{520}=0.00248 \mathrm{~Sm}^{-1}$

Step 3: Find the molar conductivity (Λ_m)

Molar conductivity shows how well ions move in a solution. Use this formula:

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

$$\begin{aligned} \Lambda_m & =\frac{0.00248 \times 1000}{0.02} \\ & =124 \mathrm{~Scm}^2 \mathrm{~mol}^{-1} \end{aligned}$$

Question8

For which of the following the $E^{\ominus}\left(M^{3+} / M^{2+}\right)$ is negative?

AP EAPCET 2025 - 22nd May Evening Shift

Options:

A.

Mn

B.

Co

C.

Fe

D.

Cr

Answer: D

Solution:

Among the given options, Cr has $E^{\circ}_{\mathrm{Cr}^{3+}/\mathrm{Cr}^{2+}}$ negative.

The value of $E^{\circ}_{\mathrm{Cr}^{3+}/\mathrm{Cr}^{2+}} = -0.41\ \mathrm{V}$.

readily oxidised to Cr^{3+} with a relatively negative potential.

Question9

If $E^{\circ}_{\mathrm{Fe}^{2+}/\mathrm{Fe}} = -0.441\ \mathrm{V}$ and $E^{\circ}_{\mathrm{Fe}^{3+}/\mathrm{Fe}^{2+}} = 0.771\ \mathrm{V}$, the standard emf of the cell reaction $\mathrm{Fe}(s) + 2\mathrm{Fe}^{3+}(aq) \longrightarrow 3\mathrm{Fe}^{2+}(aq)$ is

AP EAPCET 2025 - 22nd May Morning Shift

Options:

A.

-1.212 V

B.

+1.212 V

C.

-2.424 V

D.

+2.424 V

Answer: B

Solution:

Oxidation reaction

$\mathrm{Fe} \longrightarrow \mathrm{Fe}^{2+} + 2e$

Reduction $2\mathrm{Fe}^{3+} + 2e^- \longrightarrow 2\mathrm{Fe}^{2+}$

$$\begin{aligned} E_{\text{anode}} &= E_{\mathrm{Fe}^{2+}/\mathrm{Fe}} = -0.441\ \mathrm{V}, \\ E_{\text{cathode}} &= 0.771\ \mathrm{V} \\ E_{\text{cell}} &= E_{\text{cathode}} - E_{\text{anode}} \\ &= 0.771\ \mathrm{V} - (-0.441\ \mathrm{V}) \\ &= 1.212\ \mathrm{V} \end{aligned}$$

Question10

The specific conductance of 0.05 M NaOH solution is $0.0115\ \mathrm{S\ cm}^{-1}$ What is its molar conductance ($\wedge_{\mathrm{m}}$) in $\mathrm{Scm}^2\ \mathrm{mol}^{-1}$?

AP EAPCET 2025 - 21st May Evening Shift

Options:

A.

23

B.

5.75×10^{-7}

C.

2300

D.

230

Answer: D

Solution:

Given, $\kappa = 0.0115 \mathrm{Sm}^{-1}$

Concentration of NaOH solution $= 0.05\mathrm{M}$ or molar conductance, Λ_m is given by,

$$\begin{aligned}\Lambda_m &= \frac{\kappa}{C} \times 1000 \\ &= \frac{0.0115}{0.050} \times 1000 \text{in S cm}^2 \text{ mol}^{-1} \\ \Lambda_m &= 230 \text{ S cm}^2 \text{ mol}^{-1}\end{aligned}$$

Question11

What is E_{cell} (in V) of the following cell at 298 K?

$$\left(E^{\ominus}_{\mathrm{Zn}^{2+}/\mathrm{Zn}} = -0.76 \text{ V}; E^{\ominus}_{\mathrm{Ni}^{2+}/\mathrm{Ni}} = -0.25 \text{ V}; \frac{2.303RT}{F} = 0.06 \text{ V}\right)$$

$$1(s)\mathrm{Zn}^{2+}(0.01\mathrm{M})\mathrm{Ni}^{2+}(0.1\mathrm{M})\mathrm{Ni}(s$$

AP EAPCET 2025 - 21st May Morning Shift

Options:

A.

0.51

B.

0.48

C.

0.57

D.

0.54

Answer: D

Solution:

Using Nernst equation,

$$\begin{aligned} E_{\text{cell}} &= E_{\text{cell}}^{\circ} - \frac{0.06}{n} \log \left[\frac{\mathrm{Zn}^{2+}}{\mathrm{Ni}^{2+}}\right] \\ E_{\text{cell}}^{\circ} &= E_{\text{Cathode}}^{\circ} - E_{\text{Anode}}^{\circ} \\ &= E_{\mathrm{Ni}^{2+}/\mathrm{Ni}}^{\circ} - E_{\mathrm{Zn}^{2+}/\mathrm{Zn}}^{\circ} \\ &= (-0.25 \mathrm{~V}) - (-0.76 \mathrm{~V}) = 0.51 \mathrm{~V} \end{aligned}$$

Substituting all the given values

$$\begin{aligned} E_{\text{cell}} &= 0.51 - \frac{0.06}{2} \log \left(\frac{0.01}{0.1}\right) \\ &= 0.51 - 0.03 \log (0.1) \\ &= 0.51 + 0.03 = 0.54 \mathrm{~V} \end{aligned}$$

Question12

At 300 K, the $E_{\text{cell}}^{\ominus}$ of

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

is 1.0 V. If $\Delta_r S^{\theta}$ of this reaction is $100 \mathrm{JK}^{-1}$. What is $\Delta_r H^{\ominus}$ (in kJmol^{-1}) of this reaction?

$$\left(\mathrm{F} = 96500 \mathrm{C} \mathrm{~mol}^{-1}\right)$$

AP EAPCET 2024 - 23th May Morning Shift

Options:

A. -163

B. -223

C. -193

D. -163000

Answer: A

Solution:

Given the reaction:

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

The standard cell potential is:

$E^{\ominus}=1\,\mathrm{V}$

The standard entropy change is:

$\Delta_r S^{\ominus}=100\,\mathrm{J/K}$

The temperature is:

$T=300\,\mathrm{K}$

There is a relationship between the change in Gibbs free energy, entropy, and enthalpy, given by:

$\Delta_r G^{\ominus}=\Delta_r H^{\ominus}-T\Delta_r S^{\ominus}$

Also, the change in Gibbs free energy is related to the cell potential and the number of moles of electrons transferred (n) through the equation:

$\Delta_r G^{\ominus}=-nFE^{\ominus}$

Where F is the Faraday constant ($96500\,\mathrm{C/mol}$).

Substituting into the equations, we find:

$-nFE^{\ominus}=\Delta_r H^{\ominus}-T\Delta_r S^{\ominus}$

Rearranging and substituting the known values:

$-2\times 96500\times 1=\Delta_r H^{\ominus}-300\times 100$

Solving this, we get:

$\Delta_r H^{\ominus}=-163\times 10^3\,\mathrm{J/mol}$

Therefore:

$\Delta_r H^{\ominus}=-163\,\mathrm{kJ/mol}$

Question13

Consider the cell reaction at 300 K .

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

Its $E^{\ominus}$ is 1.0 V . The $\Delta_r H^{\ominus}$ of the reaction is $-163\mathrm{kJmol}^{-1}$.

What is $\Delta_r s^{\ominus}$ (in JK^{-1}) of the reaction?

$\left(F = 96500\mathrm{C\ mol}^{-1}\right)$

AP EAPCET 2024 - 22th May Evening Shift

Options:

A. 10

B. 100

C. 1000

D. 10000

Answer: B

Solution:

To find the change in entropy ($\Delta_r S^\ominus$) for the given cell reaction at 300 K, we start with the reaction:

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

Given parameters are:

Standard cell potential, $E^\ominus = 1.0\,\mathrm{V}$

Standard reaction enthalpy change, $\Delta_r H^\ominus = -163\,\mathrm{kJ/mol}$

Temperature, $T = 300\,\mathrm{K}$

Faraday's constant, $F = 96500\,\mathrm{C/mol}$

The relationship between Gibbs free energy change ($\Delta_r G^\ominus$), enthalpy change ($\Delta_r H^\ominus$), and entropy change ($\Delta_r S^\ominus$) is given by:

$\Delta_r G^\ominus = \Delta_r H^\ominus - T\Delta_r S^\ominus$

Additionally, the change in Gibbs free energy for the cell is also related to the cell potential by the equation:

$\Delta_r G^\ominus = -nFE^\ominus$

where n is the number of moles of electrons exchanged in the reaction. Here, $n = 2$ since two electrons are transferred.

Substituting the known values into the equation:

$-nFE^\ominus = \Delta_r H^\ominus - T\Delta_r S^\ominus$

Calculate as follows:

$-2 \times 96500 \times 1 = -163000 - T\Delta_r S^\ominus$

Simplifying:

$-193000 = -163000 - 300\Delta_r S^{\ominus}$

Rearranging to solve for $\Delta_r S^{\ominus}$:

$30000 = 300\Delta_r S^{\ominus}$

$\Delta_r S^{\ominus} = \frac{30000}{300}$

$\Delta_r S^{\ominus} = 100\,\mathrm{J/K}$

Question14

In which of the following Galvanic cells emf is maximum?

(Given, $E^{\circ}_{\mathrm{Mg}^{2+}|\mathrm{Mg}} = -2.36\ \mathrm{V}$

and $E^{\circ}_{\mathrm{Cl}_2|2\mathrm{Cl}^-} = +136\ \mathrm{V}$)

AP EAPCET 2024 - 22th May Morning Shift

Options:

A.

$\mathrm{Mg}\left|\mathrm{Mg}^{2+}(1\mathrm{M})\right|\left|2\mathrm{Cl}^-(1\mathrm{M})\right|\mathrm{Cl}_2(1\ \mathrm{atm}), \mathrm{Pt}$

B.

$\mathrm{Mg}\left|\mathrm{Mg}^{2+}(0.01\mathrm{M})\right|\left|2\mathrm{Cl}^-(1\mathrm{M})\right|\mathrm{Cl}_2(1\ \mathrm{atm}), \mathrm{Pt}$

C.

$\mathrm{Mg}\left|\mathrm{Mg}^{2+}(1\mathrm{M})\right|\left|2\mathrm{Cl}'(0.01\mathrm{M})\right|\mathrm{Cl}_2(1\ \mathrm{atm}), \mathrm{Pt}$

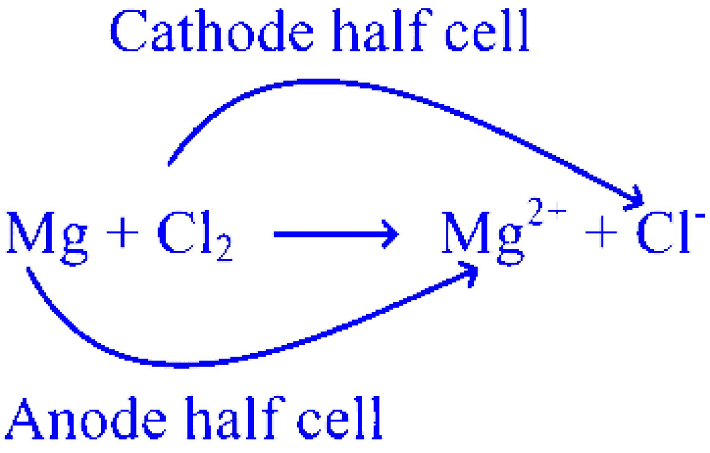
D.

$\mathrm{Mg}^2\left|\mathrm{Mg}^{2+}(0.01\mathrm{M})\right\|2\mathrm{Cl}^-(0.01\mathrm{M})\left|\mathrm{Cl}_2(1\ \mathrm{atm}), \mathrm{Pt}\right.$

Answer: D

Solution:

The cell reaction is,



$$\therefore \quad Q=\frac{\left[\mathrm{Mg}^{2+}\right]\left[\mathrm{Cl}^{-}\right]^{2}}{p_{\mathrm{Cl}_2}}$$

$$\text{and } E_{\text{cell}} = E^{\circ}-\frac{0.0591}{2} \log Q$$

$$\text{So, } E_{\text{cell}} \propto \frac{1}{Q}$$

(a) $Q=\frac{1 \times 1^{2}}{1}=1$;

(b) $Q=\frac{0.01 \times 1^{2}}{1}=10^{-2}$

(c) $Q=\frac{1 \times(0.01)^{2}}{1}=10^{-4}$;

(d) $Q=\frac{0.01 \times(0.01)^{2}}{1}=10^{-6}$

$\Rightarrow$ The order of the values of

$$ Q: d

$\Rightarrow$ The order of emf of galvanic cells:

$d>c>b>a$

The emf is maximum in

$\mathrm{Mg}\left|\mathrm{Mg}^{2+}(0.01 \mathrm{M})\right|\left|2 \mathrm{Cl}^{-}(0.01 \mathrm{M})\right|$
$\mathrm{Cl}_2(1 \text{ atm}) \mid \mathrm{Pt}$

Question15

Consider the following standard electrode potentials (E° in volts) in aqueous solution.

Element	M^{3+}/M	M^{+}/M
Al	-1.66	$+0.55$
TI	$+1.26$	-0.34

Based on this data. which of the following statements is correct?

AP EAPCET 2024 - 21th May Evening Shift

Options:

A. Tl^{3+} is more stable than Al^{3+}

B. $\mathrm{Al}^{+}$is more stable than Al^{3+}

C. $\mathrm{Al}^{+}$is more stable than Al^{3+}

D. $\mathrm{TI}^{+}$is more stable than Al^{+}

Answer: D

Solution:

Let's analyze the data step by step.

We are given the following standard electrode reactions and their potentials:

For aluminum (Al):

$\mathrm{Al}^{3+} + 3e^{-} \rightarrow \mathrm{Al}, \quad E^{\circ} = -1.66\ \mathrm{V}$

$\mathrm{Al}^{+} + e^{-} \rightarrow \mathrm{Al}, \quad E^{\circ} = +0.55\ \mathrm{V}$

For thallium (Tl):

$\mathrm{Tl}^{3+} + 3e^{-} \rightarrow \mathrm{Tl}, \quad E^{\circ} = +1.26\ \mathrm{V}$

$\mathrm{Tl}^{+} + e^{-} \rightarrow \mathrm{Tl}, \quad E^{\circ} = -0.34\ \mathrm{V}$

A key idea in interpreting these potentials is that a more positive reduction potential means the ion is a stronger oxidizing agent and is more eager to gain electrons (i.e., it is less “stable” in its oxidized form). In contrast, a more negative reduction potential indicates that the species is less likely to gain electrons and is more stable in that oxidation state.

Let’s apply this reasoning:

- For aluminum:

The reaction $Al^{3+} \rightarrow$ Al has a very negative potential (–1.66 V), meaning that once aluminum is in the +3 state, it does not easily pick up electrons to revert to the metal. This makes Al^{3+} a very stable oxidation state.

The reaction involving Al^{+} has a positive potential (+0.55 V), indicating that Al^{+} is readily reduced to aluminum, and thus is less stable in the +1 state.

- For thallium:

The reaction $Tl^{3+} \rightarrow$ Tl has a highly positive potential (+1.26 V), so Tl^{3+} is not very stable; it tends strongly to gain electrons.

In contrast, the reaction involving Tl^{+} has a negative potential (–0.34 V), which indicates that Tl^{+} does not easily accept an electron, making it a relatively stable oxidation state.

Now, let's look at the options:

Option A: "Tl^{3+} is more stable than Al^{3+}."

- Because Tl^{3+} (with +1.26 V) is eager to gain electrons (a strong oxidizing agent), it is less stable compared to Al^{3+} (with –1.66 V). Thus, Option A is incorrect.

Option B and Option C: "Al^{+} is more stable than Al^{3+}."

- Our data indicates that Al^{+} (with +0.55 V) is less stable (more readily reduced) than Al^{3+} (with –1.66 V). So, these statements are not correct.

Option D: "Tl^{+} is more stable than Al^{+}."

- Comparing the two, the reduction reaction for Tl^{+} has $E^\circ = -0.34$ V while that for Al^{+} has $E^\circ = +0.55$ V. The more negative potential for Tl^{+} shows it is less likely to be reduced to elemental Tl, meaning Tl^{+} is more stable than Al^{+}.

Thus, Option D is the correct statement.

In summary, the correct answer is:

Option D: Tl^{+} is more stable than Al^{+}.

Question16

The standard reduction potentials of $2H^{+}/H_2$, Cu^{2+}/Cu, Zn^{2+}/Zn and $NO_3^{-}, H^{-}/NO$ are 0.0 0.34 . -0.76 and 0.97 V respectively. Identify the correct statements from the following.

I. H^{+}does not oxidise Cu to Cu^{2+}

II. Zn reduces Cu^{2+} to Cu

III. NO_3^{-}oxidises Cu to Cu^{2+}

Options:

A. I, II only

B. I,II, III

C. I, III only

D. II, III only

Answer: B

Solution:

Let's analyze each statement using the given standard reduction potentials:

$2H^+ + 2e^- \rightarrow H_2, \quad E^\circ = 0.0 \text{ V}$

$Cu^{2+} + 2e^- \rightarrow Cu, \quad E^\circ = +0.34 \text{ V}$

$Zn^{2+} + 2e^- \rightarrow Zn, \quad E^\circ = -0.76 \text{ V}$

For nitrate reducing to nitric oxide (in acidic solution):

$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O, \quad E^\circ = +0.97 \text{ V}$

Now, let’s consider each statement one by one.

Statement I: “H^+ does not oxidise Cu to Cu^{2+}”

The possible reaction would be:

$Cu(s) + 2H^+ \rightarrow Cu^{2+} + H_2(g)$

To check spontaneity, write the half-reactions:

Oxidation (copper):

$Cu(s) \rightarrow Cu^{2+} + 2e^-,$

(reverse of the given reduction reaction with $E^\circ = -0.34$ V)

Reduction (hydrogen):

$2H^+ + 2e^- \rightarrow H_2, \quad E^\circ = 0.0 \text{ V}$

The cell potential is:

$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = 0.0 \text{ V} - 0.34 \text{ V} = -0.34 \text{ V}$

A negative cell potential means the reaction is non-spontaneous.

Thus, Statement I is correct.

Statement II: “Zn reduces Cu^{2+} to Cu”

The reaction is:

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

Write the half-reactions:

Oxidation (zinc):

$$Zn(s) \rightarrow Zn^{2+} + 2e^-,$$

The oxidation potential for Zn is the reverse of its reduction: $+0.76$ V.

Reduction (copper):

$$Cu^{2+} + 2e^- \rightarrow Cu(s), \quad E^\circ = +0.34\ \text{V}$$

Combine the potentials:

$$E^\circ_{\text{cell}} = 0.34\ \text{V} + 0.76\ \text{V} = +1.10\ \text{V}$$

A positive cell potential indicates the reaction is spontaneous.

Thus, Statement II is correct.

Statement III: “NO_3^- oxidises Cu to Cu^{2+}”

In reactions with nitric acid (which provides NO_3^- in acidic medium), copper is oxidized. The half-reactions are:

Oxidation (copper):

$$Cu(s) \rightarrow Cu^{2+} + 2e^-, \quad E^\circ_{\text{ox}} = -0.34\ \text{V}$$

Reduction (nitrate in acid):

$$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O, \quad E^\circ = +0.97\ \text{V}$$

Even though the electrons do not match directly (2 vs. 3 electrons), the overall potential difference can be determined by the difference in the standard electrode potentials:

$$E^\circ_{\text{net}} = 0.97\ \text{V} - 0.34\ \text{V} = +0.63\ \text{V}$$

A positive net potential shows that the oxidation of Cu by NO_3^- in acidic conditions is thermodynamically favored.

Thus, Statement III is correct.

Since all three statements (I, II, and III) are correct, the correct option is:

Option B: I, II, III

Question17

As per standard reduction potential values, which is the strongest reducing agent among the given elements?

AP EAPCET 2024 - 21th May Morning Shift

Options:

A. Rb

B. Sr

C. Na

D. Mg

Answer: A

Solution:

A reducing agent is a substance that donates electrons during a chemical reaction. The more negative the standard reduction potential (given by the half-reaction), the stronger the tendency of the substance to lose electrons—making it a stronger reducing agent.

Let's compare the given elements by their half-reactions and approximate standard reduction potentials:

For sodium (Na):

$\mathrm{Na}^{+}+\mathrm{e}^{-} \rightarrow \mathrm{Na} \quad E^{\circ} \approx-2.71 V$

For rubidium (Rb):

$\mathrm{Rb}^{+}+\mathrm{e}^{-} \rightarrow \mathrm{Rb} \quad E^{\circ} \approx-2.98 V$

For magnesium (Mg):

$\mathrm{Mg}^{2+}+2 \mathrm{e}^{-} \rightarrow \mathrm{Mg} \quad E^{\circ} \approx-2.36 V$

For strontium (Sr):

$\mathrm{Sr}^{2+}+2 \mathrm{e}^{-} \rightarrow \mathrm{Sr} \quad E^{\circ} \approx-2.89 V$

Since the strength of a reducing agent corresponds to having a very negative standard reduction potential, we compare the values:

Rb: approximately -2.98 V

Sr: approximately -2.89 V

Na: approximately -2.71 V

Mg: approximately -2.36 V

Rubidium (Rb) has the most negative reduction potential. Therefore, it is the strongest reducing agent among the given elements.

The correct answer is: Option A (Rb).

Question18

The standard reduction potentials of $2H^+/H_2$, Cu^{2+}/Cu, Zn^{2+}/Zn and $NO_3^-, H^-/NO$ are 0.0 , +0.34 . -0.76 and 0.97 V respectively. Observe the following reactions

I. $Zn + HCI \rightarrow$

II. $Cu + HCl \rightarrow$

III. $Cu + HNO_3 \rightarrow$

Which reactions does not liberate $H_2(g)$?

AP EAPCET 2024 - 20th May Evening Shift

Options:

A. II, III only

B. I, II only

C. I, III only

D. I, II, III

Answer: A

Solution:

The standard reduction potentials for the following half-reactions are:

$2H^+/H_2$: 0.0 V

Cu^{2+}/Cu: +0.34 V

Zn^{2+}/Zn: -0.76 V

$NO_3^-, H^-/NO$: 0.97 V

Let's consider the reactions:

$Zn + HCl \rightarrow$

$Cu + HCl \rightarrow$

$Cu + HNO_3 \rightarrow$

We are tasked with identifying which of these reactions do not liberate $H_2(g)$.

Reactions II and III do not produce H_2 because they are not spontaneous. This conclusion is based on their cell potentials (E°_{cell}) being negative ($E^{\circ}_{\text{cell}} < 0$).

Calculation of Cell Potentials:

Reaction II: $Cu + HCl$

$$\begin{aligned} E^{\circ}_{\text{cell}} &= E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} \\ &= 0 - 0.34 \\ &= -0.34 \text{ V} \end{aligned}$$

Reaction III: $Cu + HNO_3$

$$\begin{aligned} E^{\circ}_{\text{cell}} &= E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} \\ &= 0 - 0.97 \\ &= -0.97 \text{ V} \end{aligned}$$

Since both reactions have negative cell potentials, they are non-spontaneous and thus do not release $H_2(g)$.

Question19

Aqueous $CuSO_4$ solution was electrolysed by passing 2 amp of current for 10 min . What is the weight (in g) of copper deposited at cathode ?

$$\left(\text{Cu} = 63\text{u}; F = 96500\text{C mol}^{-1}\right)$$

AP EAPCET 2024 - 20th May Morning Shift

Options:

A. 0.195

B. 0.39

C. 0.78

D. 1.56

Answer: B

Solution:

To determine the mass of copper deposited at the cathode during electrolysis, we first need to gather the given data and use the formula for electrolysis. The relevant parameters are:

Current, $I = 2\,\mathrm{A}$

Time, $t = 10\,\text{minutes}$

Convert the time from minutes to seconds for consistency with the formula:

$t = 10 \times 60 = 600\,\text{seconds}$

Next, calculate the total charge Q passed during electrolysis using the formula:

$Q = I \times t = 2 \times 600 = 1200\,\mathrm{C}$

The electrochemical reaction for copper deposition is:

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

Using the formula for the mass of substance deposited during electrolysis:

Mass of copper deposited $= \frac{\text{Molar mass} \quad \times \text{Charge}}{\text{Number of electrons transferred} \quad \times F}$

Given:

Molar mass of Cu = 63 g/mol

Faraday constant, $F = 96500\,\mathrm{C/mol}$

Number of electrons transferred for copper, = 2

Substitute these values into the formula:

Mass of copper deposited $= \frac{63 \times 1200}{2 \times 96500} = 0.39\,\mathrm{g}$

Thus, the weight of copper deposited at the cathode is 0.39 grams.

Question20

2.644 g of metal (M) was deposited when 8040 coulombs of electricity was passed through molten $M\,\mathrm{F}_2$ salt. What is the atomic mass of M? $\left(\mathrm{F} = 96500\mathrm{C\ mol}^{-1}\right)$

AP EAPCET 2024 - 19th May Evening Shift

Options:

A. 63.47 u

B. 65.54 u

C. 31.74 u

D. 61.48 u

Answer: A

Solution:

To find the atomic mass of the metal M, given that 2.644 g of metal was deposited with a total charge of 8040 Coulombs, we follow these steps:

Calculate the moles of electrons used:

$$\text{Moles of electrons} = \frac{\text{Total charge}}{\text{Faraday's constant}} = \frac{8040\text{ C}}{96500\text{ C/mol}} = 0.0833\,\text{mol}$$

Determine the moles of metal deposited:

The reaction $M\text{F}_2 \rightarrow M^{2+} + 2\text{F}^-$ implies that 2 moles of electrons are needed to deposit 1 mole of metal.

$$\text{Moles of metal}(M) = \frac{\text{Moles of electrons}}{2} = \frac{0.0833}{2} = 0.04165\,\text{mol}$$

Calculate the atomic mass of metal M:

$$\text{Atomic mass of } M = \frac{\text{Mass of metal}}{\text{Number of moles of metal}} = \frac{2.644\text{ g}}{0.04165\text{ mol}} \approx 63.47$$

Thus, the atomic mass of metal M is approximately 63.47 u.

Question21

The anode and cathode used in electrolytic refining of copper respectively are

AP EAPCET 2024 - 19th May Evening Shift

Options:

A. pure copper, impure copper

B. impure copper, pure copper

C. pure copper, pure zinc.

D. impure copper, pure zinc

Answer: B

Solution:

In the electrolytic refining process of copper:

The anode is made up of impure copper. During the process, oxidation occurs at the anode, meaning the impure copper dissolves into the electrolyte.

The cathode is made of pure copper. Here, reduction takes place; copper ions gain electrons and deposit as pure copper on the cathode.

Mathematically, the reactions are as follows:

At the anode (oxidation):

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

At the cathode (reduction):

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

Therefore, the correct option is:

Option B: impure copper, pure copper.

Question22

The E^{-} of $M\left|M^{2+}\right|\left|\mathrm{Cu}^{2+}\right| \mathrm{Cu}$ is 0.3 V .

At what concentration of $\mathrm{Cu}^{2+}\left(\mathrm{inmolLL}^{-1}\right)$, the $\mathrm{E}_{\mathrm{cel}}$ value becomes zero ? $\left(\frac{2.303 R T}{F}=0.06\right)$

(Conc. of $\mathrm{M}^{2+}=0.1 \mathrm{M}$)

AP EAPCET 2024 - 18th May Morning Shift

Options:

A. 10^{-3}

B. 10^{-11}

C. 10^{-8}

D. 10^{-10}

Answer: B

Solution:

The given standard electrode potential is:

$E^{\circ}_{\{cell\}} = 0.3\,\mathrm{V}$

According to the Nernst equation:

$$E_{\{cell\}} = E^{\circ}_{\{cell\}} - \frac{2.303RT}{nF}\log\left[\frac{[\mathrm{Cu}^{2+}]}{[\mathrm{M}^{2+}]}\right]$$

In this context:

$n = 2$ (since it involves the transfer of 2 electrons)

$E_{\{cell\}} = 0$ (when the cell voltage is zero)

$[\mathrm{M}^{2+}] = 0.1\,\mathrm{M}$

$\frac{2.303RT}{F} = 0.06$

Substituting these values into equation (1) gives:

$$0.3 = -\frac{0.06}{2}\log\left[\frac{[\mathrm{Cu}^{2+}]}{0.1}\right]$$

Solving for the log term:

$$-\frac{0.6}{0.06} = \log\left[\frac{[\mathrm{Cu}^{2+}]}{0.1}\right]$$

Taking the antilog:

$$10^{-10} = \frac{[\mathrm{Cu}^{2+}]}{0.1}$$

This implies:

$$[\mathrm{Cu}^{2+}] = 10^{-10} \times 10^{-1} = 10^{-11}\,\mathrm{mol/L}$$

Therefore, the concentration of Cu^{2+} required to make the cell potential zero is $\boxed{10^{-11}\,\mathrm{mol/L}}$.

Question23

96.5 amperes current is passed through the molten $\mathrm{AlCl_3}$ for 100 seconds. The mass of aluminium deposited at the cathode is (atomic weight of $\mathrm{Al} = 27\mathrm{u}$)

Options:

A. 0.90 g

B. 0.45 g

C. 1.35 g

D. 1.8 g

Answer: A

Solution:

Given :

$$I = 96.5\,\text{A}$$
$$t = 100\,\text{s}$$
$$m = ?$$

The atomic weight of aluminum (Al) is 27 u.

First, calculate the total charge (Q):

$$\begin{aligned} Q &= I \times t \\ &= 96.5 \times 100 \\ &= 9650\,\text{C} \end{aligned}$$

Consider the dissociation reaction of AlCl_3:

$$\begin{aligned} \text{AlCl}_3 &\rightarrow \text{Al}^{3+} + 3\text{Cl}^- \\ \text{Al}^{3+} + 3e^- &\rightarrow \text{Al} \end{aligned}$$

Understanding that 96,500 C (or 3 Faradays) of electricity deposits 1 mole of aluminum:

$$\frac{9650\,\text{C}}{96,500\,\text{C/mol}} = 0.0333\,\text{mol}$$

The weight of aluminum deposited can be calculated as:

$$\begin{aligned} \text{Weight of Al} &= n \times \text{atomic mass} \\ &= 0.0333 \times 27 \\ &\approx 0.90\,\text{g} \end{aligned}$$

Question24

38.6 amperes of current is passed for 100 seconds through an aqueous CuSO_4 solution using platinum electrodes. The mass of copper

consumed from the solution and volume of gas liberated at STP are respectively (molar mass of $\mathrm{Cu} = 63.54 \mathrm{~g} \mathrm{~mol}^{-1}$).

AP EAPCET 2022 - 4th July Evening Shift

Options:

A. 6.37 g, 0.448 L

B. 0.63 g, 0.224 L

C. 1.27 g, 0.224 L

D. 4 g, 0.448 L

Answer: C

Solution:

Given, current, $I = 38.6$ A, Time, $t = 100$ s

Cathode $\mathrm{Cu}^{2+}(aq) + 2e^{\theta} \longrightarrow \mathrm{Cu}(s)$

Anode $2\overline{\mathrm{O}}\mathrm{H}(\mathrm{aq}) \longrightarrow \mathrm{H}_2\mathrm{O}(l) + \frac{1}{2}\mathrm{O}_2(g) + 2e^{\ominus}$

Using Faraday's 1st law,

$W = ZIt, \quad (Z \rightarrow$ electrochemical equivalent $)$

$$= \frac{EIt}{96500} = \frac{MIt}{2 \times 96500}$$
$$= \frac{63.5 \times 38.6 \times 100}{2 \times 96500} = 1.27 \mathrm{~g}$$

1.27 g Cu is deposited.

Using Faraday's second law,

$$\frac{W_{\mathrm{Cu}}}{E_{\mathrm{Cu}}} = \frac{W_{\mathrm{O}_2}}{E_{\mathrm{O}_2}} \Rightarrow \frac{1.27}{63.5} = \frac{W_{\mathrm{O}_2}}{16}$$
$$W_{\mathrm{O}_2} = \frac{1.27 \times 16}{63.5} = 0.32 \mathrm{~g}$$

Moles of $\mathrm{O}_2 = \frac{0.32}{32} = 0.01$ mol

Volume of O_2 at STP $= 0.01 \times 22.4 = 0.224$ L

Question25

The reduction potential of hydrogen electrode at $25^\circ\mathrm{C}$ in a neutral solution is ($p_{\mathrm{H}_2} = 1$ bar)

AP EAPCET 2022 - 4th July Morning Shift

Options:

A. 0.059 V

B. -0.059 V

C. -0.413 V

D. 0.0 V

Answer: C

Solution:

$$\mathrm{H}^+(\mathrm{aq}) + e^- \longrightarrow \tfrac{1}{2}\mathrm{H}_2(g)$$
$$p_{\mathrm{H}_2} = 1\mathrm{bar}$$

In a neutral electrode, concentration of H^+ is $10^{-7}\mathrm{M}$.

Temperature, $T = 298$ K, $E^\circ_{\mathrm{H}^+/\mathrm{H}_2/\mathrm{Pt}} = 0$ V

Using equation,

$$E = E^\circ_{\mathrm{H}^+/\mathrm{H}_2/\mathrm{Pt}} - \frac{0.0591}{1}\log\frac{[p_{\mathrm{H}_2}]^{1/2}}{[\mathrm{H}^+]}$$
$$= 0 - 0.0591\log\frac{1}{10^{-7}}$$
$$= -0.0591\log 10^7 = -0.0591 \times 7\log 10$$
$$= -0.0591 \times 7 = -0.4137\ \mathrm{V}$$

Question26

In the electrolysis of a CuSO_4 solution, howmany grames of Cu are plated out on thecathode, in the time that is required toliberate 5.6 L of O_2(g), measured at 1 atm and273 K, at the anode?

AP EAPCET 2021 - 20th August Evening Shift

Options:

A. 31.75 g

B. 14.2 g

C. 4.32 g

D. 3.175 g

Answer: A

Solution:

Volume of $\mathrm{O}_2 = 5.6 \mathrm{~L}$

Number of moles of $\mathrm{O}_2 = \frac{5.6}{22.4} = 0.25 \mathrm{~mol}$

Equivalent of $\mathrm{O}_2 = n$ factor $\times$ number of moles of $\mathrm{O}_2 = 4 \times 0.25 = 1$

Equivalent of $\mathrm{Cu} =$ equivalent of $\mathrm{O}_2 = 1$

$$\frac{\text{Mass of } \quad \mathrm{Cu}}{\text{Equivalent mass}} = 1 \Rightarrow \frac{\text{Mass of } \quad \mathrm{Cu}}{63.5/2} = 1$$

Mass of $\mathrm{Cu} = \frac{63.5}{2} = 31.75 \mathrm{~g}$

Question27

If hydrogen electrons dipped in two solutionsof pH = 3 and pH = 6 are connected by a saltbridge, the emf of the resulting cell is

AP EAPCET 2021 - 20th August Morning Shift

Options:

A. 0.177 V

B. 0.3 V

C. 0.052 V

D. 0.104 V

Answer: A

Solution:

$$
\begin{aligned}
\mathrm{pH}_1 &= 3 \\
\Rightarrow \left[\mathrm{H}^{+}\right]_1 &= 10^{-3}\mathrm{M} \\
\mathrm{pH}_2 &= 6 \\
\Rightarrow \left[\mathrm{H}^{+}\right]_2 &= 10^{-6}\mathrm{M}
\end{aligned}
$$

$$
\begin{aligned}
\text{Now, } E_{\text{cell}} &= E\Upsilon_{\text{cell}} + \frac{0.059}{1}\log \frac{10^{-3}}{10^{-6}} = 0 + 0.059 \times 3 \\
E_{\text{cell}} &= 0.177\ \mathrm{V}
\end{aligned}
$$

Question28

At $291\ \mathrm{K}$, saturated solution of BaSO_4 was found to have a specific conductivity of $3.648 \times 10^{-6}\mathrm{ohm}^{-1}\ \mathrm{cm}^{-1}$ and that of water being used is $1.25 \times 10^{-6}\mathrm{ohm}^{-1}\ \mathrm{cm}^{-1}$. If the ionic conductances of Ba^{2+} and SO_4^{2-} are 110 and $136.6\mathrm{ohm}^{-1}\ \mathrm{cm}^2\ \mathrm{mol}^{-1}$ respectively. The solubility of BaSO_4 at $291\ \mathrm{K}$ will be [Atomic masses of $\mathrm{Ba} = 137,\ \mathrm{S} = 32, \mathrm{O} = 16$]

AP EAPCET 2021 - 19th August Evening Shift

Options:

A. $1.435 \times 10^{-3}\mathrm{gL}^{-1}$

B. $2.266 \times 10^{-3}\mathrm{gL}^{-1}$

C. $2.843 \times 10^{-3}\mathrm{gL}^{-1}$

D. $1.768 \times 10^{-3}\mathrm{gL}^{-1}$

Answer: B

Solution:

Specific conductivity of BaSO_4 solution

$$\kappa_{\mathrm{BaSO}_4} = 3.648 \times 10^{-6}\mathrm{ohm}^{-1}\ \mathrm{cm}^{-1}$$

Specific conductivity of H_2O, $\kappa_{H_2O} = 1.25 \times 10^{-6}\ \mathrm{ohm}^{-1}\ \mathrm{cm}^{-1}$

Now, $$\begin{aligned} \kappa_{\mathrm{BaSO}_4} &= \kappa_{\mathrm{BaSO}_4}\ (\text{solution}) - \kappa_{H_2O} \\ &= 3.648 \times 10^{-6} - 1.25 \times 10^{-6} \\ &= 2.398 \times 10^{-6}\mathrm{Scm}^{-1} \end{aligned}$$

Formula used, $\lambda_{\mathrm{m}}^{\circ} = \frac{1000 \times \kappa}{S}$ ..... (i)

Here, $S =$ solubility

$\kappa =$ kappa (specific conductivity of BaSO_4)
$\lambda_{\mathrm{m}}\left(\mathrm{BaSO}_4\right) = \lambda_{\mathrm{m}}^{\circ}\left(\mathrm{Ba}^{2+}\right) + \lambda_{\mathrm{m}}^{\circ}\left(\mathrm{SO}_4^{2-}\right)$
$(110 + 136.6)\mathrm{ohm}^{-1}\ \mathrm{cm}^{-1} = 246.6\mathrm{ohm}^{-1}\ \mathrm{cm}^{-1}$

Put values in Eq (i) we get,

$$246.6\mathrm{ohm}^{-1}\ \mathrm{cm}^{-1} = \frac{1000 \times 2.398 \times 10^{-6}\ \mathrm{S\ cm}^{-1}}{S}$$
$S = 9.72 \times 10^{-6} \times 233 = 2.266 \times 10^{-3}\mathrm{gL}^{-1}$

Hence, the solubility of BaSO_4 at 291 K will be $2.266 \times 10^{-3}\ \mathrm{g\ L}^{-1}$

Question29

Find the emf of the following cell reaction. Given, $E^{\Upsilon}_{\mathrm{Cr}^{3+}/\mathrm{Cr}^{2+}} = -0.72\ \mathrm{V}$ and $E^{\Upsilon}_{\mathrm{Fe}^{2+}/\mathrm{Fe}} = -0.42\ \mathrm{V}$ at $25^{\circ}\mathrm{C}$ is

$$\mathrm{Cr}\left|\mathrm{Cr}^{3+}(0.1\mathrm{M})\right|\ \left|\ \mathrm{Fe}^{2+}(0.1\mathrm{M})\ \right|\ \mathrm{Fe}$$

AP EAPCET 2021 - 19th August Evening Shift

Options:

A. 0.30 V

B. 0.25 V

C. 1.14 V

D. 1.56 V

Answer: A

Solution:

$$\underset{\substack{\text{Anode}\\ \text{(oxidation)}}}{Cr/Cr^{3+}\ (0.1\ M)}\ ||\ \underset{\substack{\text{(Cathode}\\ \text{(reduction)}}}{Fe^{2+}\ (0.1\ M)/Fe}$$

Anode $Cr \longrightarrow Cr^{3+} + 3e^{-}$ ..... (i)

Cathode $Fe^{2+} \longrightarrow 2e^{-} + Fe$ ..... (ii)

Multiply 2 in Eq. (i) and 3 in Eq. (ii) we get,

$$\begin{array}{l} 2Cr \longrightarrow 2Cr^{3+} + 6e^{-} \\ 3Fe^{2+} \longrightarrow 3Fe + 6e^{-} \\ \hline 2Cr + 3Fe^{2+} \rightarrow 2Cr^{3+} + 3Fe \end{array}$$

For, $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$

$= -0.42 - (-0.72\text{ V}) = -0.42 + 0.72\text{ V}$

$E^{\circ}_{\text{cell}} = 0.30\text{ V}$

Using Nernst equation,

$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0591}{n} \log \frac{[Cr^{3+}]^2}{[Fe^{2+}]^3}$

Here, $n = 6$ (For $6e^{-}$)

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

$E_{\text{cell}} = 0.30 - 0 \Rightarrow E_{\text{cell}} = 0.30\text{ V}$

Hence, emf of the reaction cell is 0.30 V.

Question30

For $Cr_2O_7^{2-} + 14H^{+} + 6e^{-} \xrightarrow{\text{Yields}} 2Cr^{3+} + 7H_2O, E^{\Upsilon} = 1.33$ V at $[Cr_2O_7^{2-}] = 4.5$ millimole, $[Cr^{3+}] = 1.5$ millimole and $E = 1.067$ V, then calculate the pH of the solution.

AP EAPCET 2021 - 19th August Evening Shift

Options:

A. 2

B. 3

C. 2.5

D. 1.5

Answer: A

Solution:

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \overset{\text{yields}}{\longrightarrow} 2Cr^{3+} + 7H_2O$$

Using Nernst equation,

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{n} \log \frac{[Cr^{3+}]^2 [H_2O]^7}{\left[(Cr_2O_7)^{2-}\right][H^+]^{14}}$$

$$1.067 = 1.33 - \frac{0.0591}{6} \log \frac{\left(1.5 \times 10^{-3}\right)^2 (1)^7}{\left(4.5 \times 10^{-3}\right)[H^+]^{14}}$$

$$0.26 = \frac{0.0591}{6} \log \frac{\left[225 \times 10^{-6}\right]}{\left(4.5 \times 10^{-3}\right) \times [H^+]^{14}}$$

$$26.7 = \log \frac{50 \times 10^{-3}}{[H^+]^{14}}$$

$$26.7 = \log\left(50 \times 10^{-3}\right) - \log [H^+]^{14}$$

$$\text{For pH} = -\log [H^+]$$

$$26.7 = -1.3010 - 14 \log [H^+]$$

$$14\text{pH} = 28.041 \Rightarrow \text{pH} = \frac{28.041}{14}$$

$$\text{pH} = 2$$

Question31

Assertion (A) Sodium acetate on Kolbe'selectrolysis gives ethane.

Reason (R) Methyl free radical is formed atcathode.

AP EAPCET 2021 - 19th August Morning Shift

Options:

A. Both A and R are true and R is a correct explanation of A.

B. Both A and R are true but R is not a correct explanation of A.

C. A is true but R is false.

D. A is false but R is true.

Answer: C

Solution:

Kolbe's electrolysis

$2\mathrm{RCOOK} + 2\mathrm{H_2O} \longrightarrow R - R + 2\mathrm{CO_2} + 2\mathrm{KOH} + \mathrm{H_2}$

Reaction at anode :

$$\begin{aligned} 2\mathrm{CH_3CO\dot{O}} &\longrightarrow 2\mathrm{\dot{C}H_3} + 2\mathrm{CO_2} \\ 2\mathrm{\dot{H}_3} &\longrightarrow \mathrm{C_2H_6} \end{aligned}$$

As after completion of the reaction solution becomes alkaline, so pH of solution will increase. Sodium acetate on Kolbe's electrolysis gives ethane. It is formed at anode. $\mathrm{\dot{C}H_3}$ methyl radical is produced at anode only.

Hence, A is true but R is false.

Question32

When a current of 10 A is passes throughmolten $\mathrm{AlCl_3}$ for 1.608 minutes. The mass ofAl deposited will be

[Atomic mass of Al = 27 g]

AP EAPCET 2021 - 19th August Morning Shift

Options:

A. 0.09 g

B. 0.81 g

C. 1.35 g

D. 0.27 g

Answer: A

Solution:

Given current, $i=10 \mathrm{~A}$

Time, $t=1.608$ min or 96.5 s

Weight of aluminium $W_{\mathrm{Al}}=?$

Atomic weight of aluminium $=27 \mathrm{~g} / \mathrm{mol}(M)$

We know that, $\mathrm{AlCl}_3 \rightarrow \mathrm{Al}^{3+}+3 \mathrm{Cl}^{-}, n=3$

At cathode, reduction $\mathrm{Al}^{3+}+3 e^{-} \rightarrow \mathrm{Al}$

Formula, $\quad W_{\mathrm{Al}}=\frac{\varepsilon \times i \times t}{96500}$, where $\varepsilon=\frac{M}{n}$

$$
\begin{aligned}
W_{\mathrm{Al}} & =\frac{27}{3} \times \frac{10 \times 96.5}{96500} \\
W_{\mathrm{Al}} & =0.09 \mathrm{~g}
\end{aligned}
$$

Mass of Al deposited will be 0.09 g

Question33

The molar conductivities $\left(\lambda_{\mathrm{m}}^{\Upsilon}\right)$ at infinite dilution of $\mathrm{KBr}, \mathrm{HBr}$ and KNH_2 are 120.5, 420.6 and $90.48 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$ respectively. Find the value of $\lambda_{\mathrm{m}}^{\Upsilon}$ for NH_3.

AP EAPCET 2021 - 19th August Morning Shift

Options:

A. $511.0 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

B. $390.5 \mathrm{~s} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

C. $256.2 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

D. $240.9 \mathrm{~s} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

Answer: B

Solution:

At infinite dilution, the molar conductivity of an electrolyte can be represented as the sum individual contribution of its cations and anions.

$\lambda^{\circ}_{\mathrm{m}}\left(\mathrm{NH}_3\right) \longrightarrow \lambda^{\circ}_{\infty}\left(\mathrm{H}^{+}\right)+\lambda^{\circ}_{\infty}\left(\mathrm{NH}_2^{-}\right)$
$\lambda^{\circ}_{\mathrm{m}}\left(\mathrm{NH}_3\right)=\lambda^{\circ}_{\infty}\left(\mathrm{NH}_2^{-}\right)+\lambda^{\circ}_{\infty}\left(\mathrm{H}^{+}\right)+\lambda^{\circ}$
$\mathrm{KBr}-\lambda^{\circ}\mathrm{KBr}$
$=\lambda^{\circ}_{\infty}\left(\mathrm{KNH}_2\right)+\lambda^{\circ}_{\infty}(\mathrm{HBr})-\lambda^{\circ}_{\infty}(\mathrm{KBr})$
$=90.48+420.6-120.5$
$=390.5\ \mathrm{S\ cm\ mol}^{-1}.$
