Electrochemistry

Question1

What is percent dissociation of NH_4OH if molar conductance at zero concentration for NH_4Cl, NaCl and NaOH are 130, 109 and $213 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$ respectively and molar conductivity of $0.01 \quad \mathrm{M} \quad NH_4OH$ is $9.0 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$?

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Options:

A.

$\frac{100}{40}$

B.

$\frac{100}{35}$

C.

$\frac{100}{32}$

D.

$\frac{100}{26}$

Answer: D

Solution:

$$
\begin{aligned}
& \wedge_0\left(NH_4OH\right)=\wedge_0\left(NH_4Cl\right)+\wedge_0(NaOH)-\wedge_0(NaCl) \\
& =130+213-109 \\
& =234 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}
\end{aligned}
$$

$$\begin{aligned}\text{Percentage dissociation } &= \frac{\wedge_c}{\wedge_0} \times 100 = \frac{9}{234} \times 100 \\ &= \frac{1}{26} \times 100\% \\ &= \frac{100}{26}\%\end{aligned}$$

Question2

What is degree of dissociation of $\mathrm{CH_3COOH}$ if, $\wedge^\circ\left(\mathrm{CH_3COO^-}\right) = 50\ \mathrm{S\ cm^2\ mol^{-1}}$, $\wedge^\circ\left(\mathrm{H^+}\right) = 350\ \mathrm{S\ cm^2\ mol^{-1}}$ and molar conductivity of $5 \times 10^{-2}\mathrm{M\,CH_3COOH}$ is $20\ \mathrm{S\ cm^2\ mol^{-1}}$?

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Options:

A.

1.25×10^{-4}

B.

5×10^{-2}

C.

1.25×10^{-2}

D.

5×10^{-4}

Answer: B

Solution:

$$\begin{aligned}&\wedge_0 = \lambda^\circ\left(\mathrm{CH_3COO^-}\right) + \lambda^\circ\left(\mathrm{H^+}\right) \\ &= 50 + 350 = 400\ \mathrm{S\ cm^2\ mol^{-1}}\end{aligned}$$

$$\alpha = \frac{\wedge_c}{\wedge_0} = \frac{20}{400} = 0.05 = 5 \times 10^{-2}$$

Question3

What is the SI unit of resistivity?

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Options:

A.

$\Omega \mathrm{m}$

B.

$\Omega \mathrm{m}^{-1}$

C.

$\Omega^{-1} \mathrm{~m}^{-1}$

D.

$\Omega \mathrm{m}^{-2}$

Answer: A

Solution:

Resistivity ρ is related to resistance R by:

$R = \rho \frac{L}{A}$

where

- R has unit ohm (Ω)
- L is length (m)
- A is cross-sectional area $\left(\mathrm{m}^2\right)$

So,

$\rho = R \cdot \frac{A}{L}$

Step 2: Substitute units

$[\rho] = [\Omega] \cdot \frac{m^2}{m} = \Omega \cdot m$

Final Answer:

The SI unit of resistivity is $\Omega\, m$

✅ Option A: $\Omega\,\mathrm{m}$

Question4

What is reduction potential of hydrogen gas electrode when pure hydrogen gas is at I atmospheric pressure and platinum electrode is in contact with HCl solution of pH 1 at 298 K ?

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Options:

A.

-0.1184 V

B.

-0.0592 V

C.

-0.0296 V

D.

-0.592 V

Answer: B

Solution:

Step 1: Write the half-cell reaction

The reaction at the hydrogen gas electrode is:$\mathrm{H}^+ + \mathrm{e}^- \longrightarrow \frac{1}{2}\mathrm{H}_2$

Here, $n = 1$ (one electron is involved), the standard electrode potential $E^{\circ}_{H_2} = 0$, and the pressure of hydrogen gas $P_{H_2} = 1$ atm.

Step 2: Use the Nernst Equation

The reduction potential of the hydrogen electrode is given by:$E_{H_2} = \frac{-2.303RT}{nF} \log_{10} \frac{P_{H_2}^{1/2}}{[H^+]}$

Step 3: Relate pH to $[H^+]$

pH is related to the concentration of hydrogen ions by:$pH = -\log_{10}[H^+]$

Or, $\log_{10} \frac{1}{[H^+]} = pH$

Step 4: Substitute values

Substitute the values into the equation:$E_{H_2} = \frac{-2.303RT}{nF} \times pH$

At 298 K, the value of $\frac{-2.303RT}{nF}$ is approximately -0.0592 V.

Given pH = 1:$E_{H_2} = -0.0592 \times 1 = -0.0592$ V

Question5

What happens to the emf for the cell, $Zn_{(s)} \left| Zn^{+2}_{(1M)} \right\| Ag^{+1}(1M) \left| Ag^{+}_{(s)} \right.$ if concentration of Ag^{+1} decreases to 0.1 M ?

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Options:

A.

increase by 0.0592 V

B.

decrease by 0.0592 V

C.

increase by 0.0296 V

D.

decrease by 0.0296 V

Answer: B

Solution:

Cell reaction:

$$Zn_{(s)} + 2Ag^{+}_{(aq)} \longrightarrow Zn^{+2}_{(aq)} + Ag_{(s)}$$

$\therefore \quad n = 2$

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{2} \log \frac{[Zn^{+2}]}{[Ag^{+}]^{2}}$$

$$(E_{cell})_1 = E^{\circ}_{cell} - \frac{0.0592}{2} \log_{10} \frac{[1]}{[1]^2}$$

$$(E_{cell})_1 = E^{\circ}_{cell} \quad (\because \log 1 = 0)$$

When concentration of Ag^{+} decreases to 0.1 M .

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{n} \log \frac{(1)}{(0.1)^2}$$

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{2} \times \log (10^2)$$

$$= E^{\circ}_{cell} - \frac{0.0592}{2} \times 2$$

$$= (E_{cell})_1 - 0.0592$$

Above equation indicates the decrease in emf of cell by 0.0592 V when the concentration of Ag^{+} decreased from 1 M to 0.1 M .

Question6

Which of the following net cell reaction takes place in a galvanic cell containing copper electrode and standard hydrogen electrode?

$$E^{o}\left(Cu^{+2}_{(aq)} \mid Cu_{(s)}\right) = +0.34 \text{ V}$$

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Options:

A.

$$Cu_{(s)} + 2H^{+}_{(aq)} \longrightarrow Cu^{+2}_{(aq)} + H_{2(g)}$$

B.

$$\mathrm{H}_{2(\mathrm{~g})}+\mathrm{Cu}_{(\mathrm{aq})}^{+2} \longrightarrow 2 \mathrm{H}_{(\mathrm{aq})}^{+}+\mathrm{Cu}_{(\mathrm{s})}$$

C.

$$\mathrm{Cu}_{(\mathrm{s})}+\mathrm{H}_{2(\mathrm{~g})} \longrightarrow \mathrm{Cu}_{(\mathrm{aq})}^{+2}+2 \mathrm{H}_{(\mathrm{aq})}^{+}$$

D.

$$\mathrm{Cu}_{(\mathrm{aq})}^{+2}+2 \mathrm{H}_{(\mathrm{aq})}^{+} \longrightarrow \mathrm{Cu}_{(\mathrm{s})}+\mathrm{H}_{2(\mathrm{~g})}$$

Answer: B

Solution:

On comparing standard reduction potentials:

$$\mathrm{E}^{o}\left(\mathrm{Cu}_{(\mathrm{aq})}^{+2} \mid \mathrm{Cu}_{(\mathrm{s})}\right)=+0.34 \mathrm{~V}$$
$$\mathrm{E}^{o}\left(\mathrm{H}_{(\mathrm{aq})}^{+2} \mid \mathrm{H}_{2(\mathrm{~s})}\right)=0.00 \mathrm{~V}$$

The electrode (copper) with the higher reduction potential undergoes reduction (cathode) and the electrode (hydrogen) with the lower reduction potential undergoes oxidation (anode)

$\therefore$ The net cell reaction is:

$$\mathrm{H}_{2(\mathrm{~g})}+\mathrm{Cu}_{(\mathrm{aq})}^{+2} \longrightarrow 2 \mathrm{H}_{(\mathrm{aq})}^{+}+\mathrm{Cu}_{(\mathrm{s})}$$

Question7

Which of the following changes takes place at positive electrode during recharging of lead accumulator?

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Options:

A. Pb is oxidised to PbSO_4

B. PbSO_4 is oxidised to PbO_2

C. PbSO_4 is reduced to Pb

D. PbO_2 is reduced to $PbSO_4$

Answer: B

Solution:

During recharging in lead accumulator, the reaction occurs at positive electrode (anode) can be given as:

$PbSO_{4(s)} + 2H_2O_{(l)} \longrightarrow PbO_{2(s)} + 4H^+_{(aq)} + SO^{2-}_{4(aq)} + 2e^-$

Reaction indicates, $Pb(II)SO_4$ is getting oxidised to $Pb(IV)O_2$.

Question8

For the cell reaction, $A_{(s)} + B^{+2}_{(eq)} \longrightarrow A^{+2}_{(eq)} + B_{(s)}$ if equilibrium constant of reaction is 10^4 at 298 K . What is standard emf of cell?

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Options:

A. 0.0592 V

B. 0.1184 V

C. 0.1776 V

D. 0.2368 V

Answer: B

Solution:

$E_{cell} = E^\circ_{cell} - \frac{0.0592}{n} \log Q$

At equilibrium,

$$
\begin{aligned}
& E_{cell} = 0 \text{ and } Q = K \\
\therefore \quad & E^\circ_{cell} = \frac{0.0592}{n} \log K
\end{aligned}
$$

For cell reaction, $A_{(s)} + B^{+2}_{(aq)} \rightarrow A^{+2}_{(aq)} + B_{(s)}$

$$n = 2$$

$$\therefore \quad E^{o}_{cell} = \frac{0.0592}{2}\log_{10}\left(10^{4}\right) = \frac{0.0592}{2} \times 4$$
$$= 0.0592 \times 2 = 0.1184 \text{ V}$$

Question9

If $E^{-}\left(Cu^{+2}_{(aq)} \mid Cu_{(s)}\right) = +0.34$ V**. What is potential for** $Cu_{(s)} \rightarrow Cu^{+2}_{(aq)}(0.1M) + 2e^{-}$ **at 298 K ?**

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Options:

A.

+0.3696 V

B.

-0.3696 V

C.

+0.3104 V

D.

-0.3104 V

Answer: D

Solution:

$$E^{o}\left(Cu_{(s)} \mid Cu^{+2}_{(aq)}\right) = -0.34 \text{ V} \quad \left(E^{o}_{ox} = -E^{o}_{red}\right)$$

$$E_{electrode} = E^{\circ}_{electrode} - \frac{0.0592}{n}\log\left[Cu^{2+}\right]$$

$$= -0.34 - \frac{0.0592}{2}\log(0.1)$$
$$= -0.34 - 0.0296 \times \log 10^{-1}$$
$$= -0.34 - 0.0296 \times (-1)$$
$$= -0.34 + 0.0296$$
$$= -0.3104 \text{ V}$$

Question10

If standard reduction potential (E°) of $\left(Mg^{+2}_{(aq)} \mid Mg_{(s)}\right)$, $\left(Ag^{+}_{(aq)} \mid Ag_{(s)}\right)$, $\left(Zn^{+2}{}_{(aq)} \mid Zn_{(s)}\right)$ and $\left(Cu^{+2}{}_{(aq\ \ )} \mid Cu_{(s\ \ )}\right)$ are -2.37 V, $+0.79$ V, -0.76 V and +0.34 V respectively. Which of the following reaction is spontaneous?

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Options:

A.

$Zn_{(s)} + Mg^{+2}_{(aq)} \longrightarrow Zn^{+2}_{(aq)} + Mg_{(s)}$

B.

$2Ag_{(s)} + Zn^{+2}_{(aq)} \longrightarrow 2Ag^{+}_{(aq)} + Zn_{(s)}$

C.

$Zn_{(s)} + Cu^{+2}_{(aq)} \longrightarrow Zn^{+2}_{(aq)} + Cu_{(s)}$

D.

$Cu_{(s)} + Mg^{+2}_{(aq)} \longrightarrow Cu^{+2}_{(aq)} + Mg_{(s)}$

Answer: C

Solution:

For reaction to be spontaneous, E^0_{Cell} should be positive.

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

$(A)\ E^0_{Cell} = E^0_{(Mg^{+2}|Mg)} - E^0_{(Zn^{+2}|Zn)}$
$= -2.37\ V - (-0.76)V = -1.61\ V$

$(B)\ E^0_{Cell} = E^0_{(Zn^{+2}|Zn)} - E^0_{(Ag^{+}|Ag^{-})}$
$= -0.76\ V - (+0.79)V = -1.55\ V$

$(C)\ E^0_{Cell} = E^0_{(Cu^{+2}|ICu)} - E^0_{(Zn^{+2}|Zn)}$
$= 0.34\ V - (-0.76)V = +1.1\ V$

$(D)\ E^0_{Cell} = E^0_{(Mg^{+2}{}_{1Mg})} - E^0_{(Cu^{+2}|Cu)}$
$= -2.37\ V + 0.34\ V = -2.03\ V$

Only the reaction in Option C has a positive standard cell potential, indicating it is spontaneous.

Question11

What is molar conductivity at zero concentration in $\Omega^{-1}\ cm^2\ mol^{-1}$ for aluminium sulphate, if molar ionic conductivities at zero concentration of Al^{+3} and SO_4^{-2} are $189\Omega^{-1}\ cm^2\ mol^{-1}$ and $50.1\Omega^{-1}\ cm^2\ mol^{-1}$ respectively?

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Options:

A. 239.1

B. 428.1

C. 478.2

D. 528.3

Answer: D

Solution:

$$\wedge^0 (Al_2(SO_4)_3) = 2 \times \lambda^0_{N^{3+}} + 3 \times \lambda^0_{SO_4^{2-}}$$

$$\therefore \quad = 2 \times 189 + 3 \times 50.1$$

$$= 378 + 150.3$$

$$= 528.3 \Omega^{-1} \text{ cm}^2 \text{ mol}^{-1}$$

Question12

For the cell involving following reaction.

$$Zn_{(s)} + Ni^{+2}_{(aq)} \longrightarrow Zn^{+2}_{(aq)} + Ni_{(s)}, E^*_{cell} = 0.5 \text{ V}.$$

What is standard Gibb's energy change of cell reaction?

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Options:

A. -193 kJ

B. -905 kJ

C. -96.5 kJ

D. -89.65 kJ

Answer: C

Solution:

The cell reaction:

$$Zn_{(s)} + Ni^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Ni_{(s)}, \quad E^\circ_{cell} = 0.50 \text{ V}.$$

We want the standard Gibbs free energy change, ΔG°.

Step 1: Recall the relation

$$\Delta G^\circ = -nFE^\circ_{cell} .$$

- n = number of moles of electrons transferred.
- F = Faraday constant $= 96,500 \text{ C mol}^{-1}$.

- $E^\circ_{\text{cell}} = 0.50$ V.

Step 2: Identify electrons exchanged

Reaction:

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

So, $n = 2$.

Step 3: Substitute values

$\Delta G^\circ = -nFE^\circ_{\text{cell}} = -(2)(96,500 \text{ C/mol})(0.50 \text{ V})$.

Step 4: Compute

$\Delta G^\circ = -(2)(96,500)(0.50)$.

$= -96,500 \text{ J mol}^{-1}$.

$= -96.5 \text{ kJ mol}^{-1}$.

Final Answer:

$\boxed{-96.5 \text{ kJ}}$

Correct option: C

Question13

If $E^\circ \left(Zn^{+2}_{(aq)} \mid Zn_{(s)}\right) = -0.76$ V.

Calculate potential for $Zn_{(s)} \rightarrow Zn^{+2}(0.01M) + 2e^-$ at 298 K .

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Options:

A. +0.8192 V

B. -0.8192 V

C. +0.7008 V

D. -0.7008 V

Answer: A

Solution:

$E^0\left(Zn_{(s)} \mid Zn^{+2}_{(eq)}\right) = +0.76\ V$

$$\begin{aligned} E_{electrode} &= E^0_{electrode} - \frac{0.0592}{n}\log\left[Zn^{2+}\right] \\ &= +0.76 - \frac{0.0592}{2}\log(0.01) \\ &= +0.76 - \frac{0.0592}{2}(-2) = +0.8192\ V \end{aligned}$$

Question14

Which from following statements is true regarding the cell emf at 298 K for

$$^{\ominus}Ni_{(s)} \mid \overset{+2}{N}i(0.01M) \| \overset{+}{A}g\ (0.01M) \mid \overset{\oplus}{A}g_{(s)}$$

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Options:

A. less than E°_{cell} by 0.0592 V

B. greater than E°_{cell} by 0.0592 V

C. less than E°_{cell} by 0.0296 V

D. greater than E°_{cell} by 0.0296 V

Answer: A

Solution:

$Ni_{(s)} + 2Ag^+_{(aq)} \longrightarrow Ni^{2+}_{(aq)} + 2Ag_{(s)}$

For the given cell reaction, $n = 2$.

Using Nernst equation,

$$\begin{aligned}
E_{cell} &= E^{\circ}_{cell} - \frac{0.0592}{n}\log_{10} Q \\
E_{cell} &= E^{\circ}_{cell} - \frac{0.0592}{2}\log_{10} \frac{[Ni^{2+}]}{[Ag^{+}]^{2}}
\end{aligned}$$

Substituting the given values,

$$\begin{aligned}
E_{cell} &= E^{0}_{cell} - \frac{0.0592}{2}\log_{10} \frac{10^{-2}}{(10^{-2})^{2}} \\
&= E^{0}_{cell} - \frac{0.0592}{2}\log_{10} 10^{2} \\
&= E^{0}_{cell} - \frac{0.0592}{2} \times 2 \times 1 \\
&(\because \log_{10} 10^{2} = 2\log_{10} 10 \text{ and } \log_{10} 10 = 1) \\
\therefore \quad E_{cell} &= E^{0}_{cell} - 0.0592 \text{ V}
\end{aligned}$$

Question15

The standard emf for cell, $^{\ominus}Cd_{(s)} \left| {}^{+2}Cd(1M) \right\| \overset{+2}{Cu}\ (1M) \left| Cu_{(s)} \right.^{\oplus}$ is 0.74 V .

If concentration of $Cd^{+2}_{(aq)}$ and $Cu^{+2}_{(aq)}$ decreases by 10 times at 298 K . Calculate emf of cell.

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Options:

A. +0.074 V

B. +0.850 V

C. +0.680 V

D. +0.740 V

Answer: D

Solution:

$\mathrm{Cd}_{(s)} + \mathrm{Cu}^{2+}_{(aq)} \longrightarrow \mathrm{Cd}^{2+}_{(aq)} + \mathrm{Cu}_{(s)}$

For the given cell reaction, $n = 2$.

Using Nernst equation,

$$\mathrm{E}_{\text{cell}} = \mathrm{E}^{\circ}_{\text{cell}} - \frac{0.0592}{2}\log_{10}\frac{[\mathrm{Cd}^{2+}]}{[\mathrm{Cu}^{2+}]}$$

Substituting the given values,

$$\begin{aligned} \mathrm{E}_{\text{cell}} &= 0.74 - \frac{0.0592}{2}\log_{10}\frac{1}{1} \\ \therefore \mathrm{E}_{\text{cell}} &= +0.740 \text{ V} \quad (\because \log_{10} 1 = 0) \end{aligned}$$

Now, $[\mathrm{Cd}^{2+}]$ and $[\mathrm{Cu}^{2+}]$ are both decreased by 10 times.

$$\begin{aligned} &[\mathrm{Cd}^{2+}] = [\mathrm{Cu}^{2+}] = 0.1\mathrm{M} \\ &\mathrm{E}_{\text{cell}} = 0.74 - \frac{0.0592}{2}\log_{10}\frac{0.1}{0.1} = +0.740 \text{ V} \end{aligned}$$

Thus, there is no change in E_{cell} as both $[\mathrm{Cd}^{2+}]$ and $[\mathrm{Cu}^{2+}]$ are changed such that the value of Q is overall unchanged.

Question16

A hypothetical galvanic cell is $^{\ominus}\mathrm{A}_{(s)}\ |\overset{+}{\mathrm{A}}\ (1\mathrm{M})||\ \overset{+2}{\mathrm{B}}\ (1\mathrm{M})||\overset{\oplus}{\mathrm{B}}_{(s)}$ and emf of cell is positive. What is the possible cell reaction?

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Options:

A.

$\mathrm{A}_{(s)} + \overset{+B}{B}_{(aq)} \longrightarrow \overset{+1}{A}_{(aq)} + \mathrm{B}_{(s)}$

B.

$2\ \mathrm{A}_{(s)} + \overset{+}{\mathrm{B}}_{(aq)} \longrightarrow 2\ \overset{+1^{+}}{\mathrm{A}}_{(aq)} + \mathrm{B}_{(s)}$

C.

$$\mathrm{A_{(s)}} + 2\overset{+2B}{(\mathrm{aq})}^{+2} \xrightarrow{+1} \mathrm{A_{(aq)}} + 2\ \mathrm{B_{(s)}}$$

D.

$$2\overset{+}{\mathrm{A}}_{\mathrm{(aq)}} + \mathrm{B_{(s)}} \longrightarrow 2\ \mathrm{A_{(s)}} + \overset{+2}{\mathrm{B}}_{\mathrm{(aq)}}$$

Answer: B

Solution:

We are asked about a galvanic cell of the form:

$^{\ominus}A_{(s)} \mid A^{+}(1M) \mid\mid B^{2+}(1M) \mid B^{\oplus}_{(s)}$

Step 1: Interpret the notation

- The **left electrode** is $A_{(s)} \mid A^{+}(aq)$.
- This means $A_{(s)} \rightarrow A^{+}(aq) + e^{-}$. So **A is the anode** (oxidized).
- The **right electrode** is $B^{2+}(aq) \mid B_{(s)}$.
- This means $B^{2+}(aq) + 2e^{-} \rightarrow B(s)$. So **B is the cathode** (reduced).
- Since emf is positive, the reaction written this way occurs spontaneously.

Step 2: Write half-reactions

At anode (oxidation):

$A(s) \ \rightarrow \ A^{+}(aq) + e^{-}$

At cathode (reduction):

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

Step 3: Balance electrons

- The anode gives 1 electron per atom of A.
- The cathode requires 2 electrons per atom of B.

To balance, multiply the anode reaction by 2:

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

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

Step 4: Combine

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

Step 5: Match with given options

This corresponds to **Option B**:

$2A_{(s)} + B^{+}_{(aq)} \longrightarrow 2A^{+}_{(aq)} + B_{(s)}$

(Here the notation in option B is slightly garbled, but clearly it matches the balanced reaction.)

✅ **Final Answer: Option B**

Question17

Consider galvanic cell, $^{\ominus}A_{(s)} \left| A^{+2}_{(aq)} \right| \left| B^{+}_{(aq)} \right| B^{\oplus}_{(s)}$

if emf of cell is positive, identify a correct cell reaction from following.

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Options:

A. $A_{(s)} + B^{+2}_{(aq)} \longrightarrow A^{+2}_{(aq)} + B_{(s)}$

B. $A_{(s)} + 2\,B^{+}_{(aq)} \longrightarrow A^{+2}_{(aq)} + 2\,B_{(s)}$

C.

$2\,A^{+2}_{(aq)} + B_{(s)} \longrightarrow 2\,A_{(s)} + B^{+}_{(aq)}$

D.

$2\,A_{(s)} + B^{+}_{(aq)} \longrightarrow 2\,A^{+2}_{(aq)} + B_{(s)}$

Answer: B

Solution:

We're given a galvanic cell notation:

$^{\ominus}\,A_{(s)} \mid A^{2+}_{(aq)} \parallel B^{+}_{(aq)} \mid B^{\oplus}_{(s)}$

Step 1. Recall cell notation rules

- Left-hand electrode (anode, marked by ⊖) → oxidation occurs.
- Right-hand electrode (cathode, marked ⊕) → reduction occurs.

So:

- At anode (left): $\mathrm{A}(s) \rightarrow \mathrm{A}^{2+}(aq) + 2e^{-}$
- At cathode (right): $\mathrm{B}^{+}(aq) + e^{-} \rightarrow \mathrm{B}(s)$

Step 2. Balance the electrons

- Oxidation: $\mathrm{A}(s) \rightarrow \mathrm{A}^{2+}(aq) + 2e^{-}$
- Reduction: $2\,\mathrm{B}^{+}(aq) + 2e^{-} \rightarrow 2\,\mathrm{B}(s)$

Step 3. Add the half–reactions

$$\mathrm{A}(s) + 2\,\mathrm{B}^{+}(aq) \quad \longrightarrow \quad \mathrm{A}^{2+}(aq) + 2\,\mathrm{B}(s)$$

Step 4. Match with given options

This matches **Option B** exactly.

✅ **Correct Answer: Option B**

Cell reaction:

$$\mathrm{A}_{(s)} + 2\,\mathrm{B}^{+}_{(aq)} \quad \rightarrow \quad \mathrm{A}^{2+}_{(aq)} + 2\,\mathrm{B}_{(s)}$$

Question18

Standard potential of electrode reaction $\mathrm{Cu}^{+2}_{(\mathrm{aq})} + 2\mathrm{e}^{-} \longrightarrow \mathrm{Cu}_{(\mathrm{s})}$ is +0.34 V .

What is standard potential of reaction

$$2\mathrm{Cu}_{(\mathrm{s})} \longrightarrow 2\mathrm{Cu}^{+2}_{(\mathrm{aq})} + 4\mathrm{e}^{-}?$$

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Options:

A. +0.68 V

B. -0.68 V

C. +0.34 V

D. -0.34 V

Answer: D

Solution:

The reaction $2Cu_{(s)} \longrightarrow 2Cu^{+2}_{(aq)} + 4e^-$ is the reverse of the given reaction $Cu^{+2}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$ which has a standard potential of $+0.34$ V.

When you reverse a reaction, the sign of the standard potential (E°) also changes. So, the standard potential for the oxidation reaction is -0.34 V.

Because you have exactly double the number of copper atoms, but E° does not change with the number of electrons or atoms, the standard potential stays -0.34 V for the whole reaction.

Question19

What is value of E_{cell} at 298 K for the reaction, $Zn_{(s)} + Cu^{+2}(0.1M) \longrightarrow Zn^{+2}(0.1M) + Cu_{(s)}$ if $E^\circ_{\text{cell}} = 1.1$ V ?

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Options:

A. 1.1 V

B. 0.11 V

C. 1.0408 V

D. 0.0296 V

Answer: A

Solution:

The reaction is

$$\mathrm{Zn_{(s)} + Cu^{2+}_{(0.1M)} \longrightarrow Zn^{2+}_{(0.1M)} + Cu_{(s)}}$$

$$\mathrm{E_{cell}} = \mathrm{E^0_{cell}} - \frac{0.0592\ \mathrm{V}}{\mathrm{n}} \log_{10} \frac{[\mathrm{Zn^{2+}}]}{[\mathrm{Cu^{2+}}]}$$

$$\therefore\ \mathrm{E_{cell}} = 1.1\ \mathrm{V} - \frac{0.0592\ \mathrm{V}}{2} \log_{10} \frac{(0.1)}{(0.1)} = 1.1\ \mathrm{V} - 0$$

$$\therefore\ \mathrm{E_{cell}} = 1.1\ \mathrm{V}$$

Question20

If $\mathrm{E^\circ}\left(\mathrm{Fe^{+2}_{(aq)}} \mid \mathrm{Fe_{(s)}}\right) = -0.44\ \mathrm{V}$ and

$\mathrm{E^\circ}\left(\mathrm{Sn^{+2}_{(aq)}} \mid \mathrm{Sn_{(s)}}\right) = -0.14\ \mathrm{V}$

What is standard emf of cell containing the two electrodes?

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Options:

A. +0.30 V

B. -0.30 V

C. +0.58 V

D. -0.58 V

Answer: A

Solution:

We are tasked with finding the standard emf for the electrochemical cell:

$\mathrm{E^\circ(Fe^{2+}/Fe) = -0.44\,V}$

$\mathrm{E^\circ(Sn^{2+}/Sn) = -0.14\,V}$

Step 1: Identify which is the cathode and which is the anode

- The standard reduction potential (E°) shows how readily a species is reduced.

- The electrode with **higher** E° acts as the **cathode** (reduction).
- The electrode with **lower** E° acts as the **anode** (oxidation).

Here:

- $E^\circ(\mathrm{Sn}^{2+}/\mathrm{Sn}) = -0.14\ \mathrm{V}$ (higher, so **cathode** / reduction).
- $E^\circ(\mathrm{Fe}^{2+}/\mathrm{Fe}) = -0.44\ \mathrm{V}$ (lower, so **anode** / oxidation).

Step 2: Write the half-reactions

- Oxidation (anode): $\mathrm{Fe}(s) \rightarrow \mathrm{Fe}^{2+}(aq) + 2e^-$
- Reduction (cathode): $\mathrm{Sn}^{2+}(aq) + 2e^- \rightarrow \mathrm{Sn}(s)$

Step 3: Apply the formula

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

$$E^\circ_{\text{cell}} = (-0.14) - (-0.44)$$

$$E^\circ_{\text{cell}} = -0.14 + 0.44 = +0.30\ \mathrm{V}$$

✅ Final Answer:

The standard emf of the cell is:

Option A: +0.30 V

Question21

If $E^o\left(\mathrm{AI}^{+3}_{(eq)} \mid \mathrm{AI}_{(s)}\right) = -1.66\ \mathrm{V}$. What is potential of $\mathrm{Al}_{(s)} \longrightarrow \mathrm{Al}^{+3}(0 \cdot 1\mathrm{M}) + 3\mathrm{e}^-$ at 298 K ?

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Options:

A. +1.540 V

B. -1.540 V

C. +1.679 V

D. -1.679 V

Answer: C

Solution:

The given value, $E^o\left(\mathrm{Al}_{(s)} \mid \mathrm{Al}^{3+}_{(aq)}\right) = -1.66$ V, is the standard potential for the reaction when the concentration of Al^{3+} is 1 M.

We are asked to find the potential when Al^{3+} is at 0.1 M at 298 K.

We use the Nernst equation to find the new potential:

$$E_{\text{electrode}} = E^o_{\text{electrode}} - \frac{0.0592}{n}\log_{10}\frac{[\text{product}]}{[\text{reactant}]}$$

For this reaction: $\mathrm{Al}_{(s)} \longrightarrow \mathrm{Al}^{3+}(0.1\ \mathrm{M}) + 3e^-$, the only ion to consider is Al^{3+} because solid Al does not count in the concentration ratio.

This makes the formula:

$$E_{\text{electrode}} = E^o_{\text{electrode}} - \frac{0.0592}{3}\log_{10}[\mathrm{Al}^{3+}]$$

Putting $[\mathrm{Al}^{3+}] = 0.1$ and $n = 3$ into the equation:

$$E_{\text{electrode}} = 1.66 - \frac{0.0592}{3}\log_{10}(0.1)$$

Because $\log_{10}(0.1) = -1$:

$$E_{\text{electrode}} = 1.66 - (0.0197 \times -1)$$

So:

$$E_{\text{electrode}} = 1.66 + 0.0197 = +1.679\ \mathrm{V}$$

Question22

Which of the following reactions exhibit minimum standard reduction potential?

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Options:

A. $2\mathrm{H}^+_{(\mathrm{aq})} + 2\mathrm{e}^- \longrightarrow \mathrm{H}_{2(\mathrm{~g})}$

B. $F_{2(g)} + 2e^- \longrightarrow 2\,F^-_{(aq)}$

C. $Li^+_{(aq)} + e^- \longrightarrow Li_{(s)}$

D. $Cl_{2(g)} + 2e^- \longrightarrow 2Cl^-_{(aq)}$

Answer: C

Solution:

Half reaction	Standard reduction potential
$2H^+ + 2e^- \longrightarrow H_2$	0.00 V
$F_2 + 2e^- \longrightarrow 2\,F^-$	+2.870 V
$Li^+ + e^- \longrightarrow Li$	-3.045 V
$Cl_2 + 2e^- \longrightarrow 2Cl^-$	+1.360 V

Lithium (Li) prefers to lose electrons and form Li^+. In other words, it prefers to get oxidised which is a characteristic feature of alkali metals, This makes its reduction potential very low compared to other elements.

Question23

If standard reduction potential (E°) of $\left(Ni^{+2}_{(aq)} \mid Ni_{(s)}\right)$ and $\left(Al^{+3}_{(aq)} \mid Al_{(s)}\right)$ are -0.25 V and -1.66 V respectively. What is standard emf of cell reaction $2Al_{(s)} + 3Ni^{+2}_{(aq)} \longrightarrow 2Al^{+3}_{(aq)} + 3Ni_{(s)}$

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Options:

A. +2.57 V

B. -2.57 V

C. +1.41 V

D. -1.91 V

Answer: C

Solution:

We are tasked with finding the standard emf of the galvanic cell:

$$2\,Al_{(s)} + 3\,Ni^{2+}_{(aq)} \;\rightarrow\; 2\,Al^{3+}_{(aq)} + 3\,Ni_{(s)}$$

Step 1. Write the half-cell data

$$E^\circ(Ni^{2+}/Ni) = -0.25\,V$$

$$E^\circ(Al^{3+}/Al) = -1.66\,V$$

These are *reduction potentials*.

Step 2. Determine which undergoes oxidation vs reduction

- Higher (less negative) $E^\circ \rightarrow$ more likely reduction.
- Lower (more negative) $E^\circ \rightarrow$ more likely oxidation.

Thus:

- Ni^{2+} will be reduced to Ni.
- Al will be oxidized to Al^{3+}.

This matches the given reaction. ✅

Step 3. Standard cell potential formula

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

- Cathode (reduction): Ni^{2+}/Ni, $E^\circ = -0.25$ V.
- Anode (oxidation): Al/Al^{3+}. The tabulated value is for $Al^{3+} + 3e^- \rightarrow Al$, with $E^\circ = -1.66$ V. We can directly use this in formula.

$$E^\circ_{cell} = -0.25 - (-1.66) = +1.41\,V$$

Step 4. Select answer

The standard emf of the given cell is:

$$\boxed{+1.41\,V}$$

✅ Correct Option: **C. +1.41 V**

Question24

If $E^{\cdot}\left(Mg^{+2}_{(aq)} \mid Mg_{(s)}\right) = -2.37\ V$**. What is potential for** $Mg_{(s)} \longrightarrow Mg^{+2}(0.01M) + 2e^{-}$ **at 298 K ?**

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Options:

A. +2.3108 V

B. -2.3108 V

C. +2.4292 V

D. -2.4292 V

Answer: C

Solution:

$E^{p}\left(Mg_{(s)} \mid Mg^{+2}_{(aq)}\right) = +2.37\ V$

$(\because$ reduction potential is given in the question)

$E_{electrode} \quad = E^{\circ}_{electrode} \quad - \frac{0.0592}{n} \log\left[Mg^{2+}\right]$

$$\begin{aligned}
&= +2.37 - \frac{0.0592}{2} \log(0.01) \\
&= +2.37 - \frac{0.0592}{2} \times \log 10^{-2} \\
&= +2.37 - \frac{0.0592}{2} \times (-2) \\
&= +2.37 + 0.0592 \\
&= +2.4292\ V
\end{aligned}$$

Question25

If $E^{\circ}\left(Ag^{+}{}_{(aq)} \mid Ag_{(s)}\right) = +0.80\ V$ **What is potential developed for** $Ag_{(s)} \longrightarrow Ag^{+1}(0.01M) + e^{-}$ **at 298 K ?**

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Options:

A. +0.9184 V

B. -0.9184 V

C. +0.6816 V

D. -0.6816 V

Answer: D

Solution:

Nernst equation can be used to calculate electrode potential.

$$\begin{aligned} E_{\text{electrode}} &= E^{\circ}_{\text{electrode}} - \frac{0.0592\ \text{V}}{\text{n}} \log_{10} \frac{[\text{ Product }]}{[\text{ Reactant }]} \\ E_{\text{electrode}} &= E^{\circ}_{\text{electrode}} - \frac{0.0592\ \text{V}}{1} \log_{10} \left[\text{Ag}^{+}\right] \end{aligned}$$

The standard oxidation potential for given electrode will be 0.80 V

$$\begin{aligned} \therefore \quad \text{Eelectrode} &= -0.8\ \text{V} - \frac{0.0592\ \text{V}}{1} \log_{10} 0.01 \\ &= -0.8\ \text{V} + 0.0592\ \text{V} \times (2) \quad \left[\because \log_{10}\left(10^{-2}\right) = -2\right] \\ &= -0.8\ \text{V} + 0.1184 \\ &= -0.6816\ \text{V} \end{aligned}$$

Question26

Which of the following relations about $\text{E}^{\circ}_{\text{cell}}$ is false?

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Options:

A. $\text{E}^{\circ}_{\text{cell}} = \text{E}^{\circ}_{(\text{cathode})} - \text{E}^{\circ}_{(\text{anode})}$

B. $E^\circ_{cell} = \frac{0.0592}{n} \times \log_{10} K$

C. $E^\circ_{cell} = \frac{\Delta G^\circ}{nF}$

D. $E^\circ_{cell} = E^\circ_{(anode)} + E^\circ_{(cathode)}$

Answer: C

Solution:

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

Question27

For cell reaction,

$$2Al_{(s)} + 3Cu^{+2}_{(aq)} \longrightarrow 2Al^{+3}_{(aq)} + 3Cu_{(s)}$$

If $\Delta G^\circ = -1158$ kJ, what is E°_{cell} ?

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Options:

A. 3 V

B. 2 V

C. 2.5 V

D. 1.5 V

Answer: B

Solution:

Reaction:

$$2Al_{(s)} + 3Cu^{2+}_{(aq)} \longrightarrow 2Al^{3+}_{(aq)} + 3Cu_{(s)}$$

Given:

$\Delta G^\circ = -1158 \text{ kJ} = -1,158,000 \text{ J}$

We need standard cell potential E°_{cell} .

Step 1: Relation between ΔG° and E°_{cell}

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

where:

- n = total moles of electrons transferred
- $F \approx 96485 \text{ C/mol}$

Step 2: Electrons transferred in this reaction

- Al goes from 0 to +3: each Al loses 3 e^-, and there are 2 Al → total 6 electrons.
- Cu goes from +2 to 0: each Cu^{2+} gains 2 e^-, and there are 3 Cu^{2+} → total 6 electrons.

So $n = 6$.

Step 3: Solve for E°_{cell}

$E^\circ_{\text{cell}} = -\frac{\Delta G^\circ}{nF} = -\frac{-1{,}158{,}000}{6\times 96485}$

$= \frac{1{,}158{,}000}{578910} \approx 2.0 \text{ V}$

Final Answer:

$E^\circ_{\text{cell}} \approx 2 \text{ V}$

✅ **Correct option: B (2 V)**

Question28

Select false statement from the following regarding electrochemical cell.

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Options:

A. Standard cell potential of electro chemical cell is an intensive property.

B. Electrode potential depends on concentration of salt solution in contact with electrode.

C. Standard free energy change of cell is an intensive property.

D. Electrical work done in galvanic cell is equal to decrease in Gibb's energy.

Answer: C

Solution:

Let us analyse each statement:

Option A: "Standard cell potential of electrochemical cell is an intensive property."

- **Intensive properties** do not depend on the amount of the substance. Standard cell potential (E°_{cell}) is indeed **intensive**.

Option B: "Electrode potential depends on concentration of salt solution in contact with electrode."

- Electrode potential **does** depend on the concentration (see Nernst equation), so this statement is **true**.

Option C: "Standard free energy change of cell is an intensive property."

- **Standard free energy change** (ΔG°) depends on the number of moles of electrons (n), as

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

  So, ΔG° **is an extensive property**, not intensive.

Option D: "Electrical work done in galvanic cell is equal to decrease in Gibb's energy."

- This is **true**: Electrical work = $-\Delta G$.

Conclusion:

The **false statement** is:

$\boxed{\text{Option C}}$

Reason: Standard free energy change (ΔG°) is an extensive property, not intensive.

Question29

If $\mathrm{E}^* \left(\mathrm{Mg^{+2}}_{(\mathrm{aq})} \mid \mathrm{Mg}_{(\mathrm{s})}\right) = -2.37\ \mathrm{V}$.

What is potential for $\mathrm{Mg}_{(\mathrm{s})} \rightarrow \mathrm{Mg}^{+2}(0.1\mathrm{M}) + 2\mathrm{e}^-$ **at 298 K ?**

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Options:

A. +2.3996 V

B. -2.3996 V

C. +2.3404 V

D. -2.3404 V

Answer: A

Solution:

$$E_{\text{electrode}} = E^{\circ}_{\text{electrode}} - \frac{0.0592\ \text{V}}{n} \log_{10} \frac{[\text{Product}]}{[\text{Reactant}]}$$

Electrode reaction:

$\mathrm{Mg(s)} \longrightarrow \mathrm{Mg^{2+}(0.1M)} + 2\mathrm{e}^{-}$

Using formula,

$$\begin{aligned} \mathrm{E}_{\left(\mathrm{Mg}|\mathrm{Mg}^{+2}\right)} &= \mathrm{E}^{\circ}_{\left(\mathrm{Mg}|\mathrm{Mg}^{+2}\right)} - \frac{0.0592\ \mathrm{V}}{\mathrm{n}} \log_{10}\left[\mathrm{Mg}^{2+}\right] \\ &= 2.37\ \mathrm{V} - \frac{0.0592\ \mathrm{V}}{2} \log_{10}(0.1) \\ &= 2.37\ \mathrm{V} - \frac{0.0592\ \mathrm{V}}{2} \times (-1) \\ &= 2.37\ \mathrm{V} + 0.0296\ \mathrm{V} \\ &= 2.3996\ \mathrm{V} \end{aligned}$$

Question30

If standard emf of cell $\mathrm{Zn_{(s)}} \left| \mathrm{Zn^{+2}_{(iM)}} \right| \mathrm{Ag^{+1}_{(iM)}} \mid \mathrm{Ag_{(s)}}$ is 1.55 V .

What is electrical work done under standard conditions?

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Options:

A. -144.750 kJ

B. -193.00 kJ

C. -299.150 kJ

D. -386.00 kJ

Answer: C

Solution:

Given:

- Standard emf of cell, $E^\circ_{cell} = 1.55\,\text{V}$
- Cell: $\text{Zn}_{(\text{s})}|\text{Zn}^{2+}_{(1\,\text{M})}||\text{Ag}^{+}_{(1\,\text{M})}|\text{Ag}_{(\text{s})}$

Step 1: Write the cell reaction and find n

$$\text{At anode (oxidation): } \text{Zn(s)} \rightarrow \text{Zn}^{2+}\text{(aq)} + 2e^-$$
$$\text{At cathode (reduction): } \text{Ag}^{+}\text{(aq)} + e^- \rightarrow \text{Ag(s)}$$

To balance electrons, multiply the cathode reaction by 2:

$$2\,\text{Ag}^{+}\text{(aq)} + 2e^- \rightarrow 2\,\text{Ag(s)}$$

Overall cell reaction:

$$\text{Zn(s)} + 2\,\text{Ag}^{+}\text{(aq)} \rightarrow \text{Zn}^{2+}\text{(aq)} + 2\,\text{Ag(s)}$$

So, number of electrons transferred, $n = 2$.

Step 2: Formula for maximum electrical work

Electrical work done under standard conditions is:

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

Where,

- n = number of moles of electrons exchanged $(= 2)$
- F = Faraday's constant $(= 96,500$ C/mol)
- E°_{cell} = emf of the cell $(= 1.55$ V)

Step 3: Substitute and calculate

$$\Delta G^\circ = -2 \times 96,500 \times 1.55$$

$$\Delta G^\circ = -2 \times 149,575 = -299,150\ \text{J}$$

Convert J to kJ:

$$\Delta G^\circ = \frac{-299,150}{1000}\ \text{kJ} = -299.15\ \text{kJ}$$

The correct answer is:

$\boxed{-299.150 \text{ kJ}}$

Option C is correct.

Question31

Assume the cell reaction,

$$\mathrm{A_{(s)} + B^{+2}_{(aq)} \longrightarrow A^{+2}_{(aq)} + B_{(s)}}$$

if $\Delta G^\circ = -386 \text{ kJ}$ at 298 K . What is $\mathrm{E^0_{Cell}}$?

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Options:

A. 1 V

B. 1.5 V

C. 2 V

D. 2.5 V

Answer: C

Solution:

Given:

- Cell reaction:

  $$\mathrm{A_{(s)} + B^{+2}_{(aq)} \longrightarrow A^{+2}_{(aq)} + B_{(s)}}$$

- $\Delta G^\circ = -386 \text{ kJ} = -386 \times 10^3$ J at $T = 298$ K.

We are to find E°_{cell} .

Step 1: Write the relationship between ΔG° and E°_{cell} :

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

Where:

- $n =$ number of electrons exchanged,
- $F =$ Faraday constant $= 96500$ C~mol^{-1}

Step 2: Find the value of n

From the reaction:

$A_{(s)} \rightarrow A^{2+}_{(aq)} + 2e^-$

$B^{2+}_{(aq)} + 2e^- \rightarrow B_{(s)}$

So, $n = 2$ electrons are exchanged.

Step 3: Substitute the given values in the equation

$-386 \times 10^3 \text{ J} = -2 \times 96500 \text{ C~mol}^{-1} \times E^\circ_{\text{cell}}$

Step 4: Solve for E°_{cell}

$386 \times 10^3 = 2 \times 96500 \times E^\circ_{\text{cell}}$

$E^\circ_{\text{cell}} = \frac{386 \times 10^3}{2 \times 96500}$

$E^\circ_{\text{cell}} = \frac{386000}{193000}$

$E^\circ_{\text{cell}} = 2.0 \text{ V}$

Final Answer:

Option C

$\boxed{2 \text{ V}}$

Question32

Which of the following change takes place at positive electrode during working of dry cell?

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Options:

A. $\text{Zn}_{(s)}$ is oxidised to $\text{Zn}^{+2}{}_{(aq)}$

B. $\text{Zn}^{+2}_{(eq)}$ is reduced to $\text{Zn}_{(s)}$

C. $MnO_{2(s)}$ is reduced to $Mn_2O_{3(s)}$

D. $Mn_2O_{3(s)}$ is oxidised to $MnO_{2(s)}$

Answer: C

Solution:

Let's carefully analyse step by step.

1. **Dry cell construction (NCERT):**

- Anode (negative electrode): Zinc container.
- Cathode (positive electrode): Carbon rod surrounded by a mixture of manganese dioxide (MnO_2) and carbon.

1. **At the anode (negative electrode):**

   Zinc gets **oxidised**:

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

2. **At the cathode (positive electrode):**

   Manganese dioxide **gets reduced** in the presence of ammonium ions:

   $MnO_2(s) + NH_4^+(aq) + e^- \rightarrow MnO(OH)(s) + NH_3(aq)$

   On simplifying, we can say:

- Mn in MnO_2 is in +4 state.
- It reduces to +3 oxidation state (as in Mn_2O_3 or $MnO(OH)$).

   So, the change at the **positive electrode (cathode)** is:

   MnO_2 is reduced to Mn_2O_3

1. **Correct option:**

   Option C: $MnO_{2(s)}$ is reduced to $Mn_2O_{3(s)}$ ✅

Question33

Which from following statement is correct regarding emf of cell for the cell reaction,

$$Cd_{(s)} + Cu^{+2}_{(aq)} \longrightarrow Cd^{+2}_{(aq)} + Cu_{(s)}$$

if concentration of Cd^{+2} is 10 times greater than concentration of $\mathrm{Cu}^{+2}_{(\mathrm{aq})}$ at 298 K .

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Options:

A. greater than $\mathrm{E}^{0}_{\text{cell}}$ cell by 0.0592 V

B. less than $\mathrm{E}^{0}_{\text{cell}}$ by 0.0592 V

C. greater than $\mathrm{E}^{0}_{\text{cell}}$ by 0.0296 V

D. less than $\mathrm{E}^{0}_{\text{cell}}$ by 0.0296 V

Answer: D

Solution:

Given cell reaction:

$\mathrm{Cd}_{(s)} + \mathrm{Cu}^{2+}_{(aq)} \longrightarrow \mathrm{Cd}^{2+}_{(aq)} + \mathrm{Cu}_{(s)}$

Step 1: Write the Nernst Equation

For a cell reaction:

$\mathrm{M}_1(s) + \mathrm{M}_2^{2+}(aq) \rightarrow \mathrm{M}_1^{2+}(aq) + \mathrm{M}_2(s)$

the Nernst equation is:

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0591}{n}\log\frac{[\mathrm{M}_1^{2+}]}{[\mathrm{M}_2^{2+}]}$$

For this reaction, $n = 2$ electrons are transferred.

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0591}{2}\log\frac{[\mathrm{Cd}^{2+}]}{[\mathrm{Cu}^{2+}]}$$

Step 2: Substitute Given Concentrations

Given:

- $[\mathrm{Cd}^{2+}] = 10 \times [\mathrm{Cu}^{2+}]$
- So, $\frac{[\mathrm{Cd}^{2+}]}{[\mathrm{Cu}^{2+}]} = 10$

Plug in the values:

$$E_{cell} = E^\circ_{cell} - \frac{0.0591}{2}\log(10)$$

$\log(10) = 1$

$$E_{cell} = E^\circ_{cell} - \frac{0.0591}{2} \times 1$$

$$E_{cell} = E^\circ_{cell} - 0.02955$$

Rounding to match options: 0.0296 V

It is less than E°_{cell} by 0.0296 V.

Correct answer:

Option D: less than E°_{cell} by 0.0296 V

Question34

Which of the following species acts as reducing agent during working of hydrogen-oxygen fuel cell?

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Options:

A. H_2

B. O_2

C. H^+

D. NaOH

Answer: A

Solution:

The correct answer is **Option A:** H_2.

Step-by-step explanation:

1. **In a hydrogen-oxygen fuel cell:**

- Hydrogen gas (H_2) is supplied at the anode.
- Oxygen gas (O_2) is supplied at the cathode.

1. **Reactions at electrodes:**

- At the **anode** (oxidation):

  $H_2(g) \rightarrow 2H^+(aq) + 2e^-$

  Hydrogen gas loses electrons. This means hydrogen is oxidised.

- At the **cathode** (reduction):

  $\frac{1}{2}O_2(g) + 2H^+(aq) + 2e^- \rightarrow H_2O(l)$

  Oxygen gains electrons (it is reduced).

1. **Reducing agent:**

- The substance that *loses electrons* and causes reduction is called the reducing agent.
- Here, H_2 **loses electrons**.
- Therefore, H_2 acts as the **reducing agent**.

Final Answer:

$\boxed{H_2}$

acts as the reducing agent during the working of a hydrogen-oxygen fuel cell.

Question35

Standard potential (E°) of $Zn^{+2}_{(aq)} + 2e^- \longrightarrow Zn_{(s)}$ is -0.76 V .

What is standard potential of reaction $2Zn_{(s)} \longrightarrow 2Zn^{+2}_{(aq)} + 4e^-$?

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Options:

A. -1.52 V

B. +1.52 V

C. -0.76 V

D. +0.76 V

Answer: D

Solution:

Given:

Standard reduction potential,

$\mathrm{Zn}^{2+}(aq) + 2e^- \rightarrow \mathrm{Zn}(s), \quad E^\circ = -0.76\,\mathrm{V}$

We need to find the standard potential for:

$2\mathrm{Zn}(s) \rightarrow 2\mathrm{Zn}^{2+}(aq) + 4e^-$

Step 1: Write the reverse reaction and adjust coefficients

Original reduction reaction:

$\mathrm{Zn}^{2+}(aq) + 2e^- \rightarrow \mathrm{Zn}(s), \quad E^\circ = -0.76\,\mathrm{V}$

Reverse (oxidation) reaction:

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

For the reverse reaction, the standard potential changes sign:

$E^\circ_{\text{oxidation}} = -(-0.76\,\mathrm{V}) = +0.76\,\mathrm{V}$

Step 2: Multiply by 2 (stoichiometric coefficients)

$2\mathrm{Zn}(s) \rightarrow 2\mathrm{Zn}^{2+}(aq) + 4e^-$

Changing the coefficients does **not** change the value of E° (it depends only on the electrode, not on the number of electrons).

So,

$E^\circ = +0.76\,\mathrm{V}$

Final Answer:

Option D: $+0.76\,\mathrm{V}$

Question36

Which from following formulae is used to obtain value of $\mathrm{E}^{r}_{\text{cell}}$ for a reaction taking place in Dry cell?

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Options:

A. $\frac{-\Delta \mathrm{G}^{\circ}}{\mathrm{F}}$

B. $\frac{-\Delta G^{\circ}}{2F}$

C. $\frac{-\Delta \mathrm{G}^{\circ}}{3 \mathrm{~F}}$

D. $\frac{-\Delta G^{\circ}}{4F}$

Answer: B

Solution:

The standard cell potential, E°_{cell} , for a reaction is related to the standard Gibbs free energy change, ΔG°, by the formula:

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

where:

- n = number of electrons involved in the reaction,
- F = Faraday constant.

Rearranging,

$$E^{\circ}_{\text{cell}} = \frac{-\Delta G^{\circ}}{nF}$$

For a dry cell, the net reaction involves **2 electrons** ($n = 2$).

Thus, the correct formula is:

$$E^{\circ}_{\text{cell}} = \frac{-\Delta G^{\circ}}{2F}$$

Correct answer: Option B

Question37

For a Galvanic cell consisting zinc electrode and standard hydrogen electrode,

$$\mathrm{E}^{\mathrm{o}}\left(\mathrm{Zn}^{+2}_{(\mathrm{aq})} \mid \mathrm{Zn}_{(\mathrm{s})}\right) = -0.76 \mathrm{~V}$$

Identify the reaction that takes place at positive electrode during working of cell?

Options:

A. $Zn_{(s)} \longrightarrow Zn^{+2}_{(aq)} + 2e^-$

B. $Zn^{+2}_{(aq)} + 2e^- \longrightarrow Zn_{(s)}$

C.

$H_{2(g)} \longrightarrow 2H^+_{(g)} + 2e^-$

D. $2H^+_{(g)} + 2e^- \longrightarrow H_{2(g)}$

Answer: D

Solution:

As $E_{red\left(Zn^{+2\circ}_{(aq)|Zn_{(s)}}\right)} < E_{red\left(H^{+\circ}_{(aq)|H_{2(g)}}\right)}$,

Zn acts as anode and H_2 as cathode. Cell representation:

$Zn_{(s)} \left|Zn^{2+}_{(aq)}\right\| H^+_{(aq,1M)} \left|H_{2(g,1\ atm)}\right| Pt$

Cell reactions:

At anode (negative electrode):

$Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)} + 2e^-$

At cathode (positive electrode):

$2H^+_{(aq)} + 2e^- \longrightarrow H_{2(g)}$

Question38

What is the number of faraday required to form $1\ molH_2$ by reduction of H^+ions?

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Options:

A. 4

B. 2

C. 0.5

D. 1

Answer: B

Solution:

Step 1: Write the reduction half-reaction

$2H^{+} + 2e^{-} \rightarrow H_2$

Step 2: Observe carefully

To form **1 mole of** H_2, the equation shows that **2 moles of electrons** are required.

Step 3: Recall definition of 1 Faraday

- 1 Faraday = charge of 1 mole of electrons.

Step 4: Calculate

- For 2 moles of electrons → 2 Faradays are required.

Faradays required $= 2F$

Correct Option: B (2) ✅

Question39

For the cell,

$$^{\ominus}\mathrm{Zn}_{(s)}\left|\mathrm{Zn}^{+2}(1\mathrm{M})\right|\left|\mathrm{Ag}^{+1}(1\mathrm{M})\right|\mathrm{Ag}_{(s)}{}^{\oplus}$$

If concentration of Zn^{+2} decreases to 0.1 M at 298 K , then emf of cell

Options:

A. increase by 0.0592 V

B. decrease by 0.0592 V

C. increase by 0.0296 V

D. decrease by 0.0296 V

Answer: C

Solution:

$Zn_{(s)} + 2Ag^{+}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + 2Ag_{(s)}$

For the given cell reaction, $n = 2$.

Using Nernst equation,

$$\begin{aligned} E_{cell} &= E^{\circ}_{cell} - \frac{0.0592}{n}\log_{10} Q \\ E'_{cell} &= E^{\circ}_{cell} - \frac{0.0592}{2}\log_{10}\frac{[Zn^{2+}]}{[Ag^{+}]^{2}} \end{aligned}$$

When Zn^{2+} and $Ag^{+} = 1M$,

$$\begin{aligned} E_{cell} &= E^{\circ}_{cell} - 0.0296\log_{10}\frac{1}{1} \\ \therefore \quad E_{cell} &= E^{\circ}_{cell} \qquad \ldots(i) \end{aligned}$$

When $Zn^{2+} = 0.1M$ and $Ag^{+} = 1M$,

$$\begin{aligned} E_{cell} &= E^{\circ}_{cell} - 0.0296\log_{10}\frac{0.1}{1} \\ \therefore E_{cell} &= E^{\circ}_{cell} - 0.0296(-1) \\ \therefore E_{cell} &= E^{\circ}_{cell} + 0.0296\ \mathrm{V} \ldots\ (ii) \end{aligned}$$

Thus, from (i) and (ii), the emf of the cell increases by 0.0296 V .

Question40

For the cell reaction,

$$\mathrm{Zn}_{(s)}+2\mathrm{Ag}^{+}_{(aq)} \longrightarrow \mathrm{Zn}^{+2}_{(aq)}+2\mathrm{Ag}_{(s)}$$

Cell potential is less than $\mathrm{E}^{\circ}_{\text{cell}}$ by 0.0592 V at 298 K when

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Options:

A. $\left[\mathrm{Zn}^{+2}\right]=1\mathrm{M}$ and $\left[\mathrm{Ag}^{+}\right]=0.1\mathrm{M}$

B. $\left[\mathrm{Zn}^{+2}\right]=1\mathrm{M}$ and $\left[\mathrm{Ag}^{+}\right]=0.01\mathrm{M}$

C. $\left[\mathrm{Zn}^{+2}\right]=0.1\mathrm{M}$ and $\left[\mathrm{Ag}^{+}\right]=1\mathrm{M}$

D. $\left[\mathrm{Zn}^{+2}\right]=0.01\mathrm{M}$ and $\left[\mathrm{Ag}^{+}\right]=1\mathrm{M}$

Answer: A

Solution:

Given reaction:

$\mathrm{Zn}_{(s)}+2\mathrm{Ag}^{+}_{(aq)} \longrightarrow \mathrm{Zn}^{2+}_{(aq)}+2\mathrm{Ag}_{(s)}$

According to the **Nernst Equation** at 298 K,

$E_{\text{cell}}=E^{\circ}_{\text{cell}}-\frac{0.0592}{n}\log Q$

Here,

- $n=2$ (number of electrons transferred)
- $Q=$ reaction quotient

Write the reaction quotient, Q:

$\text{Products}^{\text{coeff}}/\text{Reactants}^{\text{coeff}}$ (excluding solids):

$Q=\frac{[\mathrm{Zn}^{2+}]}{[\mathrm{Ag}^{+}]^{2}}$

We are told:

$E_{\text{cell}}=E^{\circ}_{\text{cell}}-0.0592\ V$

So,

$E_{\text{cell}}-E^{\circ}_{\text{cell}}=-0.0592\ V$

Let us compare with the Nernst equation:

$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0592}{2} \log Q$

Therefore,

$-0.0592 = -\frac{0.0592}{2} \log Q$

Divide both sides by -0.0592:

$1 = \frac{1}{2} \log Q \log Q = 2Q = 10^2 = 100$

We need:

$\frac{[\mathrm{Zn}^{2+}]}{[\mathrm{Ag}^+]^2} = 100$

Now, let's check each option and see which values give $Q = 100$.

Option A

$\left[\mathrm{Zn}^{2+}\right] = 1\ \mathrm{M}, \left[\mathrm{Ag}^{+}\right] = 0.1\ \mathrm{M}$

$Q = \frac{1}{(0.1)^2} = \frac{1}{0.01} = 100$

Option B

$\left[\mathrm{Zn}^{2+}\right] = 1\ \mathrm{M}, \left[\mathrm{Ag}^{+}\right] = 0.01\ \mathrm{M}$

$Q = \frac{1}{(0.01)^2} = \frac{1}{0.0001} = 10000$

Option C

$\left[\mathrm{Zn}^{2+}\right] = 0.1\ \mathrm{M}, \left[\mathrm{Ag}^{+}\right] = 1\ \mathrm{M}$

$Q = \frac{0.1}{(1)^2} = 0.1$

Option D

$\left[\mathrm{Zn}^{2+}\right] = 0.01\ \mathrm{M}, \left[\mathrm{Ag}^{+}\right] = 1\ \mathrm{M}$

$Q = \frac{0.01}{(1)^2} = 0.01$

Correct answer:

$\boxed{\text{Option A}}$

Question41

Which from following is the correct relationship between molar conductivity (Λ), conductivity (k) and molarity (M) of solution for electrolyte?

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Options:

A. $\mathrm{k}=\frac{\wedge \times \mathrm{C}}{1000}$

B. $\wedge=\frac{100 \times \mathrm{k}}{\mathrm{C}}$

C. $\wedge=\frac{\mathrm{k} \times \mathrm{C}}{1000}$

D. $\mathrm{k}=\frac{1000 \times \mathrm{C}}{\wedge}$

Answer: A

Solution:

The correct relationship between molar conductivity (Λ), conductivity (k), and concentration (molarity, C) is:

$\Lambda=\frac{k \times 1000}{C}$

where,

Λ = molar conductivity (S cm^2 mol^{-1}),

k = conductivity (S cm^{-1}),

C = concentration in mol L^{-1}.

Let's analyze the options:

- **Option A:** $k=\frac{\Lambda \times C}{1000}$ (incorrect manipulation)
- **Option B:** $\Lambda=\frac{100 \times k}{C}$ (incorrect factor)
- **Option C:** $\Lambda=\frac{k \times C}{1000}$ (incorrect formula)
- **Option D:** $k=\frac{1000 \times C}{\Lambda}$ (incorrect manipulation)

None of these match the correct formula. However, if we rearrange the correct relationship:

$\Lambda=\frac{k \times 1000}{C}$

OR

$k=\frac{\Lambda \times C}{1000}$

So, Option A is correct.

Final Answer:

Option A $\boxed{k=\frac{\Lambda \times C}{1000}}$ is correct.

Question42

How long should aqueous NaCl be electrolysed by passing 100 ampere current, so that 0.5 mol chlorine is released at anode?

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Options:

A. 96500 seconds

B. 9650 seconds

C. 965 seconds

D. 96.5 seconds

Answer: C

Solution:

The required reaction is:

$2\mathrm{Cl}^{-}_{(\mathrm{aq})}+2\mathrm{e}^{-} \longrightarrow \mathrm{Cl}_{2(\mathrm{~g})}$

Mole ratio $=\frac{1 \text{ mol}}{2 \text{ mole}^{-}}$

Moles of product formed

$$
\begin{aligned}
& =\frac{\mathrm{I} \times \mathrm{t}}{96500\left(\mathrm{C} / \mathrm{mole}^{-}\right)} \times \text{mole ratio} \\
& 0.5 \text{ mol}=\frac{100 \times \mathrm{t}}{96500 \mathrm{C} / \mathrm{mole}^{-}} \times \frac{1 \text{ mol}}{2 \text{ mole}^{-}} \\
\therefore \quad & \mathrm{t}=\frac{0.5 \times 96500 \times 2}{100 \times 1}=965 \text{ seconds}
\end{aligned}
$$

Question43

Calculate the cell constant of conductivity cell containing 0.1 M KCl solution having resistance 60Ω and conductivity $0.014 \Omega^{-1} \mathrm{~cm}^{-1}$ at $25^{\circ} \mathrm{C}$.

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Options:

A. $0.42 \mathrm{~cm}^{-1}$

B. $0.84 \mathrm{~cm}^{-1}$

C. $0.60 \mathrm{~cm}^{-1}$

D. $1.04 \mathrm{~cm}^{-1}$

Answer: B

Solution:

The cell constant of a conductivity cell can be calculated using the relationship between conductivity, resistance, and the cell constant.

The formula to calculate the cell constant (K) is:

$K=\kappa \cdot R$

where:

K is the cell constant.

κ is the conductivity of the solution ($0.014 \Omega^{-1} \mathrm{~cm}^{-1}$).

R is the resistance of the solution (60Ω).

Substituting the given values:

$K=0.014 \Omega^{-1} \mathrm{~cm}^{-1} \times 60 \Omega$

This simplifies to:

$K=0.84 \mathrm{~cm}^{-1}$

Therefore, the cell constant of the conductivity cell is $\boxed{0.84 \mathrm{~cm}^{-1}}$.

Question44

What is the number of faraday required to produce 0.18 g aluminium at cathode during electrolysis of molten AlCl_3 ?(Molar mass of $\mathrm{Al}=27 \mathrm{~g} \mathrm{~mol}^{-1}$)

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Options:

A. 0.02 F

B. 0.03 F

C. 0.25 F

D. 0.30 F

Answer: A

Solution:

To calculate the number of Faradays required to produce 0.18 g of aluminum at the cathode during the electrolysis of molten $AlCl_3$, we can follow these steps:

Determine the moles of aluminum:

Given the mass of aluminum produced is 0.18 g, and its molar mass is 27 g/mol, the moles of aluminum is calculated as:

$$\text{Moles of Al} = \frac{\text{mass}}{\text{molar mass}} = \frac{0.18\text{ g}}{27\text{ g/mol}} = 0.00667\,\text{mol}$$

Understand the electrochemical reaction:

The electrochemical reaction at the cathode for Al^{3+} ions is given by:

$$Al^{3+} + 3e^{-} \rightarrow Al$$

This shows that 3 moles of electrons are required to deposit 1 mole of aluminum.

Calculate the number of Faradays:

Using the relationship between moles of electrons, moles of aluminum, and Faraday's constant (1 Faraday = 1 mole of electrons), the number of Faradays required is given by:

$$\text{Faradays required} = \text{Moles of Al} \times 3 = 0.00667\,\text{mol} \times 3 = 0.02001\,\text{F}$$

Therefore, 0.02 Faradays are required.

The correct option is **A: 0.02 F**.

Question45

Which from following equations represents a correct relationship between standard cell potential and equilibrium constant for cell reaction?

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Options:

A. $E^\circ_{cell} = -\frac{2.303RT}{nF}$

B. $E^\circ_{cell} = \frac{0.0592}{nF}\log_{10} K$

C. $E^\circ_{cell} = \frac{0.0592}{n}\log_{10} K$

D. $E^\circ_{cell} = \frac{0.0592}{n}\ln K$

Answer: C

Solution:

We know from electrochemistry that the relationship between the standard cell potential E°_{cell} and the equilibrium constant K for the cell reaction (at 25°C) is:

$$E^\circ_{cell} = \frac{0.0592}{n}\log_{10}(K).$$

This comes from combining the thermodynamic and electrochemical relationships:

$$\Delta G^\circ = -RT\ln K \quad \text{and} \quad \Delta G^\circ = -nF\,E^\circ_{cell},$$

and at 25°C, using $2.303\,\frac{RT}{F} \approx 0.0592\,V$.

Checking the options:

Option A: $E^\circ_{cell} = -\frac{2.303RT}{nF}$ — This lacks $\log K$, so it is incorrect.

Option B: $E^\circ_{cell} = \frac{0.0592}{nF}\log_{10} K$ — There's an extra F in the denominator that does not belong.

Option C: $E^\circ_{cell} = \frac{0.0592}{n}\log_{10} K$ — **This matches the well-known formula**.

Option D: $E^\circ_{cell} = \frac{0.0592}{n}\ln K$ — The 0.0592 factor is valid for a base-10 logarithm, not for the natural log (ln).

Therefore, the **correct relationship** is:

$$\boxed{E^\circ_{cell} = \frac{0.0592}{n}\log_{10} K \quad \text{(Option C)}}$$

Question46

What is the conductivity of 0.05 M NaOH solution having resistance 31.5 ohm and cell constant $0.315\ \mathrm{cm}^{-1}$?

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Options:

A. $100\ \mathrm{ohm}^{-1}\ \mathrm{cm}^{-1}$

B. $0.02\ \mathrm{ohm}^{-1}\ \mathrm{cm}^{-1}$

C. $0.09\ \mathrm{ohm}^{-1}\ \mathrm{cm}^{-1}$

D. $0.01\ \mathrm{ohm}^{-1}\ \mathrm{cm}^{-1}$

Answer: D

Solution:

The conductivity (κ) of a solution can be calculated using the formula:

$$\kappa = \frac{\text{Cell Constant}}{\text{Resistance}}$$

Given:

Resistance $R = 31.5\,\Omega$

Cell Constant $= 0.315\,\mathrm{cm}^{-1}$

Substituting the given values into the formula:

$$\kappa = \frac{0.315\ \mathrm{cm}^{-1}}{31.5\ \Omega}$$

Simplifying the above expression:

$$\kappa = \frac{0.315}{31.5}\ \mathrm{ohm}^{-1}\,\mathrm{cm}^{-1}$$

$$\kappa = 0.01\,\mathrm{ohm}^{-1}\,\mathrm{cm}^{-1}$$

Thus, the conductivity of the 0.05 M NaOH solution is $0.01\ \mathrm{ohm}^{-1}\ \mathrm{cm}^{-1}$, which corresponds to Option D.

Question47

Which from the following concentrations of a weak electrolyte solution exhibits maximum molar conductivity?

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Options:

A. 0.004 M

B. 0.002 M

C. 0.005 M

D. 0.001 M

Answer: D

Solution:

The molar conductivity of weak electrolytes increases rapidly on decreasing concentration.

$\therefore$ 0.001 M weak electrolyte solution exhibits maximum molar conductivity.

Question48

Consider the following cellE° cell $= 1.007$ V and E° calomel $= 0.242$ V What is the standard potential of Zn ?

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Options:

A. -0.765 V

B. 0.765 V

C. -1.247 V

D. 1.247 V

Answer: A

Solution:

The standard electrode potential for the cell is given by the equation:

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

Here, the given values are:

$E^{\circ}_{cell} = 1.007\ \text{V}$

$E^{\circ}_{calomel} = 0.242\ \text{V}$ (this is the calomel electrode, typically used as the reference electrode)

Assuming the calomel electrode acts as the cathode, we can plug these values into the equation:

$1.007\ \text{V} = 0.242\ \text{V} - E^{\circ}_{Zn}$

To find the standard potential of Zn, rearrange the equation for E°_{Zn} :

$E^{\circ}_{Zn} = 0.242\ \text{V} - 1.007\ \text{V}$

$E^{\circ}_{Zn} = -0.765\ \text{V}$

Therefore, the standard potential of Zn is:

Option A: $-0.765\ \text{V}$

Question49

Calculate the percentage dissociation of 0.05 M solution of weak electrolyte if its molar conductivity and molar conductivity at infinite dilution are respectively. $3.3\Omega^{-1}\ \text{cm}^2\ \text{mol}^{-1}$ and $132\Omega^{-1}\ \text{cm}^2\ \text{mol}^{-1}$.

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Options:

A. 4.0%

B. 3.5%

C. 2.5%

D. 10.0%

Answer: C

Solution:

To calculate the percentage dissociation of a weak electrolyte, use the formula for the degree of dissociation (α):

$\alpha = \frac{\Lambda_c}{\Lambda_0}$

where:

Λ_c is the molar conductivity at concentration c.

Λ_0 is the molar conductivity at infinite dilution (also known as the limiting molar conductivity).

Given the values:

$\Lambda_c = 3.3\,\Omega^{-1}\,\mathrm{cm}^2\,\mathrm{mol}^{-1}$

$\Lambda_0 = 132\,\Omega^{-1}\,\mathrm{cm}^2\,\mathrm{mol}^{-1}$

Substitute these into the formula:

$\alpha = \frac{3.3}{132}$

Calculate α:

$\alpha = \frac{3.3}{132} = \frac{1}{40} = 0.025$

The degree of dissociation α is 0.025, which means the percentage dissociation is:

Percentage dissociation = $\alpha \times 100\% = 0.025 \times 100\% = 2.5\%$

Thus, the percentage dissociation of the electrolyte is **2.5%**.

Therefore, the correct answer is Option C: 2.5%.

Question50

The conductivity of 0.02 M solution of $\mathrm{AgNO_3}$ is $0.00216\Omega^{-1}\,\mathrm{cm}^{-1}$ at 298 K . What is its molar conductivity?

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Options:

A. $232.0 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

B. $402.0 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

C. $108.0 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

D. $150.0 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

Answer: C

Solution:

$$
\begin{aligned}
\Lambda & =\frac{1000 \mathrm{k}}{\mathrm{c}}=\frac{1000 \mathrm{~cm}^3 \mathrm{~L}^{-1} \times 0.00216 \Omega^{-1} \mathrm{~cm}^{-1}}{0.02 \mathrm{~mol} \mathrm{~L}^{-1}} \\
& =108.0 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}
\end{aligned}
$$

Question51

What is the decreasing order of deposition of metal on electrode if standard reduction potentials are given as -

$$
\mathrm{Ag}^{+}\left|\mathrm{Ag}=0.80 \mathrm{~V}, \mathrm{Cu}^{2+}\right| \mathrm{Cu}=0.337 \mathrm{~V}
$$

$$
\mathrm{Sn}^{2+}\left|\mathrm{Sn}=-0.136 \mathrm{~V}, \mathrm{Cd}^{2+}\right| \mathrm{Cd}=-0.403 \mathrm{~V}
$$

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Options:

A. $\mathrm{Ag}>\mathrm{Cu}>\mathrm{Sn}>\mathrm{Cd}$

B. $\mathrm{Cu}>\mathrm{Sn}>\mathrm{Cd}>\mathrm{Ag}$

C. $\mathrm{Sn}>\mathrm{Cd}>\mathrm{Ag}>\mathrm{Cu}$

D. $\mathrm{Cd}>\mathrm{Sn}>\mathrm{Cu}>\mathrm{Ag}$

Answer: A

Solution:

Higher the standard reduction potential (E°) value, greater the tendency of the species to accept electrons and undergo reduction.

Order of E_{red}^{o} : $Ag^{+} > Cu^{2+} > Sn^{2+} > Cd^{2+}$

$\therefore$ Decreasing order of deposition of metal on electrode is as follows: $Ag > Cu > Sn > Cd$.

Question52

The E_{cell}^{+} of $Cu_{(s)} \left| Cu_{(1M)}^{++} \right\| Ag_{(1M)}^{+} \left| Ag_{(s)} \right.$ is 0.647 volt. Calculate the E_{Ag}° if E_{Cu}° is 0.153 V .

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Options:

A. 0.8 V

B. 0.5 V

C. -0.8 V

D. -0.5 V

Answer: A

Solution:

The given cell is composed of copper electrode as the anode and silver electrode as the cathode. The standard cell potential is given by

$$
\begin{aligned}
& E_{cell}^{\circ} = E_{cathode}^{\circ} - E_{anode}^{\circ} \\
& \therefore \quad E_{cell}^{\circ} = E_{Ag^{+}|Ag}^{\circ} - E_{Cu^{2+}|Cu}^{o} \\
& \therefore \quad E_{Ag^{+}|Ag}^{\circ} = E_{cell}^{\circ} + E_{Cu^{2+}|Cu}^{\circ} \\
& = (0.647 \text{ V}) + (0.153 \text{ V}) = 0.8 \text{ V}
\end{aligned}
$$

Question53

Calculate the mass of ' Ca ' deposited at cathode by passing 0.8 ampere current through molten $\mathrm{CaCl_2}$ in 60 minutes.[Molar mass of $\mathrm{Ca} = 40 \mathrm{~g} \mathrm{~mol}^{-1}$]

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Options:

A. 0.4 g

B. 0.5 g

C. 0.6 g

D. 0.7 g

Answer: C

Solution:

To calculate the mass of calcium ('Ca') deposited at the cathode, Faraday's laws of electrolysis can be used. Here's a step-by-step calculation:

Step 1: Calculate the total charge (Q) passed through the electrolyte.

The formula to find the total charge is:

$Q = I \times t$

where:

$I = 0.8\,\mathrm{A}$ (current)

$t = 60 \times 60 = 3600\,\mathrm{s}$ (time converted to seconds)

So,

$Q = 0.8\,\mathrm{A} \times 3600\,\mathrm{s} = 2880\,\mathrm{Coulombs}$

Step 2: Use Faraday's Second Law of Electrolysis.

The mass of substance deposited/liberated is given by:

$m = \frac{Q \times M}{n \times F}$

where:

m = mass of substance deposited (in grams)

$Q = 2880\,\mathrm{C}$ (total charge)

$M = 40\,\mathrm{g/mol}$ (molar mass of calcium)

$n = 2$ (number of electrons exchanged; calcium forms Ca^{2+})

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

Step 3: Calculate the mass of calcium deposited.

Substituting in the values:

$m = \frac{2880 \times 40}{2 \times 96500}$

Simplifying:

$m = \frac{115200}{193000}$

$m \approx 0.5969\,\mathrm{g}$

Rounding to one decimal place gives:

$m \approx 0.6\,\mathrm{g}$

Therefore, the mass of calcium deposited at the cathode is approximately **0.6 grams**.

Option C: 0.6 g

Question54

What is the concentration of an electrolyte solution to have molar conductivity of $101 \Omega^{-1}\ \mathrm{cm}^2\ \mathrm{mol}^{-1}$ and conductivity of $1.01 \times 10^{-2} \Omega^{-1}\ \mathrm{cm}^{-1}$ at 298 K?

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Options:

A. 0.05 M

B. 0.1 M

C. 0.15 M

D. 0.2 M

Answer: B

Solution:

$$\text{Molar conductivity } (\wedge) = \frac{1000\mathrm{k}}{\mathrm{c}}$$

$$101\Omega^{-1}\,\mathrm{cm}^2\,\mathrm{mol}^{-1} = \frac{1000 \times 1.01\Omega^{-1}\,\mathrm{cm}^{-1} \times 10^{-2}}{\mathrm{c}}$$

$$\therefore \quad \mathrm{c} = \frac{1000 \times 1.01 \times 10^{-2}}{101} = 0.1\mathrm{M}$$

Question55

Cell constant of a conductivity cell is $0.9\ \mathrm{cm}^{-1}$ and resistance shown by $\mathrm{AgNO_3}$ solution is 6530 ohm. What is the conductivity of $\mathrm{AgNO_3}$ solution?

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Options:

A. $725\Omega^{-1}\,\mathrm{cm}^{-1}$

B. $5870\Omega^{-1}\,\mathrm{cm}^{-1}$

C. $1.38 \times 10^{-4}\Omega^{-1}\,\mathrm{cm}^{-1}$

D. $4.72 \times 10^{-4}\Omega^{-1}\,\mathrm{cm}^{-1}$

Answer: C

Solution:

The conductivity (κ) of a solution can be calculated using the formula:

$\kappa = \frac{K}{R}$

where:

K is the cell constant, given in this case as $0.9\,\mathrm{cm}^{-1}$,

R is the resistance of the solution, given as $6530\,\Omega$.

Substituting the given values into the formula:

$\kappa = \frac{0.9 \text{ cm}^{-1}}{6530 \, \Omega}$

Perform the calculation:

Calculate the division:

$\kappa = \frac{0.9}{6530}$

Simplify the fraction:

$\kappa \approx 1.38 \times 10^{-4} \, \Omega^{-1} \text{ cm}^{-1}$

Thus, the conductivity of the $AgNO_3$ solution is approximately $1.38 \times 10^{-4} \, \Omega^{-1} \text{ cm}^{-1}$. Therefore, the correct option is:

Option C

$1.38 \times 10^{-4} \, \Omega^{-1} \text{ cm}^{-1}$

Question56

Which of the following cannot be used as standard solution for determination of cell constant of conductivity cell?

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Options:

A. 1 M KCl

B. 0.1 M KCl

C. 0.01 M KCl

D. Saturated KCl

Answer: D

Solution:

What is a Standard Solution for Conductivity?

To determine the cell constant of a conductivity cell, one typically uses a standard solution whose conductivity is well-known and stable. Potassium chloride (KCl) solutions at specific molarities (e.g., 0.1 M

KCl, 0.01 M KCl, and even 1 M KCl) have well-established conductivities at certain temperatures and are commonly used for calibration.

Suitability of Different KCl Concentrations :

1 M KCl : Conductivity at this concentration is well-tabulated and can be used for calibration.

0.1 M KCl : This is a very common standard for calibrating conductivity cells. Its conductivity is thoroughly documented.

0.01 M KCl : Another known standard solution with a well-established conductivity value, suitable for more sensitive or low-conductivity measurements.

Saturated KCl Solution :

A saturated solution of KCl is not typically used as a standard solution for conductivity calibration. The reason is that a saturated KCl solution may have variable composition due to temperature changes, partial crystallization, and difficulty in ensuring a uniform, reproducible concentration. Thus, it does not have a precisely known or easily reproducible conductivity value as the other standard concentrations do.

Conclusion :

Out of the given options, the saturated KCl solution (Option D) is not suitable for use as a standard solution in determining the cell constant because its concentration and thus its conductivity are not stable or easily reproducible.

Correct Answer :

$\boxed{\text{Saturated KCl (Option D)}}$

Question57

When fused NaCl undergoes electrolysis the product formed at cathode is

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Options:

A. $Cl_{2(g)}$

B. $Na_{(s)}$

C. $Na-Hg_{(l)}$

D. $H_{2(g)}$

Answer: B

Solution:

Reaction at cathode: $2Na^+_{(l)} + 2e^- \longrightarrow 2Na_{(l)}$

During the electrolysis of fused NaCl , molten silvery-white sodium is formed at cathode.

Question58

Which of the following element is used in photoelectric cell?

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Options:

A. Beryllium

B. Caesium

C. Liquid sodium

D. Calcium

Answer: B

Solution:

Among the listed options, **caesium** is the element commonly used in photoelectric cells due to its low ionization energy and hence its ability to emit electrons upon exposure to light.

Answer: (B) Caesium.

Question59

The resistance of conductivity cell filled with 0.1 M KCl solution is 100 ohm and conductivity is 1.70×10^{-4} S cm^{-1}. What is the cell constant of the cell?

Options:

A. $1.70 \mathrm{~cm}^{-1}$

B. $0.058 \mathrm{~cm}^{-1}$

C. $0.017 \mathrm{~cm}^{-1}$

D. $0.012 \mathrm{~cm}^{-1}$

Answer: C

Solution:

The cell constant of a conductivity cell is calculated using the formula:

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

Given:

Conductivity, $\kappa = 1.70 \times 10^{-4} \, \mathrm{S\,cm^{-1}}$

Resistance, $R = 100 \, \mathrm{ohm}$

Substitute these values into the formula:

$\text{Cell Constant} = (1.70 \times 10^{-4} \, \mathrm{S\,cm^{-1}}) \times 100 \, \Omega$

Calculate the cell constant:

$\text{Cell Constant} = 1.70 \times 10^{-2} \, \mathrm{cm^{-1}}$

Therefore, the cell constant of the conductivity cell is:

Option C: $0.017 \, \mathrm{cm^{-1}}$

Question60

The conductivity of 0.02 M KCl solution is 0.00250 ohm $^{-1} \mathrm{~cm}^{-1}$. What is its molar conductivity?

Options:

A. $125 \mathrm{~ohm}^{-1} \mathrm{~cm}^{-2} \mathrm{~mol}^{-1}$

B. $0.05 \mathrm{~ohm}^{-1} \mathrm{~cm}^{-2} \mathrm{~mol}^{-1}$

C. $725 \mathrm{~ohm}^{-1} \mathrm{~cm}^{-2} \mathrm{~mol}^{-1}$

D. $800 \mathrm{~ohm}^{-1} \mathrm{~cm}^{-2} \mathrm{~mol}^{-1}$

Answer: A

Solution:

Molar conductivity (Λ_m) is calculated using the formula:

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

where κ is the conductivity of the solution, and c is the concentration of the solution in moles per liter (M).

Given:

Conductivity, $\kappa = 0.00250 \, \mathrm{ohm}^{-1} \, \mathrm{cm}^{-1}$

Concentration, $c = 0.02 \, \mathrm{M}$

Substitute these values into the formula:

$\Lambda_m = \frac{0.00250 \, \mathrm{ohm}^{-1} \, \mathrm{cm}^{-1}}{0.02 \, \mathrm{M}}$

Perform the division:

$\Lambda_m = \frac{0.00250}{0.02} \, \mathrm{ohm}^{-1} \, \mathrm{cm}^2 \, \mathrm{mol}^{-1}$

$\Lambda_m = 0.125 \, \mathrm{ohm}^{-1} \, \mathrm{cm}^2 \, \mathrm{mol}^{-1}$

Therefore, the molar conductivity is $125 \, \mathrm{ohm}^{-1} \, \mathrm{cm}^2 \, \mathrm{mol}^{-1}$.

The correct answer is Option A:

$125 \, \mathrm{ohm}^{-1} \, \mathrm{cm}^2 \, \mathrm{mol}^{-1}$

Question61

The molar conductivity of 0.01 M acetic acid at $25^{\circ} \mathrm{C}$ is $16.5 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$ and its molar conductivity at zero concentration

is $390.7 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$. What is its degree of dissociation?

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Options:

A. 0.0223

B. 0.0422

C. 0.0642

D. 0.0821

Answer: B

Solution:

The degree of dissociation (α) of acetic acid can be calculated using the formula for molar conductivity :

$\alpha=\frac{\Lambda_m}{\Lambda_m^0}$

where :

Λ_m is the molar conductivity at the given concentration.

Λ_m^0 is the molar conductivity at zero concentration (infinite dilution).

Given :

$\Lambda_m=16.5 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

$\Lambda_m^0=390.7 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

Substitute the values into the formula :

$\alpha=\frac{16.5}{390.7}$

Calculating the above expression gives :

$\alpha \approx 0.0422$

Therefore, the degree of dissociation of acetic acid is approximately 0.0422.

Option B : 0.0422 is the correct answer.

Question62

Which of the following is released at cathode during electrolysis of aqueous sodium chloride?

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Options:

A. $Cl_{2(g)}$

B. $Na_{(s)}$

C. $O_{2(g)}$

D. $H_{2(g)}$

Answer: D

Solution:

During the electrolysis of aqueous sodium chloride (brine solution), the products formed at the electrodes depend on the ions present and their respective reduction and oxidation potentials.

In an aqueous solution of sodium chloride, the ions present are Na^+, Cl^-, H_2O, H^+, and OH^-. At the cathode, the reduction process takes place to form a neutral element or molecule. The two important reactions considered for reduction at the cathode are:

The reduction of sodium ions:

$Na^+ + e^- \rightarrow Na_{(s)}$

The reduction of water (leading to hydrogen gas):

$2H_2O + 2e^- \rightarrow H_{2(g)} + 2OH^-$

Between these two reactions, water is more readily reduced than sodium ions due to its lower reduction potential. As a consequence, hydrogen gas is released at the cathode.

Therefore, the correct answer is:

Option D

$H_{2(g)}$

Question63

Calculate the E°_{cell} of $Al\left|Al_2(SO_4)_{3(1M)}\right|\left|HCl_{(IM)}\right|H_{2(\ g)(1atm)}, Pt($
Given $E^{0}_{AB^{3+}|A|} = -1.66\ V$ **)**

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Options:

A. 1.66 V

B. −1.66 V

C. 0.0533 V

D. 1.14 V

Answer: A

Solution:

The standard cell potential, E°_{cell} , is calculated from the standard reduction potentials of the cathode and anode half-reactions. Given:

The standard reduction potential for the half-reaction involving aluminum:

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

The standard hydrogen electrode (SHE) is used as the reference with a potential of 0 V:

$H^+ + e^- \rightarrow \frac{1}{2}H_2 \quad E^\circ = 0\ V$

Since the cell notation provided is:

$Al|Al_2(SO_4)_{3(1M)}||HCl_{(1M)}|H_{2(g)(1atm)}, Pt$

The cell is an $Al|H_2$ cell, where aluminum acts as the anode and hydrogen as the cathode.

Half-reactions:

Anode (oxidation):

$Al \rightarrow Al^{3+} + 3e^- \quad (E^\circ = +1.66\ V \text{ since oxidation is the reverse process})$

Cathode (reduction):

$2H^+ + 2e^- \rightarrow H_2 \quad (E^\circ = 0\ V)$

Calculate E°_{cell} :

The cell potential is given by the difference in the reduction potentials of the cathode and the anode:

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

Substituting the values,

$E^{\circ}_{\text{cell}} = 0\text{ V} - (-1.66\text{ V}) = 1.66\text{ V}$

Thus, the standard cell potential E°_{cell} is:

Option A: 1.66 V

Question64

Which of the following units of electrical measurement is not equivalent to 1 Siemen?

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Options:

A. Ω^{-1}

B. AV^{-1}

C. $\text{CV}^{-1}\text{s}^{-1}$

D. Ω

Answer: D

Solution:

Option D, Ω, is not equivalent to 1 Siemen. The Siemen (S) is the unit of conductance and is equivalent to the reciprocal of resistance, which is measured in ohms (Ω). Therefore, Ω^{-1} is equivalent to 1 Siemen.

Here's why the other options represent Siemens:

Option A: Ω^{-1} is the definition of the Siemen (S), as it is the reciprocal of the ohm.

Option B: AV^{-1} can be simplified to Ω^{-1} when considering Ohm's Law, which states $V = IR$, so I/V indeed gives conductance ($1/R$).

Option C: $\text{CV}^{-1}\text{s}^{-1}$ can be related to Siemen by considering the units:

Since C is charge in coulombs, and V is voltage, the unit $\text{CV}^{-1}\text{s}^{-1}$ translates to amperes per volt, which is conductance.

Option D (Ω) represents resistance, not conductance, hence it does not equate to 1 Siemen.

Question65

A conductivity cell dipped in 0.05 MKCl has resistance 600 ohm. If conductivity is $0.0012 \mathrm{ohm}^{-1} \mathrm{~cm}^{-1}$. What is the value of cell constant?

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Options:

A. $0.50 \mathrm{~cm}^{-1}$

B. $0.72 \mathrm{~cm}^{-1}$

C. $1.5 \mathrm{~cm}^{-1}$

D. $2.0 \mathrm{~cm}^{-1}$

Answer: B

Solution:

$$\begin{aligned} & \text{Conductivity }(\mathrm{k})=\frac{\text{Cell constant}}{\mathrm{R}} \\ \therefore \quad & \text{Cell constant }=\mathrm{k} \times \mathrm{R} \\ & =0.0012 \mathrm{ohm}^{-1} \mathrm{~cm}^{-1} \times 600 \mathrm{ohm}=0.72 \mathrm{~cm}^{-1}\end{aligned}$$

Question66

What is the conductivity of 0.05 M KCl solution if cell constant is $1.32 \mathrm{~cm}^{-1}$ and resistance is 528 ohm?

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Options:

A. $0.0401 \text{ ohm}^{-1} \text{ cm}^{-1}$

B. $0.0051 \text{ ohm}^{-1} \text{ cm}^{-1}$

C. $0.0025 \text{ ohm}^{-1} \text{ cm}^{-1}$

D. $0.0691 \text{ ohm}^{-1} \text{ cm}^{-1}$

Answer: C

Solution:

$$\begin{aligned}\text{Conductivity (k)} &= \frac{\text{Cell constant}}{\text{R}} \\ &= \frac{1.32 \text{ cm}^{-1}}{528 \text{ ohm}} \\ &= 0.0025 \text{ ohm}^{-1} \text{ cm}^{-1}\end{aligned}$$

Question67

Which of the following substances conducts electricity?

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Options:

A. Diamond

B. Sulphur (solid)

C. Molten NaCl

D. Crystalline NaCl

Answer: C

Solution:

Molten NaCl (Option C) conducts electricity.

In its solid state, NaCl (sodium chloride) is composed of ions that are locked in place within the crystal lattice structure. However, when NaCl is melted, the rigid lattice is broken, allowing the free movement of its ions. These free ions are what enable the conduction of electricity. In contrast, diamond is an allotrope of carbon where each carbon atom is covalently bonded to four other carbon atoms, forming a non-conductive structure. Sulphur, in its solid state, consists of S_8 rings and does not have free-moving charges for conduction. Crystalline NaCl, similar to solid NaCl, has ions fixed in position, preventing electrical conductivity in its solid form.

Question68

Resistance and conductivity of a cell containing 0.1 M KCl solution at 298 K are 115 ohm and $1.90 \times 10^{-6}\ \mathrm{S\ cm}^{-1}$ respectively. What is the value of cell constant?

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Options:

A. $0.165\ \mathrm{cm}^{-1}$

B. $1.601\ \mathrm{cm}^{-1}$

C. $2.185\ \mathrm{cm}^{-1}$

D. $0.218\ \mathrm{cm}^{-1}$

Answer: D

Solution:

The cell constant (G^*) can be calculated using the formula that relates it to resistance (R) and conductivity (κ):

$G^* = R \times \kappa$

Given:

Resistance $R = 115\ \mathrm{ohm}$

Conductivity $\kappa = 1.90 \times 10^{-6}\ \mathrm{S\ cm}^{-1}$

Substitute these values into the formula:

$G^* = 115\ \mathrm{ohm} \times 1.90 \times 10^{-6}\ \mathrm{S\ cm}^{-1}$

Perform the multiplication:

$G^* = 218.5 \times 10^{-6}\,\mathrm{cm}^{-1}$

Convert to a more conventional form:

$G^* = 0.2185\,\mathrm{cm}^{-1}$

Thus, the value of the cell constant is approximately $0.218\,\mathrm{cm}^{-1}$.

The correct option is **Option D**: $0.218\,\mathrm{cm}^{-1}$.

Question69

Which of the following is used to avoid leakage of electrolyte in dry cell?

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Options:

A. Paste of MnO_2

B. Paste of NH_4Cl

C. Starch

D. Paste of $ZnCl_2$

Answer: C

Solution:

Option C, Starch, is often used in dry cells to avoid the leakage of the electrolyte. The starch acts as a binder and thickening agent to stabilize the electrolyte mixture, preventing it from leaking out of the cell and maintaining the integrity and functionality of the dry cell.

Question70

What is the quantity of electricity required to produce 4.8 g of Mg (molar mass $= 24\ \mathrm{g\ mol}^{-1}$) from its salt solution?

Options:

A. 10 F

B. 4 F

C. 1 F

D. 0.4 F

Answer: D

Solution:

To determine the quantity of electricity required to produce 4.8 g of Mg from its salt solution, it's crucial to first understand the electrochemical process involved in reducing Mg^{2+} ions to metallic magnesium. The half-reaction is:

$Mg^{2+} + 2e^{-} \rightarrow Mg$

This indicates that 2 moles of electrons (2 Faradays) are needed to reduce 1 mole of Mg^{2+} ions.

Calculate the number of moles of Mg to be produced:

$\text{Moles of Mg} = \frac{4.8 \text{ g}}{24 \text{ g/mol}} = 0.2 \text{ mol}$

Determine the total charge required:

Since 1 mole of Mg requires 2 Faradays of charge, 0.2 moles will require:

$0.2 \text{ mol} \times 2 \text{ F/mol} = 0.4 \text{ F}$

Therefore, the quantity of electricity required is **0.4 Faradays**.

Thus, Option D (0.4 F) is the correct answer.

Question71

For which electrolyte the molar conductivity at infinite dilution can not be obtained graphically?

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Options:

A. HNO_3

B. H_2SO_4

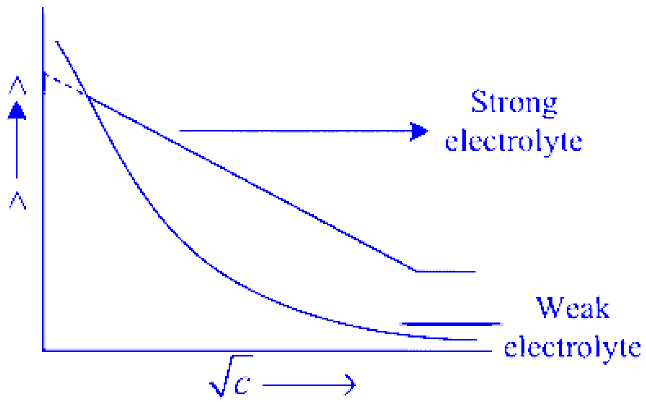
C. CH_3COOH

D. KCl

Answer: C

Solution:

The molar conductivity of weak electrolyte like CH_3COOH does not vary linearly with square root of concentration as shown in the figure. As a result, Λ_0 cannot be determined by extrapolation method graphically unlike strong electrolytes.



Variation of ∧ with $\sqrt{c}$

Question72

What volume of chlorine gas (molar mass 71) is evolved at STP during electrolysis of fused NaCl by passage of 1 amp current for 965 second? (At STP, V $= 22.4\ \mathrm{dm}^3$)

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Options:

A. 0.112 L

B. 0.224 L

C. 1.12 L

D. 2.24 L

Answer: A

Solution:

During the electrolysis of fused NaCl, chlorine gas (Cl_2) is evolved at the anode. The balanced chemical equation for this process is:

$2NaCl \rightarrow 2Na + Cl_2$

Each mole of Cl_2 is produced by the transfer of 2 moles of electrons.

1 Faraday (1 F or 96500 C) is the quantity of charge required to deposit 1 mole of electrons or equivalent of substance. For Cl_2, which involves 2 moles of electrons, the charge needed is:

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

To find out how much Cl_2 is evolved, calculate the moles of electrons using:

$\text{Charge supplied} = \text{Current} \times \text{Time} = 1 \text{ A} \times 965 \text{ s} = 965 \text{ C}$

Moles of electrons:

$\text{Moles of electrons} = \frac{\text{Charge supplied}}{\text{1 Faraday}} = \frac{965}{96500} = 0.01$

Since each mole of Cl_2 requires 2 moles of electrons:

$\text{Moles of } Cl_2 = \frac{0.01}{2} = 0.005$

Using the molar volume at STP:

$\text{Volume of } Cl_2 = \text{Moles of } Cl_2 \times 22.4 \text{ L/mol} = 0.005 \times 22.4 = 0.112 \text{ L}$

Thus, the volume of chlorine gas evolved at STP is 0.112 L. Therefore, the correct option is:

Option A: 0.112 L

Question73

If E° cell for $Cd_{(s)} \left| Cd^{2+}_{(IM)} \square Ag^{+}_{(1M)} \right| Ag_{(s)}$ is 1.2 V . What is the emf of the cell at 25°C ?

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Options:

A. $-1.2 \mathrm{~V}$

B. $2.4 \mathrm{~V}$

C. $-2.4 \mathrm{~V}$

D. $1.2 \mathrm{~V}$

Answer: D

Solution:

The electrode reaction is:

$\mathrm{Cd}+2 \mathrm{Ag}^{+} \longrightarrow \mathrm{Cd}^{2+}+2 \mathrm{Ag}$

For this cell reaction, $\mathrm{n}=2$.

Using Nernst equation at 298 K :

$$
\begin{aligned}
\mathrm{E}_{\text {cell }} & =\mathrm{E}_{\text {cell }}^{o} \quad-\frac{0.0592}{2} \log _{10} \frac{\left[\mathrm{Cd}^{2+}\right]}{\left[\mathrm{Ag}^{+}\right]^{2}} \\
& =1.2-\frac{0.0592}{2} \log _{10} \frac{1}{1} \\
\therefore \quad \mathrm{E}_{\text {cell }} & =1.2-0=1.2 \mathrm{~V} \quad\left(\because \log _{10} 1=0\right)
\end{aligned}
$$

Question74

What is the conductivity of $0.02 \mathrm{~M} \mathrm{AgNO}_{3}$ solution having cell constant $1.1 \mathrm{~cm}^{-1}$ and resistance is 94.5 ohms?

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Options:

A. $0.0216 \mathrm{ohm}^{-1} \mathrm{~cm}^{-1}$

B. $0.0116 \text{ ohm}^{-1} \text{ cm}^{-1}$

C. $0.0416 \text{ ohm}^{-1} \text{ cm}^{-1}$

D. $0.0859 \text{ ohm}^{-1} \text{ cm}^{-1}$

Answer: B

Solution:

The conductivity (κ) of a solution is calculated using the formula:

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

where R is the resistance of the solution and the Cell Constant is specific to the conductivity cell used. Given:

Resistance (R) = 94.5 ohms

Cell Constant = 1.1 cm^{-1}

Substituting the known values:

$\kappa = \frac{1}{94.5} \times 1.1$

Calculating the value:

First, find the reciprocal of the resistance:

$\frac{1}{94.5} \approx 0.010582$

Multiply by the Cell Constant:

$0.010582 \times 1.1 = 0.0116402$

Thus, the conductivity of the 0.02 M AgNO_3 solution is approximately:

$\kappa \approx 0.0116 \text{ ohm}^{-1} \text{ cm}^{-1}$

Therefore, the correct option is **B**: $0.0116 \text{ ohm}^{-1} \text{ cm}^{-1}$.

Question75

Which of the following expressions for conductivity of solution of an electrolyte is NOT correct?

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Options:

A. $\mathrm{k} = \frac{1}{\varrho}$

B. $\mathrm{k} = \mathrm{G}\frac{1}{\mathrm{a}}$

C. $\mathrm{k} = \frac{1}{\mathrm{R}}\frac{\mathrm{a}}{1}$

D. $\mathrm{k} = \Lambda.\mathrm{c}$

Answer: B

Solution:

Step-by-Step Reasoning:

Recall standard definitions and formulas:

Conductivity (κ) is defined as the reciprocal of resistivity (ρ). Thus:

$\kappa = \frac{1}{\rho}.$

For a given cell, the relationship between conductivity and conductance (G) is:

$\kappa = G \times (\text{cell constant}).$

The cell constant is often denoted by $\frac{l}{A}$ (length of the conducting path over the cross-sectional area).

Checking each option:

Option A: $\kappa = \frac{1}{\rho}$

This is the fundamental definition: Conductivity is the inverse of resistivity.

This is correct.

Option B: $\kappa = G\frac{1}{a}$

If a represents the cell constant (which is usually $\frac{l}{A}$), the correct formula is:

$\kappa = G \times a$

not $\kappa = G\frac{1}{a}.$

This does not match the standard relationship, making this suspicious.

Likely incorrect.

Option C: $\kappa = \frac{1}{R}\frac{a}{1}$

Since $G = \frac{1}{R}$, this can be rewritten as:

$\kappa = G \times a.$

This is the correct formula if a is the cell constant.

This is correct.

Option D: $\kappa = \Lambda \cdot c$

Molar conductivity Λ is defined by:

$\Lambda = \frac{\kappa}{c} \implies \kappa = \Lambda \times c.$

This is a standard relation.

This is correct.

Conclusion:

The only formula that does not align with the standard relationships is:

$\boxed{\text{Option B: } \kappa = G \frac{1}{a}}$

This formula is not correct based on standard definitions.

Final Answer: Option B is NOT correct.

Question76

Which of the following reactions occurs at cathode during discharging of lead accumulator?

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Options:

A. $PbSO_{4(s)} + 2e^- \longrightarrow Pb_{(s)} + SO_{4\,(aq.)}^{2-}$

B. $Pb_{(s)} + SO_{4(oq)}^{2-} \longrightarrow PbSO_{4(s)} + 2e^-$

C. $PbO_{2(s)} + 4H_{(aq)}^{+} + SO_{4(qq)}^{2-} + 2e^- \longrightarrow PbSO_{4(s)} + 2H_2O_{(1)}$

D. $PbSO_{4(s)} + 2H_2O_{(1)} \longrightarrow PbO_{2(s)} + 4H_{(aq)}^{+} + SO_{4(aq)}^{2-} + 2e^-$

Answer: C

Solution:

During the discharging of a lead accumulator, the reaction that occurs at the cathode is:

$$PbO_{2(s)} + 4H^+_{(aq)} + SO^{2-}_{4(aq)} + 2e^- \longrightarrow PbSO_{4(s)} + 2H_2O_{(l)}$$

This reaction involves the reduction of lead dioxide (PbO_2) present at the cathode. Hydrogen ions (H^+) and sulfate ions (SO_4^{2-}) from the sulfuric acid electrolyte participate in this reaction, forming solid lead sulfate ($PbSO_4$) and water. The presence of electrons ($2e^-$) on the reactant side indicates that reduction is taking place, which is characteristic of the cathode during the discharging process in a lead-acid battery.

Question77

The conductivity of 0.005 M NaI solution at $25^\circ C$ is $6.07 \times 10^{-4} \Omega^{-1}\ cm^{-1}$. Calculate its molar conductivity

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Options:

A. $121.4 \Omega^{-1}\ cm^2\ mol^{-1}$

B. $110.1 \Omega^{-1}\ cm^2\ mol^{-1}$

C. $201.1 \Omega^{-1}\ cm^2\ mol^{-1}$

D. $241.4 \Omega^{-1}\ cm^2\ mol^{-1}$

Answer: A

Solution:

The molar conductivity (Λ_m) of an electrolyte solution is calculated using the formula:

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

where:

Λ_m is the molar conductivity in $\Omega^{-1} cm^2 mol^{-1}$,

κ is the conductivity in $\Omega^{-1} cm^{-1}$,

c is the concentration in mol cm^{-3}.

Given:

$$\kappa = 6.07 \times 10^{-4} \Omega^{-1} cm^{-1}$$

$c = 0.005$ M, which converts to mol $\text{cm}^{-3} = 0.005 \times 10^{-3}$ mol $\text{cm}^{-3} = 5 \times 10^{-6}$ mol cm^{-3}

Now, substituting these values into the formula:

$\Lambda_m = \frac{6.07 \times 10^{-4}}{5 \times 10^{-6}}$

Perform the division:

$\Lambda_m = 121.4 \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$

Thus, the molar conductivity of the 0.005 M NaI solution at 25°C is:

Option A: $121.4 \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$

Question78

Calculate the amount of electricity required in coulombs to convert 0.08 mol of MnO_4^- to Mn^{2+}.

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Options:

A. 96500 C

B. 38600 C

C. 48250 C

D. 19300 C

Answer: B

Solution:

The conversion of MnO_4^- to Mn^{2+} involves a reduction process. The balanced half-reaction for this process in acidic solution is:

$\text{MnO}_4^- + 8\text{H}^+ + 5e^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$

In this reaction, one mole of MnO_4^- requires 5 moles of electrons ($5e^-$) to be reduced to Mn^{2+}.

To calculate the total charge in coulombs needed to reduce 0.08 moles of MnO_4^-, follow these steps:

Determine the total moles of electrons needed:

Moles of electrons $= 0.08\,\text{mol} \times 5 = 0.40\,\text{mol}\,e^-$

Use Faraday's constant to convert moles of electrons into coulombs.

Faraday's constant (F) is approximately 96500 coulombs per mole of electrons:

$Q = \text{Moles of electrons} \times F = 0.40\,\text{mol} \times 96500\,\frac{\text{C}}{\text{mol}} = 38600\,\text{C}$

Thus, the amount of electricity required is **38600 C**.

So the correct answer is **Option B: 38600 C**.

Question79

The limiting molar conductivities (Λ_0) for NaCl , KBr and KCl are 126,152 and $150\ \mathrm{S\ cm^2\ mol^{-1}}$ respectively. What is the Λ_0 of NaBr ?

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Options:

A. $128\ \mathrm{S\ cm^2\ mol^{-1}}$

B. $302\ \mathrm{S\ cm^2\ mol^{-1}}$

C. $278\ \mathrm{S\ cm^2\ mol^{-1}}$

D. $176\ \mathrm{S\ cm^2\ mol^{-1}}$

Answer: A

Solution:

The limiting molar conductivity (Λ_0) of an electrolyte can be determined using Kohlrausch's Law of Independent Migration of Ions. According to this law, the limiting molar conductivity of an electrolyte can be expressed as the sum of the limiting molar conductivities of its constituent ions.

Given:

$\Lambda_0(\mathrm{NaCl}) = 126\ \mathrm{S\ cm^2\ mol^{-1}}$

$\Lambda_0(\mathrm{KBr}) = 152\ \mathrm{S\ cm^2\ mol^{-1}}$

$\Lambda_0(\mathrm{KCl}) = 150\ \mathrm{S\ cm^2\ mol^{-1}}$

The limiting molar conductivity of NaBr, $\Lambda_0(\mathrm{NaBr})$, can be calculated as follows:

Express the given data in terms of ionic conductivities:

$\Lambda_0(\mathrm{NaCl}) = \Lambda_0(\mathrm{Na}^+) + \Lambda_0(\mathrm{Cl}^-)$

$\Lambda_0(\mathrm{KBr}) = \Lambda_0(\mathrm{K}^+) + \Lambda_0(\mathrm{Br}^-)$

$\Lambda_0(\mathrm{KCl}) = \Lambda_0(\mathrm{K}^+) + \Lambda_0(\mathrm{Cl}^-)$

Solve for the individual ionic contributions ($\Lambda_0(\mathrm{Na}^+)$, $\Lambda_0(\mathrm{Cl}^-)$, $\Lambda_0(\mathrm{K}^+)$, $\Lambda_0(\mathrm{Br}^-)$) using simultaneous equations:

From NaCl and KCl:

$\Lambda_0(\mathrm{Na}^+) + \Lambda_0(\mathrm{Cl}^-) = 126$

$\Lambda_0(\mathrm{K}^+) + \Lambda_0(\mathrm{Cl}^-) = 150$

Subtracting the first equation from the second:

$\Lambda_0(\mathrm{K}^+) - \Lambda_0(\mathrm{Na}^+) = 24$

$\Lambda_0(\mathrm{K}^+) = \Lambda_0(\mathrm{Na}^+) + 24$

From KBr and KCl:

$\Lambda_0(\mathrm{K}^+) + \Lambda_0(\mathrm{Br}^-) = 152$

Since $\Lambda_0(\mathrm{K}^+) = \Lambda_0(\mathrm{Na}^+) + 24$, substitute into the above:

$(\Lambda_0(\mathrm{Na}^+) + 24) + \Lambda_0(\mathrm{Br}^-) = 152$

$\Lambda_0(\mathrm{Na}^+) + \Lambda_0(\mathrm{Br}^-) = 128$

Substitute $\Lambda_0(\mathrm{Na}^+)$ and $\Lambda_0(\mathrm{Br}^-)$ back to find $\Lambda_0(\mathrm{NaBr})$:

$\Lambda_0(\mathrm{NaBr}) = \Lambda_0(\mathrm{Na}^+) + \Lambda_0(\mathrm{Br}^-) = 128$

Thus, the limiting molar conductivity of NaBr is:

$\Lambda_0(\mathrm{NaBr}) = 128\ \mathrm{S\ cm^2\ mol^{-1}}$

Correct answer: Option A: $128\ \mathrm{S\ cm^2\ mol^{-1}}$

Question80

A conductivity cell dipped on 0.05 M KCl has resistance 600 ohm. If conductivity is $0.0015 \mathrm{ohm}^{-1}\ \mathrm{cm}^{-1}$. What is the value of cell constant?

Options:

A. $0.47 \mathrm{~cm}^{-1}$

B. $0.90 \mathrm{~cm}^{-1}$

C. $1.5 \mathrm{~cm}^{-1}$

D. $2.5 \mathrm{~cm}^{-1}$

Answer: B

Solution:

The cell constant can be calculated using the formula:

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

Given:

Conductivity $\sigma = 0.0015 \, \text{ohm}^{-1} \, \text{cm}^{-1}$

Resistance $R = 600 \, \text{ohm}$

Substitute the values into the formula:

$\text{Cell Constant} = 0.0015 \, \text{ohm}^{-1} \, \text{cm}^{-1} \times 600 \, \text{ohm}$

Calculating the cell constant:

$\text{Cell Constant} = 0.9 \, \text{cm}^{-1}$

Thus, the value of the cell constant is:

Option B: $0.90 \, \text{cm}^{-1}$

Question81

What is the cell constant, if two platinum electrodes in conductivity cell are separated by 0.92 cm and area of cross section is $1.2 \mathrm{~cm}^2$?

Options:

A. $0.767 \mathrm{~cm}^{-1}$

B. $0.906 \mathrm{~cm}^{-1}$

C. $1.304 \mathrm{~cm}^{-1}$

D. $1.104 \mathrm{~cm}^{-1}$

Answer: A

Solution:

The cell constant is calculated using the formula:

Cell Constant $= \frac{L}{A}$

where:

L is the distance between the electrodes,

A is the area of cross-section.

Given:

$L = 0.92\,\mathrm{cm}$,

$A = 1.2\,\mathrm{cm}^2$.

Substituting these values into the formula gives:

Cell Constant $= \frac{0.92\,\mathrm{cm}}{1.2\,\mathrm{cm}^2} = 0.767\,\mathrm{cm}^{-1}$

Thus, the cell constant is $0.767\,\mathrm{cm}^{-1}$. Therefore, the correct option is:

Option A: $0.767\,\mathrm{cm}^{-1}$

Question82

The molar conductivity of 0.02 M KCl solution is $410 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$ at $25^{\circ} \mathrm{C}$. Calculate its conductivity?

Options:

A. $8.2 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}$

B. $2.8 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}$

C. $4.1 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}$

D. $5.4 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}$

Answer: A

Solution:

The relationship between molar conductivity (Λ_m) and conductivity (κ) is given by the formula:

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

where Λ_m is the molar conductivity, κ is the conductivity, and C is the concentration of the solution in mol/L.

Given:

Molar conductivity, $\Lambda_m = 410\,\Omega^{-1}\,\mathrm{cm}^2\,\mathrm{mol}^{-1}$

Concentration, $C = 0.02\,\mathrm{M}$

Rearranging the formula to solve for conductivity κ:

$\kappa = \Lambda_m \times C$

Substituting the given values:

$\kappa = 410\,\Omega^{-1}\,\mathrm{cm}^2\,\mathrm{mol}^{-1} \times 0.02\,\mathrm{mol/L}$

Calculating κ:

$\kappa = 8.2\,\Omega^{-1}\,\mathrm{cm}^{-1}$

Therefore, the conductivity of the solution is $8.2 \times 10^{-3}\,\Omega^{-1}\,\mathrm{cm}^{-1}$.

Thus, the correct option is **Option A**: $8.2 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}$.

Question83

Calculate the amount of electricity required to convert 1.1 mol of $Cr_2O_7^{2-}$ to Cr^{+3} in acidic medium.

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Options:

A. $6.369 \times 10^5 C$

B. $1.462 \times 10^5 C$

C. $4.839 \times 10^5 C$

D. $3.419 \times 10^5 C$

Answer: A

Solution:

To calculate the amount of electricity required to convert 1.1 moles of $Cr_2O_7^{2-}$ to Cr^{3+}, follow these steps:

Step 1: Write the half-reaction for the reduction of $Cr_2O_7^{2-}$ in acidic medium.

$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

Step 2: Determine the number of electrons needed for the reduction of one mole of $Cr_2O_7^{2-}$.

From the balanced half-reaction, 6 moles of electrons are required to convert 1 mole of $Cr_2O_7^{2-}$ to 2 moles of Cr^{3+}.

Step 3: Calculate the total charge required using Faraday's Law of Electrolysis.

Faraday's constant (F) is 96485 C/mol.

For 1 mole of $Cr_2O_7^{2-}$, the charge required is:

$6 \text{ mol } e^- \times 96485 \text{ C/mol} = 578910 \text{ C}$

Since we have 1.1 moles of $Cr_2O_7^{2-}$, compute the total charge:

$1.1 \text{ mol} \times 578910 \text{ C/mol} = 636801 \text{ C}$

Therefore, the amount of electricity required to convert 1.1 mol of $Cr_2O_7^{2-}$ to Cr^{3+} is approximately $\boxed{6.369 \times 10^5 \text{ C}}$, which corresponds to Option A.

Question84

Which of the following statements is NOT correct for $H_2 - O_2$ fuel cell?

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Options:

A. $H_{2(g)}$ serves as fuel

B. $O_{2(g)}$ is an oxidising agent

C. Platinum wires are used as anode and cathode

D. Hot aqueous KOH solution acts as an electrolyte

Answer: C

Solution:

To determine which statement is **NOT** correct for an H_2-O_2 fuel cell, let's analyze each option:

Option A: $H_{2(g)}$ *serves as fuel.*

Explanation: In a hydrogen-oxygen fuel cell, hydrogen gas (H_2) acts as the fuel at the **anode** where it gets oxidized.

Reaction at Anode:

$$H_2 + 2OH^- \rightarrow 2H_2O + 2e^-$$

Conclusion: This statement is **correct**.

Option B: $O_{2(g)}$ *is an oxidizing agent.*

Explanation: Oxygen gas (O_2) acts as the oxidizing agent at the **cathode** where it gets reduced.

Reaction at Cathode:

$$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$$

Conclusion: This statement is **correct**.

Option C: *Platinum wires are used as anode and cathode.*

Explanation:

In an H_2-O_2 fuel cell, the electrodes are typically **porous carbon electrodes** impregnated with catalysts like platinum, palladium, or nickel to facilitate the reactions.

The use of **platinum wires** as electrodes is **not standard practice** because:

Platinum is expensive, and using it as solid wires would not be cost-effective.

Porous electrodes increase the surface area for the reaction and improve efficiency.

Conclusion: This statement is **NOT correct**.

Option D: *Hot aqueous KOH solution acts as an electrolyte.*

Explanation:

An aqueous solution of **potassium hydroxide (KOH)** is commonly used as the **electrolyte** in alkaline fuel cells.

The electrolyte must conduct ions (here, OH^- ions) to complete the electrical circuit.

Operating at elevated temperatures improves the ionic conductivity of the electrolyte.

Conclusion: This statement is **correct**.

Final Answer:

Option C is **NOT correct** because **platinum wires are not used as the anode and cathode** in an H_2-O_2 fuel cell; instead, porous carbon electrodes with catalytic coatings are used.

Answer: Option C

Question85

The resistance of a conductivity cell of 0.1 M KCl solution is 120 ohm and conductivity is $1.64 \times 10^{-4}\ \mathrm{S\ cm^{-1}}$. What is the value of cell constant?

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Options:

A. $0.0136\ \mathrm{cm}^{-1}$

B. $0.0196\ \mathrm{cm}^{-1}$

C. $0.0618 \mathrm{~cm}^{-1}$

D. $0.0731 \mathrm{~cm}^{-1}$

Answer: B

Solution:

The cell constant (G^*) of a conductivity cell is determined by the relationship between the conductivity (κ) of the solution and the resistance (R) of the cell. The formula for the cell constant is given by:

$G^* = \kappa \times R$

Given that the conductivity of the 0.1 M KCl solution is $\kappa = 1.64 \times 10^{-4} \mathrm{~S~cm}^{-1}$ and the resistance $R = 120 \, \Omega$, we can calculate the cell constant as follows:

Substitute the known values into the formula:

$G^* = (1.64 \times 10^{-4} \mathrm{~S~cm}^{-1}) \times (120 \, \Omega)$

Calculate the result:

$G^* = 1.64 \times 120 \times 10^{-4} \mathrm{~cm}^{-1}$

$G^* = 196.8 \times 10^{-4} \mathrm{~cm}^{-1}$

Convert to standard form:

$G^* = 0.01968 \mathrm{~cm}^{-1}$

Therefore, the value of the cell constant is approximately $0.0196 \mathrm{~cm}^{-1}$.

Hence, the correct option is **Option B:** $0.0196 \mathrm{~cm}^{-1}$.

Question86

Which of the following changes occurs during the discharging of lead accumulator?

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Options:

A. $Pb_{(s)}$ is reduced

B. H_2SO_4 is consumed

C. $PbSO_4$ is consumed

D. PbO_2 is produced

Answer: B

Solution:

During discharge, the following reaction occurs:

During discharge, the
$Pb_{(s)} + PbO_{2(\mathrm{~s})} + 2H_2SO_{4(aq)} \rightarrow 2PbSO_{4(\mathrm{~s})} + 2H_2O_{(1}$

As the cell operates to generate current, H_2SO_4 is consumed.

Question87

Calculate the quantity of electricity required to liberate 0.1 mole of chlorine gas during electrolysis of molten sodium chloride.

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Options:

A. 9665 C

B. 19300 C

C. 14500 C

D. 96500 C

Answer: B

Solution:

During electrolysis of molten NaCl , the following reaction takes place at anode:

$2Cl^{-} \longrightarrow Cl_2 + 2e^{-}$

1 mole of Cl_2 is liberated with 2 moles of electrons.

$\therefore$ 0.1 mole of Cl_2 will be liberated with 0.2 moles of electrons.

$\therefore$ Quantity of electricity required $= 0.2 \times 96500 C$
$= 19300 C$

Question88

Which from following expressions is used to calculate E_{cell} for the following cell at $25^{\circ}C$?

$$Pb_{(s)} \left| Pb^{++}_{(1M)} \| Ag^{+}_{(10M)} \right| Ag_{(s)}$$

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Options:

A. $E_{cell} = (E^{\circ}_{cell} + 0.0592) V$

B. $E_{cell} = \left(E^{\circ}_{cell} - 0.0592\right) V$

C. $E_{cell} = \left(E^{\circ}_{cell} - 0.0296\right) V$

D. $E_{cell} = (E^{\circ}_{cell} + 0.0296) V$

Answer: A

Solution:

Cell representation:

$Pb_{(s)} \left| Pb^{++}_{(1M)} \right| \left| Ag^{+}_{(10M)} \right| Ag_{(s)}$

Reaction: $Pb_{(s)} + 2Ag^{+}_{(10M)} \rightarrow Pb^{++}_{(1M)} + 2Ag_{(s)}$

Nernst equation at $25^{\circ}C$:

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{2} \log_{10} \frac{\left[Pb^{++}\right]}{\left[Ag^{+}\right]^{2}}$$

$$\begin{aligned}
E_{cell} &= E^{\circ}_{cell} - \frac{0.0592}{2}\log_{10}\frac{1}{100} \\
E_{cell} &= E^{\circ}_{cell} - \frac{0.0592}{2}\log_{10}10^{-2} \\
E_{cell} &= E^{\circ}_{cell} - \frac{0.0592}{2} \times (-2) \\
E_{cell} &= (E^{\circ}_{cell} + 0.0592)\,V
\end{aligned}$$

Question89

Which from following is NOT true about electrolysis of molten NaCl ?

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Options:

A. Cl_2 gas is liberated at anode.

B. Na is deposited at cathode.

C. The decomposition of NaCl into $Na_{(s)}$ and $Cl_{2(g)}$ is spontaneous.

D. Electrical energy is used to carry out the reaction.

Answer: C

Solution:

Option C is NOT true about the electrolysis of molten NaCl.

During the electrolysis of molten NaCl:

At the anode, chloride ions (Cl^-) are oxidized to chlorine gas (Cl_2), which matches Option A.

At the cathode, sodium ions (Na^+) are reduced to solid sodium ($Na_{(s)}$), which matches Option B.

The process requires electrical energy to decompose NaCl, which is a non-spontaneous reaction. Option D correctly states that electrical energy is used to carry out the reaction.

Therefore, the decomposition of NaCl into $Na_{(s)}$ and $Cl_{2(g)}$ without external energy input (as described in Option C) is not spontaneous, making Option C incorrect.

Question90

The resistance of decimolar solution of NaCl is 30 ohms. Calculate the conductivity of solution if the cell constant is 0.33 cm^{-1}.

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Options:

A. $0.025 \Omega^{-1} \text{ cm}^{-1}$

B. $0.035 \Omega^{-1} \text{ cm}^{-1}$

C. $0.011 \Omega^{-1} \text{ cm}^{-1}$

D. $0.029 \Omega^{-1} \text{ cm}^{-1}$

Answer: C

Solution:

The conductivity of a solution is given by the formula:

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

where:

κ is the conductivity (in $\Omega^{-1} \text{ cm}^{-1}$),

R is the resistance (in Ω),

G is the cell constant (in cm^{-1}).

Given:

Resistance, $R = 30\,\Omega$

Cell constant, $G = 0.33\,\text{cm}^{-1}$

Substituting these values into the formula, we get:

$\kappa = \frac{1}{30} \cdot 0.33$

Calculate the conductivity:

$\kappa = 0.011\,\Omega^{-1}\,\text{cm}^{-1}$

Therefore, the correct option is **Option C**: $0.011\Omega^{-1}\text{ cm}^{-1}$.

Question91

A conductivity cell containing 0.001 M AgNO_3 solution develops resistance 6530ohm at 25°C. Calculate the electrical conductivity of solution at same temperature if the cell constant is 0.653 cm^{-1}.

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Options:

A. $1.3\times10^{-4}\Omega^{-1}\text{ cm}^{-1}$

B. $1.5\times10^{-4}\Omega^{-1}\text{ cm}^{-1}$

C. $1.7\times10^{-4}\Omega^{-1}\text{ cm}^{-1}$

D. $1.0\times10^{-4}\Omega^{-1}\text{ cm}^{-1}$

Answer: D

Solution:

$$\begin{aligned} &\text{Cell constant } = \text{k}\times\text{R} \\ \therefore\quad & 0.653 = \text{k}\times6530 \\ \therefore\quad & \text{k} = \tfrac{0.653}{6530} = 1\times10^{-4}\Omega^{-1}\text{ cm}^{-1} \end{aligned}$$

Question92

Calculate E^0_{cell} for $\text{Cd}_{(\text{s})}\left|\text{Cd}^{++}_{(1\text{M})}\right|\left|\text{Ag}^{+}_{(1\text{M})}\right|\text{Ag}_{(\text{s})}$.

$$\left[\text{E}^0_{\text{Cd}} = -0.403\text{ V}; \text{E}^0_{\text{Ag}} = 0.799\text{ V}\right]$$

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Options:

A. 1.202 V

B. -1.202 V

C. 0.396 V

D. -0.396 V

Answer: A

Solution:

For the given cell, anode is Cd and cathode is Ag.

$$\begin{aligned} E^0_{cell} &= E^0_{cathode} - E^0_{anode} \\ &= 0.799-(-0.403) \\ &= 1.202 \text{ V} \end{aligned}$$

Question93

Which from following expressions is used to find the cell potential of $\mathrm{Cd}_{(s)}\left|\mathrm{Cd}^{++}_{(aq)}\right|\left|\mathrm{Cu}^{+}_{(aq)}\right|\mathrm{Cu}_{(s)}$ cell at $25^\circ\mathrm{C}$?

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Options:

A. $E_{cell} = E^o_{cell} - 0.0296 \log \frac{[Cd^{++}]}{[Cu^{++}]}$

B. $E_{cell} = E^o_{cell} + 0.0296 \log \frac{[Cd^{++}]}{[Cu^{++}]}$

C. $E_{cell} = E^o_{cell} - 0.0592 \log \frac{[Cu^{++}]}{[Cd^{++}]}$

D. $E_{cell} = E^\circ_{cell} + 0.0592 \log \frac{[Cu^{++}]}{[Cd^{++}]}$

Answer: A

Solution:

Cell reaction

$$Cd_{(s)} \longrightarrow Cd^{2+}_{(s)} + 2e \quad \text{(oxidation at anode)}$$

$$Cu^{2+}_{(aq)} + 2e^- \longrightarrow Cu_{(s)} \quad \text{(reduction at cathode)}$$

$$Cd_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Cd^{2+}_{(aq)} + 2Cu_{(s)}$$

$$\text{(overall reaction)}$$

Now,

$$E_{cell} = E^o_{cell} - \frac{0.0592\ V}{n} \log_{10} \frac{[\text{ Product }]}{[\text{ Reactant }]}$$

$$= 0.02 - \frac{0.0592\ V}{2} \log_{10} \frac{[Cd^{++}]}{[Cu^{++}]}$$

$$= E^o_{cell} - 0.0296 \log \frac{[Cd^{++}]}{[Cu^{++}]}$$

Question94

Which from following is NOT true about voltaic cell?

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Options:

A. The anode acts as negative electrode.

B. The cathode acts as a positive electrode.

C. It converts electrical energy into chemical energy

D. Dry cell is an example of voltaic cell.

Answer: C

Solution:

The correct answer is Option C: "It converts electrical energy into chemical energy." This statement is not true about a voltaic cell. Let's break down each of the given options to understand why Option C is incorrect:

Option A: **The anode acts as negative electrode**. This statement is true. In a voltaic cell, the anode is where oxidation occurs, meaning that it is the site where electrons are lost by the reactants. Because electrons are negatively charged, the anode accumulates negative charge relative to the cathode, so it acts as the negative electrode.

Option B: **The cathode acts as a positive electrode**. This statement is also true. In a voltaic cell, the cathode is where reduction occurs, meaning that it is the site where the electrons are gained by the reactants. Since the electrons come from the anode, the cathode acquires a positive charge relative to the anode, and hence it serves as the positive electrode.

Option C: **It converts electrical energy into chemical energy**. This statement is not true and is a characteristic of an electrolytic cell, not a voltaic cell. A voltaic cell, also known as a galvanic cell, operates on the principle of a spontaneous chemical reaction that produces electrical energy. It converts chemical energy into electrical energy, which can be used to do work.

Option D: **Dry cell is an example of a voltaic cell**. This statement is true. A dry cell is a common type of voltaic cell that is used in many household batteries. It generates electrical energy through spontaneous redox reactions occurring within the cell.

So, to reiterate, Option C is NOT true about a voltaic cell, as a voltaic cell indeed converts chemical energy into electrical energy, not the other way around.

Question95

Calculate molar conductivity of $\mathrm{NH_4OH}$ at infinite dilution if molar conductivities of $\mathrm{Ba(OH)_2}$ $\mathrm{BaCl_2}$ and $\mathrm{NH_4Cl}$ at infinite dilution are $520, 280, 129 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$ respectively.

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Options:

A. $249.0 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

B. $498.0 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

C. $125.0 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

D. $369.0\ \Omega^{-1}\ \mathrm{cm}^2\ \mathrm{mol}^{-1}$

Answer: A

Solution:

According to Kohlrausch law,

i. $\wedge_0\left(\mathrm{Ba(OH)}_2\right)=\lambda^0_{\mathrm{Ba}^{2+}}+2\lambda^0_{\mathrm{OH}^-}$

ii. $\wedge_0\left(\mathrm{BaCl}_2\right)=\lambda^0_{\mathrm{Ba}^{2+}}+2\lambda^0_{\mathrm{Cl}^-}$

iii. $\wedge_0\left(\mathrm{NH_4Cl}\right)=\lambda^0_{\mathrm{NH_4}^+}+\lambda^0_{\mathrm{Cl}^-}$

Eq. (i) + $\frac{1}{2}$ Eq (ii) − $\frac{1}{2}$ Eq (iii) gives

$$\wedge_0\left(\mathrm{NH_4OH}\right)=\wedge_0\left(\mathrm{NH_4Cl}\right)+\frac{1}{2}\wedge_0\left(\mathrm{Ba(OH)}_2\right)-\frac{1}{2}\wedge_0\left(\mathrm{BaCl}_2\right)$$

$$\begin{aligned}\wedge_0\left(\mathrm{NH_4OH}\right)&=129+\frac{1}{2}520-\frac{1}{2}280\\&=249.0\Omega^{-1}\ \mathrm{cm}^2\ \mathrm{mol}^{-1}\end{aligned}$$

When using Kohlrausch law, remember to multiply molar ionic conductivities of cation and anion with the number of cations and anions, respectively, as in the chemical formula of the compound.

Question96

Which from the following expression represents molar conductivity of an electrolyte A_2B_3 type?

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Options:

A. $2\lambda^0_{A^{+++}}+3\lambda^0_{B^{--}}$

B. $3\lambda^0_{A^{+++}}+2\lambda^0_{B^{--}}$

C. $2\lambda^\circ_{A^{+++}}+\lambda^\circ_{B^{--}}$

D. $\lambda^\circ_{A^{+++}}+3\lambda^\circ_{B^{--}}$

Answer: A

Solution:

For A_2B_3 type electrolyte molar conductivity is given by Kohlrausch's law as.

$\Lambda_{\mathrm{m}}^{\circ}\left(A_2B_3\right)=2\lambda_{A^{+++}}^{\circ}+3\lambda_{B^{--}}^{\circ}$

Question97

What is the molar conductivity of $0.005 \mathrm{~M} \mathrm{NaI}$ solution if it's conductivity is $6.065 \times 10^{-4} \Omega^{-1} \mathrm{~cm}^{-1}$?

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Options:

A. $121.3 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

B. $115.1 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

C. $126.5 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

D. $131.2 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

Answer: A

Solution:

The molar conductivity, often denoted as Λ_m, can be calculated by the formula:

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

where:

- κ represents the conductivity of the solution.
- C is the concentration of the solution in moles per liter.

In the given problem, the conductivity κ is provided as $6.065 \times 10^{-4} \Omega^{-1} cm^{-1}$, and the concentration C of the NaI solution is given as $0.005 M$.

For calculating molar conductivity, we must also ensure that the units are consistent, converting concentration to moles per cubic centimeter (since the conductivity unit involves cm) if necessary.

Here is how the calculation is performed:

$$\Lambda_m = \frac{6.065\times10^{-4}\ \Omega^{-1}\text{cm}^{-1}}{0.005\ mol\ L^{-1}} \times \frac{1000\ \text{cm}^3}{1\ \text{L}}$$

After multiplying and dividing as required, the result is:

$$\Lambda_m = 121.3\Omega^{-1}\text{cm}^2\text{mol}^{-1}$$

The answer reflects the molar conductivity of the NaI solution.

Question98

Calculate E° cell for following.

$\text{Zn}(s)\,|\text{Zn}^{++}(1\text{M})||\text{Pb}^{++}(1\text{M})|\text{Pb}(s)$ if $E^\circ_{\text{Zn}} = -0.763\ \text{V}$ and $E^\circ_{\text{Pb}} = -0.126\ \text{V}$

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Options:

A. $0.637\ \text{V}$

B. $-0.530\ \text{V}$

C. $-0.889\ \text{V}$

D. $0.789\ \text{V}$

Answer: A

Solution:

For the given reaction,

$$\begin{aligned} &\text{Zn}(s)\mid\text{Zn}^{2+}(1\text{M})||\text{Pb}^{2+}(1\text{M})\mid\text{Pb}(s) \\ &E^\circ_{\text{Zn}} = -0.763\ \text{V and } E^\circ_{\text{Pb}} = -0.126\ \text{V} \\ &E^\circ_{\text{cell}} \quad =? \end{aligned}$$

Using formula,

$$\begin{aligned} E^\circ_{\text{cell}} &= E_{\text{cathode}} - E_{\text{anode}} \\ &= -0.126 - (-0.763) \\ E^\circ_{\text{cell}} &= 0.637\ \text{V} \end{aligned}$$

Question99

Calculate E^0_{cell} if the equilibrium constant for following reaction is 1.2×10^6.

$$2Cu^+_{(aq)} \longrightarrow Cu^{++}_{(aq)} + Cu_{(s)}$$

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Options:

A. 0.36 V

B. -0.36 V

C. -0.18 V

D. 0.18 V

Answer: A

Solution:

$$E^o_{cell} = \frac{0.0592}{n} \log_{10} K \text{ at } 298 \text{ K}$$
$$Cu^+_{(aq)} + e^- \longrightarrow Cu_{(s)}$$
$$Cu^+_{(aq)} \longrightarrow Cu^{2+}_{(aq)} + e^-$$
$$2Cu^+_{(aq)} \longrightarrow Cu^{2+}_{(aq)} + Cu_{(s)}$$
$$\therefore \quad n = 1$$
$$\therefore \quad E^\circ_{cell} = \frac{0.0592}{1} \log_{10}\left(1.2 \times 10^6\right)$$
$$= 0.0592\left(\log 1.2 + \log 10^6\right)$$
$$= 0.0592(0.079 + 6)$$
$$= 0.0592 \times 6.079$$
$$= 0.36 \text{ V}$$

Question100

Identify the overall oxidation reaction that occurs in lead storage cell during discharge.

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Options:

A. $\mathrm{Pb}_{(\mathrm{aq})}^{++}+\mathrm{SO}_{4(\mathrm{aq})}^{--} \longrightarrow \mathrm{PbSO}_{4(\mathrm{s})}$

B. $\mathrm{PbSO}_{4(\mathrm{s})}+2 \mathrm{H}_2 \mathrm{O}_{(f)} \longrightarrow \mathrm{PbO}_{2(\mathrm{B})}+4 \mathrm{H}_{(\mathrm{aq})}^{+}+\mathrm{SO}_{4(\mathrm{aq})}^{--}+2 \mathrm{e}^{-}$

C. $\mathrm{Pb}_{(\mathrm{s})}+\mathrm{SO}_{4(\mathrm{aq})}^{--} \longrightarrow \mathrm{PbSO}_{4(\mathrm{s})}+2 \mathrm{e}^{-}$

D. $\mathrm{PbSO}_{4(\mathrm{s})}+2 \mathrm{e}^{-} \longrightarrow \mathrm{Pb}_{(\mathrm{s})}+\mathrm{SO}_{4(\mathrm{aq})}^{--}$

Answer: C

Solution:

When the lead storage cell provides current (1.e., during discharge), spongy lead (Pb) is oxidised to Pb^{2+} ions and negative charge accumulates on lead plates. The Pb^{2+} ions so formed combine with SO_4^{2-} ions from $\mathrm{H}_2\mathrm{SO}_4$ to form insoluble PbSO_4. The overall oxidation is the sum of these two processes.

$\mathrm{Pb}_{(\mathrm{s})} \longrightarrow \mathrm{Pb}_{(\mathrm{aq})}^{2+}+2 \mathrm{e}^{-}$ (oxidation)

$\mathrm{Pb}_{(\mathrm{aq})}^{2+}+\mathrm{SO}_{4(\mathrm{aq})}^{2-} \longrightarrow \mathrm{PbSO}_{4(\mathrm{s})}$ (precipitation)

$\mathrm{Pb}_{(\mathrm{s})}+\mathrm{SO}_{4(\mathrm{ag})}^{2-} \longrightarrow \mathrm{PbSO}_{4(\mathrm{s})}+2 \mathrm{e}^{-}$ (Overall oxidation)

Question101

Calculate the time required in second to deposit 6.35 g copper from its salt solution by passing 5 ampere current.[Molar mass of Cu $=63.5 \mathrm{~g} \mathrm{~mol}^{-1}$]

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Options:

A. 3600

B. 3700

C. 3860

D. 4000

Answer: C

Solution:

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

$$\text{Mole ratio} = \frac{1 \text{ mol}}{2 \text{ mole}^-}$$

$$\mathrm{W} = \frac{\mathrm{I(A)} \times \mathrm{t(s)}}{96500\,(\mathrm{C/mole^-})} \times \text{mole ratio} \times \text{molar mass}$$

$$6.35 \text{ g} = \frac{5 \times \mathrm{t}}{96500\,(\mathrm{C/mole^-})} \times \frac{1 \text{ mol}}{2 \text{ mole}^-} \times 63.5 \text{ g mol}^{-1}$$

$$\mathrm{t} = \frac{6.35 \times 96500 \times 2}{5 \times 63.5} = 3860 \text{ seconds}$$

Question102

Which of the following expressions represents molar conductivity of $\mathrm{AB_3}$ type electrolyte?

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Options:

A. $3\lambda_{\mathrm{A}^{\cdots}}^{\circ} + \lambda_{\mathrm{B}^-}^{\circ}$

B. $\lambda_{\mathrm{A}^{\cdots}}^{\circ} + \lambda_{\mathrm{B}^-}^{\circ}$

C. $\lambda_{\mathrm{A}^{\cdots}}^{\circ} + 3\lambda_{\mathrm{B}^-}^{\circ}$

D. $2\lambda_{\mathrm{A}^{\cdots}}^{\circ} + \lambda_{\mathrm{B}^-}^{\circ}$

Answer: C

Solution:

$\Lambda_0 = n_+ \lambda_+^0 + n_- \lambda_-^0$

Where, λ_+ and λ_- are molar conductivities of cation and anion, respectively, and n_+ and n are the number of moles of cation and anion specified in the chemical formula of an electrolyte.

Question103

Calculate the E_{cell} for $Zn_{(s)} \left| Zn^{++}_{(0.1M)} \right| \left| Cr^{+++}_{(0.1M)} \right| Cr_{(s)}$ at $25°C$ if E°_{cell} is 0.02 V

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Options:

A. -0.05 V

B. 0.03 V

C. -0.06 V

D. 0.07 V

Answer: B

Solution:

The reaction is

$3Zn_{(3)} + 2Cr^{3+}_{(0.1M)} \longrightarrow 3Zn^{2+}_{(0.1)} + 2Cr_{(s)}$

$$E_{cell} = E^o_{cell} - \frac{0.0592\ V}{n} \log_{10} \frac{\left[Zn^{2+}\right]^3}{\left[Cr^{3+}\right]^2}$$

$$\therefore \quad E_{cell} = 0.02\ V - \frac{0.0592\ V}{6} \log_{10} \frac{(0.1)^3}{(0.1)^2}$$

$$
\begin{aligned}
&= 0.02 \text{ V} - \frac{0.0592 \text{ V}}{6} \log_{10} 0.1 \\
&= 0.02 \text{ V} + \frac{0.0592 \text{ V}}{6} \times 1 \\
&= 0.02 \text{ V} + 0.0099 \text{ V} \\
&= 0.03 \text{ V}
\end{aligned}
$$

Question104

A conductivity cell containing 5×10^{-4} M NaCl solution develops resistance 14000 ohms at 25°C. Calculate the conductivity of solution if the cell constant is 0.84 cm^{-1}

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Options:

A. $6.0 \times 10^{-5} \Omega^{-1} \text{ cm}^{-1}$

B. $3.0 \times 10^{-5} \Omega^{-1} \text{ cm}^{-1}$

C. $9.0 \times 10^{-5} \Omega^{-1} \text{ cm}^{-1}$

D. $12.0 \times 10^{-5} \Omega^{-1} \text{ cm}^{-1}$

Answer: A

Solution:

$$
\begin{aligned}
k &= \frac{\text{cell constant}}{R} \\
&= \frac{0.84 \text{ cm}^{-1}}{14000 \Omega} \\
&= 6.0 \times 10^{-5} \Omega^{-1} \text{ cm}^{-1}
\end{aligned}
$$

Question105

Calculate ΔG° for the reaction
$Mg_{(s)} + Sn^{++}_{(aq)} \longrightarrow Mg^{++}_{(aq)} + Sn_{(s)}$ **if E^0_{cell} is 2.23 V.**

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Options:

A. -430.4 kJ

B. 215.2 kJ

C. 645.6 kJ

D. -860.8 kJ

Answer: A

Solution:

$$\begin{aligned} \Delta G^{\circ} &= -nFE^{\circ}_{cell} \\ &= -2 \times 96500 \times 2.23 \\ &= -430390 \text{ J} \\ &= -430.4 \text{ kJ} \end{aligned}$$

Question106

Electrolytic cells containing Zn and Al salt solutions are connected in series. If 6.5 g of Zn is deposited in one cell calculate mass of Al deposited in second cell (molar mass : $Zn = 65, Al = 27$) by passing definite quantity of electricity?

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Options:

A. 2.4 g

B. 2.1 g

C. 2.7 g

D. 1.8 g

Answer: D

Solution:

Cell 1:

$$\mathrm{Zn}_{(\mathrm{sq})}^{2+}+2\mathrm{e}^{-} \longrightarrow \mathrm{Zn}_{(\mathrm{s})}$$
$$(\text{mole ratio}))_1=\frac{1 \text{ mol}}{2 \text{ mole}^{-}}$$

Cell 2:

$$\mathrm{Al}_{(\mathrm{aq})}^{3+}+3\mathrm{e}^{-} \longrightarrow \mathrm{Al}_{(\mathrm{s})}$$
$$(\text{ mole ratio })_2=\frac{1 \text{ mol}}{3 \text{ mole}^{-}}$$

$$\frac{\mathrm{W}_1}{(\text{ mole ratio })_1 \times \mathrm{M}_1}=\frac{\mathrm{W}_2}{(\text{ mole ratio })_2 \times \mathrm{M}_2}$$
$$\frac{6.5 \text{ g}}{1 \text{ mol}/2 \text{ mole}^{-} \times 65 \text{ g mol}^{-1}}$$
$$=\frac{\mathrm{W}_2}{1 \text{ mol}/3 \text{ mole}^{-} \times 27 \text{ g mol}^{-1}}$$
$$\frac{6.5 \text{ g} \times 2}{65}=\frac{\mathrm{W}_2 \times 3}{27}$$
$$\mathrm{W}_2=\frac{6.5 \times 2 \times 27}{65 \times 3}=1.8 \text{ g}$$

Question107

Which from following is strongest reducing agent?

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Options:

A. K

B. Al

C. Mg

D. Ag

Answer: A

Solution:

The strength of reducing agents increases from top to bottom in the electrochemical series as E^0 values decrease. Among the given options, K has the lowest value of $\mathrm{E}^{\circ}(-2.925 \mathrm{~V})$.

Question108

Which from following sentences is NOT correct?

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Options:

A. $\Delta \mathrm{G}^{\circ}$ is an extensive property.

B. $\mathrm{E}_{\text{cell}}^{\circ}$ is an intensive property.

C. Electrical work is equal to $\mathrm{nFE}_{\text{cell}}$.

D. For a chemical reaction to be spontaneous $\mathrm{E}_{\text{cell}}^{0}$ must be negative.

Answer: D

Solution:

For a chemical reaction to be spontaneous $\mathrm{E}_{\text{cell}}^{\circ}$ must be positive.

Question109

Calculate the $\mathrm{E}_{\text{cell}}^{\circ}$ for $\mathrm{Zn}_{(\mathrm{s})}\left|\mathrm{Zn}_{(\mathrm{IM})}^{++}\right|\left|\mathrm{Cd}_{(\mathrm{IM})}^{++}\right| \mathrm{Cd}_{(\mathrm{s})}$ at $25^{\circ} \mathrm{C}\left[\mathrm{E}_{\mathrm{Zn}}^{0}=-0.763 \mathrm{~V} ; \mathrm{E}_{\mathrm{Cd}}^{\circ}=-0.403 \mathrm{~V}\right]$

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Options:

A. 0.36 V

B. 1.17 V

C. -0.36 V

D. -1.17 V

Answer: A

Solution:

For the given cell reaction, anode is Zn and cathode is Cd.

$$\begin{aligned} E^0_{cell} &= E^0_{cathode} - E^0_{anode} \\ &= -0.403 - (-0.763) \\ &= 0.36 \text{ V} \end{aligned}$$

Question110

Find the number of faradays of electricity required to produce 45 g of Al from molten $\mathrm{Al_2O_3}$.

(At. mass of $\mathrm{Al} = 27$)

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Options:

A. 1 F

B. 3 F

C. 5 F

D. 7 F

Answer: C

Solution:

$Al^{3+} + 3e^- \longrightarrow Al$

The equation shows that 3 moles of electrons are required to produce 1 mole of Al (i.e., 27 g of Al).

$\therefore$ 3 F of electricity is required to produce 27 g of Al from molten Al_2O_3.

$\therefore$ Faradays of electricity required to produce 45 g of Al $= \frac{3}{27} \times 45 = 5$ F

Question111

The molar conductivity of 0.02 M AgI at 298 K is $142.3\Omega^{-1}\ cm^2\ mol^{-1}$. What is its conductivity?

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Options:

A. $1.42 \times 10^{-3}\Omega^{-1}\ cm^{-1}$

B. $2.41 \times 10^{-3}\Omega^{-1}\ cm^{-1}$

C. $2.85 \times 10^{-3}\Omega^{-1}\ cm^{-1}$

D. $7.11 \times 10^{-3}\Omega^{-1}\ cm^{-1}$

Answer: C

Solution:

$$\wedge = \frac{1000k}{c}$$
$$k = \frac{\Lambda c}{1000}$$
$$k = \frac{142.3\Omega^{-1}\ cm^2\ mol^{-1} \times 0.02\ mol\ L^{-1}}{1000\ cm^3\ L^{-1}}$$
$$\therefore \quad k = 2.85 \times 10^{-3}\Omega^{-1}\ cm^{-1}$$

Question112

Which among the following is CORRECT formula for determination of cell constant?

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Options:

A. $\frac{l}{\mathrm{a}}=\frac{\mathrm{k}}{\mathrm{R}}$

B. $\frac{l}{\mathrm{a}}=\mathrm{k} \cdot \mathrm{R}$

C. $\frac{l}{\mathrm{a}}=\frac{\mathrm{R}}{\mathrm{k}}$

D. $\frac{l}{\mathrm{a}}=\frac{1}{\mathrm{R}}$

Answer: B

Solution:

$$\begin{aligned} & \mathrm{k}=\mathrm{G} \frac{l}{\mathrm{a}}=\frac{1}{\mathrm{R}} \frac{l}{\mathrm{a}} \\ \therefore \quad & \text{Cell constant }=\frac{l}{\mathrm{a}}=\mathrm{k} \cdot \mathrm{R}\end{aligned}$$

Question113

What happens when solution of an electrolyte is diluted?

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Options:

A. Both $\wedge$ and k increases

B. Both $\wedge$ and k decreases

C. $\wedge$ increases and k decreases

D. $\wedge$ decreases and k increases

Answer: C

Solution:

The molar conductivity ($\wedge$) of an electrolyte increases on dilution while conductivity (k) decreases on dilution.

Question114

Calculate E°_{cell} in which following reaction occurs. $\mathrm{Mg}_{(\mathrm{s})}+2\mathrm{Ag}^{+}_{(\mathrm{IM})} \rightarrow \mathrm{Mg}^{+}_{(1\mathrm{M})}+2\mathrm{Ag}_{(\mathrm{s})}$ if $\mathrm{E}^{\circ}_{\mathrm{Ag}}=0.8\ \mathrm{V}$ and $\mathrm{E}^{\circ}_{\mathrm{Mg}}=-2.37\ \mathrm{V}$

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Options:

A. -3.17 V

B. 3.17 V

C. -1.57 V

D. 1.57 V

Answer: B

Solution:

For the given cell reaction, anode is Mg and cathode is Ag.

$$\begin{aligned} E^{o}_{\text{edl}} &= E^{\circ}_{\text{cuthode}} - E^{\circ}_{\text{unode}} \\ &= 0.8-(-2.37) \\ &= 3.17\ \mathrm{V} \end{aligned}$$

Question115

Calculate the conductivity of $0.02 \mathrm{~M}$ electrolyte solution if its molar conductivity $407.2 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$?

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Options:

A. $8.144 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}$

B. $4.072 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}$

C. $7.15 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}$

D. $6.055 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}$

Answer: A

Solution:

$$
\begin{aligned}
\Lambda_{\mathrm{m}} & =\frac{1000 \mathrm{k}}{\mathrm{c}} \\
\mathrm{k} & =\frac{\Lambda_{\mathrm{m}} \times \mathrm{c}}{1000}=\frac{407.2 \times 0.02}{1000} \\
& =8.144 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}
\end{aligned}
$$

Question116

Conductivity of a solution is $1.26 \times 10^{-2} \Omega^{-1} \mathrm{~cm}^{-1}$ Calculate molar conductivity for $0.01 \mathrm{~M}$ solution.

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Options:

A. $1.26 \times 10^3 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

B. $2.52 \times 10^{3} \Omega^{-1} \mathrm{~cm}^{2} \mathrm{~mol}^{-1}$

C. $4.82 \times 10^{3} \Omega^{-1} \mathrm{~cm}^{2} \mathrm{~mol}^{-1}$

D. $6.30 \times 10^{3} \Omega^{-1} \mathrm{~cm}^{2} \mathrm{~mol}^{-1}$

Answer: A

Solution:

$$\begin{aligned} \Lambda &= \frac{1000\mathrm{k}}{\mathrm{c}} \\ &= \frac{1000 \mathrm{~cm}^{3} \mathrm{~L}^{-1} \times 1.26 \times 10^{-2} \Omega^{-1} \mathrm{~cm}^{-1}}{0.01 \mathrm{~mol} \mathrm{~L}^{-1}} \\ &= 1.26 \times 10^{3} \Omega^{-1} \mathrm{~cm}^{2} \mathrm{~mol}^{-1} \end{aligned}$$

Question117

Which of the following is NOT a difficulty in setting SHE?

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Options:

A. To obtain pure hydrogen gas

B. To obtain dry hydrogen gas

C. To maintain exactly 1 atm pressure

D. To bring the reaction in reverse direction

Answer: D

Solution:

In standard hydrogen electrode (SHE), a platinum plate coated with platinum black is used as electrode. The platinum black is capable of adsorbing large quantities of H_2 gas. This allows the change from gaseous to ionic form and the reverse process to occur.

Question118

What mass of Mg **is produced during electrolysis of molten** $\mathrm{MgCl_2}$ **by passing** 2 amp **current for 482.5 second?**

(Molar mass $\mathrm{Mg} = 24\ \mathrm{g\ mol^{-1}}$**)**

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Options:

A. $0.12\ \mathrm{g}$

B. $0.24\ \mathrm{g}$

C. $1.2\ \mathrm{g}$

D. $0.4\ \mathrm{g}$

Answer: A

Solution:

$$\mathrm{Mg^{2+}_{(s)} + 2e^- \longrightarrow Mg_{(s)}}$$

$$\text{Mole ratio } = \frac{1 \text{ mol}}{2 \text{ mole}^-}$$

$$\mathrm{W} = \frac{\mathrm{I(A)} \times \mathrm{t(s)}}{96500\,(\mathrm{C/mole^-})} \times \text{mole ratio} \times \text{molar mass}$$

$$\mathrm{W} = \frac{2 \times 482.5}{96500\,(\mathrm{C/mole^-})} \times \frac{1 \text{ mol}}{2 \text{ mole}^-} \times 24\ \mathrm{g\ mol^{-1}}$$

$$\mathrm{W} = 0.12\ \mathrm{g}$$

Question119

What is the conductivity of $0.05\ \mathrm{M\ BaCl_2}$ **solution if its molar conductivity is** $220\ \Omega^{-1}\ \mathrm{cm^2\ mol^{-1}}$ **?**

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Options:

A. $0.011\ \Omega^{-1}\ \mathrm{cm}^{-1}$

B. $0.022\ \Omega^{-1}\ \mathrm{cm}^{-1}$

C. $0.033\ \Omega^{-1}\ \mathrm{cm}^{-1}$

D. $0.044\ \Omega^{-1}\ \mathrm{cm}^{-1}$

Answer: A

Solution:

$$\begin{aligned} \Lambda &= \frac{1000\mathrm{k}}{\mathrm{c}} \\ \mathrm{k} &= \frac{\Lambda \mathrm{c}}{1000} \\ \mathrm{k} &= \frac{220\Omega^{-1}\ \mathrm{cm}^2\ \mathrm{mol}^{-1} \times 0.05\ \mathrm{mol\ L}^{-1}}{1000\ \mathrm{cm}^3\ \mathrm{L}^{-1}} \\ &= 0.011\Omega^{-1}\ \mathrm{cm}^{-1} \end{aligned}$$

Question120

A reaction, $\mathrm{Ni}_{(s)} + \mathrm{Cu}^{+}_{((M))} \rightarrow \mathrm{Ni}^{+}_{(IM)} + \mathrm{Cu}_{(s)}$ occurs in a cell. Calculate $\mathrm{E}^{\circ}_{\text{cell}}$ if $\mathrm{E}^{\circ}_{\mathrm{Cu}} = 0.337\ \mathrm{V}$ and $\mathrm{E}^{\circ}_{\mathrm{Ni}} = -0.257\ \mathrm{V}$

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Options:

A. $0.594\ \mathrm{V}$

B. $-0.594\ \mathrm{V}$

C. $-0.08\ \mathrm{V}$

D. $0.08\ \mathrm{V}$

Answer: A

Solution:

The standard cell potential is given by

$$\begin{aligned} E^\circ_{cell} &= E^\circ_{cathode} - E^\circ_{anode} \\ E^\circ_{cell} &= E^\circ_{Cu} - E^\circ_{Ni} \\ &= (0.337\ V) - (-0.257\ V) \\ &= 0.337\ V + 0.257\ V \\ &= 0.594\ V \end{aligned}$$

Question121

Calculate ΔG° for the cell:

$$Sn_{(s)} \left| Sn^{2+}_{(1M)} \right| \left| Ag^{+}_{(1M)} \right| Ag_{(s)} \text{ at } 25^\circ C \left(E^\circ_{cell} = 0.90\ V \right)$$

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Options:

A. -173.7 kJ

B. -225.3 kJ

C. -100.2 kJ

D. -290.8 kJ

Answer: A

Solution:

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

$$= -2 \times 96500 \times 0.90$$

$$= -173700\ J$$

$$= -173.7\ kJ$$

Question122

Calculate current in ampere required to deposit $4.8 \mathrm{~g} \mathrm{Cu}$ from it's salt solution in 30 minutes. [Molar mass of $\mathrm{Cu}=63.5 \mathrm{~g} \mathrm{~mol}^{-1}$]

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Options:

A. 8.1 ampere

B. 6.4 ampere

C. 10.5 ampere

D. 12.3 ampere

Answer: A

Solution:

$$
\begin{aligned}
& \mathrm{Cu}_{(\mathrm{s})}^{2+}+2 \mathrm{e}^{-} \longrightarrow \mathrm{Cu}_{(\mathrm{s})} \\
& \text { Mole ratio }=\frac{1 \mathrm{~mol}}{2 \text { mole }^{-}} \\
& \mathrm{W}=\frac{\mathrm{I}(\mathrm{A}) \times \mathrm{t}(\mathrm{s})}{96500\left(\mathrm{C} / \text { mole }^{-}\right)} \times \text { mole ratio } \times \text { molar mass } \\
& 4.8 \mathrm{~g}=\frac{\mathrm{I}(\mathrm{A}) \times 30 \times 60}{96500\left(\mathrm{C} / \text { mole }^{-}\right)} \times \frac{1 \mathrm{~mol}}{2 \mathrm{~mol} \mathrm{e}} \times 63.5 \mathrm{~g} \mathrm{~mol}^{-1} \\
& \mathrm{I}(\mathrm{A})=\frac{4.8 \times 96500 \times 2}{63.5 \times 30 \times 60}=8.1 \mathrm{~A}
\end{aligned}
$$

Question123

Identify the gas produced due to reduction of NH_4^{+} ions at cathode during working of dry cell.

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Options:

A. Ammonia

B. Hydrogen

C. Hydrogen chloride

D. Chlorine

Answer: B

Solution:

In a dry cell, at cathode, NH_4^+ ions are reduced and hydrogen gas is produced.

$$2NH^+_{4(aq)} + 2e^- \longrightarrow 2NH_{3(aq)} + H_{2(g)}$$

Question124

How long will it take to produce 5.4 g of Ag from molten AgCl by passing 5 amp current?

(Molar mass Ag = 108 g mol^{-1})

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Options:

A. 1930 second

B. 193 second

C. 965 second

D. 9650 second

Answer: C

Solution:

$t=\frac{m \times 96500}{\text{mol. ratio}}=\frac{5.4 \times 96500 \times 1}{108 \times 5}=965$ second

----

Question125

Which of the following is NOT an example of secondary voltaic cell?

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Options:

A. Lead storage battery

B. Dry cell

C. Nickel-cadmium cell

D. Mercury cell

Answer: B

Solution:

Dry cell is an example of primary voltaic cell.

Question126

Calculate $\wedge_0$ of $\mathrm{CH_2ClCOOH}$ if $\wedge_0$ for HCl, KCl and $\mathrm{CH_2ClCOOK}$ are $4.2, 1.5$ and $1.1 \Omega^{-1}\ \mathrm{cm}^2\ \mathrm{mol}^{-1}$ respectively?

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Options:

A. $2.7 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

B. $3.8 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

C. $1.9 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

D. $4.2 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

Answer: B

Solution:

$$\begin{aligned} \wedge_0 & =\left(\wedge_{\mathrm{CH}_2 \mathrm{ClCOOK}}+\wedge_{\mathrm{HCl}}\right)-\wedge_{\mathrm{KCl}} \\ & =(1.1+4.2)-1.5 \\ & =3.8 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1} \end{aligned}$$

Question127

The electrical conductance of unit volume ($1 \mathrm{~cm}^3$) of solution is called as

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Options:

A. electrical resistance

B. resistivity

C. molar conductivity

D. conductivity

Answer: D

Solution:

The correct answer is Option D: conductivity.

The electrical conductance of a unit volume ($1\ cm^3$) of solution is referred to as conductivity. Conductivity is a measure of how well a material can conduct an electric current. It is the inverse of resistivity and is commonly used to describe the ability of solutions to conduct electricity.

Therefore, the electrical conductance of unit volume ($1\ cm^3$) of solution is called conductivity.

Question128

What is the conductivity of 0.02 M HCl solution if molar conductivity of the solution at 25°C is $412.3\ \Omega^{-1}\ cm^{-1}\ mol^{-1}$?

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Options:

A. $8.880 \times 10^{-3}\ \Omega^{-1}\ cm^{-1}$

B. $8.414 \times 10^{-3}\ \Omega^{-1}\ cm^{-1}$

C. $8.624 \times 10^{-3}\ \Omega^{-1}\ cm^{-1}$

D. $8.246 \times 10^{-3}\ \Omega^{-1}\ cm^{-1}$

Answer: D

Solution:

$$c = 0.02\mathrm{M}, \wedge = 412.3\Omega^{-1}\ \mathrm{cm}^2\ \mathrm{mol}^{-1}$$

$$\wedge = \frac{1000\mathrm{k}}{\mathrm{c}}$$

$$\therefore \mathrm{k} = \frac{\wedge \mathrm{c}}{1000}$$

$$= \frac{412.3\Omega^{-1}\ \mathrm{cm}^2\ \mathrm{mol}^{-1} \times 0.02\mathrm{molL}^{-1}}{1000\ \mathrm{cm}^3\ \mathrm{L}^{-1}}$$

$$= 8.246 \times 10^{-3}\Omega^{-1}\ \mathrm{cm}^{-1}$$

Question129

What is the charge required for the reduction of moles of Cu^{2+} to Cu?

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Options:

A. 2.89×10^5 C

B. 1.93×10^5 C

C. 9.65×10^5 C

D. 3.86×10^5 C

Answer: D

Solution:

The electrode reaction is

$$\begin{aligned} & Cu^{2+} + 2e^- \longrightarrow Cu \\ \therefore \quad & 2Cu^{2+} + 4e^- \longrightarrow 2Cu \end{aligned}$$

The quantity of charge required for reduction of 2 moles of $Cu^{2+} = 4$ F

$$\begin{aligned} & = 4 \times 96500C \\ & = 3.86 \times 10^5 C \end{aligned}$$

Question130

The resistance of $0.2M$ solution of an electrolyte is 30 ohm and conductivity is $1.2\ S\ m^{-1}$. What is the value of cell constant?

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Options:

A. $0.47\ cm^{-1}$

B. 0.1 cm^{-1}

C. 0.36 cm^{-1}

D. 0.2 cm^{-1}

Answer: C

Solution:

To find the cell constant, we can use the relationship between conductivity (κ), resistance (R), and the cell constant (G*). The formula to calculate the cell constant is given by:

$G^* = \kappa \times R$

Given data:

Resistance, $R = 30\,\Omega$

Conductivity, $\kappa = 1.2\,\mathrm{S\,m^{-1}}$

First, we substitute the given values into the formula:

$G^* = 1.2\,\mathrm{S\,m^{-1}} \times 30\,\Omega$

Simplifying the expression:

$G^* = 1.2 \times 30$

$G^* = 36\,\mathrm{m^{-1}}$

However, the units of cell constant are usually given in $\mathrm{cm^{-1}}$ rather than $\mathrm{m^{-1}}$. To convert from $\mathrm{m^{-1}}$ to $\mathrm{cm^{-1}}$, we use the conversion factor $1\,\mathrm{m} = 100\,\mathrm{cm}$:

$G^* = 36\,\mathrm{m^{-1}} \times \frac{1}{100}$

$G^* = 0.36\,\mathrm{cm^{-1}}$

Therefore, the cell constant is $0.36\,\mathrm{cm^{-1}}$. The correct option is:

Option C: $0.36\,\mathrm{cm^{-1}}$

Question131

Which from the following is the correct relationship between standard Gibbs energy change and standard cell potential?

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Options:

A. $-\Delta G^0=-nFE_{cell}^0$

B. $\Delta G^{\circ}=\frac{E_{cell}^{o}}{nF}$

C. $E_{cell}^{o}\quad=\Delta G^{\circ}\times nF$

D. $\Delta G^{o}=-nFE_{cell}^{o}$

Answer: D

Solution:

The correct relationship between the standard Gibbs energy change (ΔG°) and the standard cell potential (E_{cell}°) is given by the equation:

$\Delta G^{o}=-nFE_{cell}^{o}$

Here, ΔG° is the standard Gibbs energy change, E_{cell}° is the standard cell potential, n is the number of moles of electrons transferred in the reaction, and F is the Faraday constant (approximately 96485 C/mol). This equation reflects that the Gibbs free energy change is directly related to the electrical work done by the cell potential during the reaction.

Therefore, the correct option is:

Option D

$\Delta G^{o}=-nFE_{cell}^{o}$

Question132

The conductivity of $0.04\ M\ BaCl_2$ solution is $0.0112\Omega^{-1}\ cm^{-1}$ at $25^{\circ}C$. What is it's molar conductivity?

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Options:

A. $357.0\Omega^{-1}\ cm^2\ mol^{-1}$

B. $140.0\Omega^{-1}\ cm^2\ mol^{-1}$

C. $44.8 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

D. $280.0 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

Answer: D

Solution:

To solve this problem, we need to find the molar conductivity of the 0.04 M $BaCl_2$ solution given its conductivity. The formula for molar conductivity (Λ_m) is:

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

where:

Λ_m = molar conductivity ($\Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$)

κ = conductivity ($\Omega^{-1} \mathrm{~cm}^{-1}$)

c = concentration (in mol per cm^3)

Given:

$\kappa = 0.0112 \Omega^{-1} \mathrm{~cm}^{-1}$

$c = 0.04 \mathrm{~M} = 0.04 \mathrm{~mol} \mathrm{~L}^{-1}$

First, convert concentration from $\mathrm{mol} \mathrm{~L}^{-1}$ to $\mathrm{mol} \mathrm{~cm}^{-3}$:

$0.04 \mathrm{~mol} \mathrm{~L}^{-1} = 0.04 \mathrm{~mol} \mathrm{~dm}^{-3} = 0.04 \times 10^{-3} \mathrm{~mol} \mathrm{~cm}^{-3}$

So, we have:

$c = 0.04 \times 10^{-3} \mathrm{~mol} \mathrm{~cm}^{-3}$

Now, plug these values into the formula:

$\Lambda_m = \frac{0.0112}{0.04 \times 10^{-3}}$

Calculate this expression:

$\Lambda_m = \frac{0.0112}{0.00004} = 280 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

Therefore, the correct answer is:

Option D: $280.0 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

Question133

The conductivity of $0.012 \mathrm{~M}$ NaBr solution is $2.67 \times 10^{-4} \mathrm{~S} \mathrm{~cm}^{-1}$. What is it's molar conductivity?

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Options:

A. $26.7 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

B. $32.04 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

C. $12.2 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

D. $22.2 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

Answer: D

Solution:

$$
\begin{aligned}
& \mathrm{C}=0.012 \mathrm{M}, \mathrm{k}=2.67 \times 10^{-4} \mathrm{~S} \mathrm{~cm}^{-1} \\
& \quad \wedge=\frac{1000 \mathrm{k}}{\mathrm{c}} \\
& \quad=\frac{1000 \mathrm{~cm}^3 \mathrm{~L}^{-1} \times 2.67 \times 10^{-4} \mathrm{~S} \mathrm{~cm}^{-1}}{0.012 \mathrm{~mol} \mathrm{~L}^{-1}} \\
& \therefore \wedge=22.25 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}
\end{aligned}
$$

Question134

During the electrolysis of fused NaCl, the product obtained at anode is

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Options:

A. $\mathrm{Na}_{(\mathrm{S})}$

B. $\mathrm{Cl}_{2\,(8)}$

C. $\mathrm{O}_{2\,(\mathrm{B})}$

D. $Na_{(l)}$

Answer: B

Solution:

The correct answer is **Option B: $Cl_{2(g)}$**.

Here's why:

During the electrolysis of molten NaCl, the following reactions occur:

At the anode (positive electrode):

The chloride ions (Cl^-) are oxidized to chlorine gas (Cl_2):

$2Cl^- \rightarrow Cl_2 + 2e^-$

At the cathode (negative electrode):

The sodium ions (Na^+) are reduced to sodium metal (Na):

$Na^+ + e^- \rightarrow Na$

Therefore, the product obtained at the anode is chlorine gas (Cl_2).

Let's analyze why the other options are incorrect:

Option A: $Na_{(s)}$ - Sodium metal is produced at the cathode, not the anode.

Option C: $O_{2(g)}$ - Oxygen gas is not produced during the electrolysis of molten NaCl. This would occur if water was present in the electrolyte.

Option D: $Na_{(l)}$ - Sodium is produced as a solid at the cathode, not a liquid.

Question135

What is the weight of Al deposited at cathode when 1 ampere current is passed through molten AlCl_3 for 9650 seconds? (At mass of $\mathrm{Al} = 27$)

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Options:

A. 3.0 g

B. 9.0 g

C. 13.6 g

D. 0.9 g

Answer: D

Solution:

$Al^{3+} + 3e^{-} \longrightarrow Al$

$I = 1A, t = 9650\ s$

$$\text{Mass of product } = \frac{I(A)\times t(s)}{96500(C/mole^{-})} \times \text{ mole ratio } \times \text{ Molar mass of product}$$

$$\begin{aligned}\therefore \text{ Weight of Al deposited } &= \frac{1\times 9650}{96500} \times \frac{1}{3} \times 27 \\ &= 0.9\ g\end{aligned}$$

Question136

How many electrons flow through the wire if a current of 1.5 ampere flow through it for 3 hours?

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Options:

A. 1.60×10^{19}

B. 1.01×10^{23}

C. 1.01×10^{19}

D. 1.60×10^{23}

Answer: B

Solution:

$I = 1.5\ A, t = 3hr = 3 \times 60\ min \times 60\ s = 10800\ s$
$Q = It = 1.5 \times 10800 = 16200C$

Now, $1.602 \times 10^{-19} \mathrm{C} = 1 \mathrm{e}^{-}$

$$\therefore \quad 16200 \mathrm{C} = \frac{1 \mathrm{e}^{-} \times 16200 \mathrm{C}}{1.6 \times 10^{-19} \mathrm{C}}$$
$$= 10125 \times 10^{19} \mathrm{e}^{-}$$
$$= 1.01 \times 10^{23} \mathrm{e}^{-}$$

Question137

The conductivity of $0.3 \mathrm{M}$ solution of KCl at $298 \mathrm{~K}$ is $0.0627 \mathrm{~S} \mathrm{~cm}^{-1}$. What is it's molar conductivity?

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Options:

A. $104 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

B. $188 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

C. $209 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

D. $109 \mathrm{~S} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

Answer: C

Solution:

Molar conductivity (Λ_m) is given by the formula:

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

where:

κ = conductivity of the solution (in $\mathrm{S} \mathrm{~cm}^{-1}$) and

c = concentration of the solution (in $\mathrm{mol} \mathrm{~L}^{-1}$).

Given, the conductivity κ is $0.0627 \mathrm{~S} \mathrm{~cm}^{-1}$ and the concentration c is $0.3 \mathrm{M}$ (or $0.3 \mathrm{~mol} \mathrm{~L}^{-1}$).

Substituting these values into the formula:

$\Lambda_m = \frac{0.0627 \mathrm{~S} \mathrm{~cm}^{-1}}{0.3 \mathrm{~mol} \mathrm{~L}^{-1}}$

$\Lambda_m = 0.209 \text{ S cm}^2 \text{ mol}^{-1}$

Hence, the molar conductivity of the 0.3M solution of KCl at 298 K is 209 S cm^2 mol^{-1}.

The correct answer is Option C 209 S cm^2 mol^{-1}.

Question138

In the cell represented as $\mathrm{Ni}_{(s)}\left|\mathrm{Ni}_{(\mathrm{IM})}^{2\oplus}\right|\left|\mathrm{Ag}_{(\mathrm{IM})}^{\oplus}\right|\mathrm{Ag}_{(s)}$, the reducing agent is

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Options:

A. Ag

B. $\mathrm{Ag}^{\oplus}$

C. Ni

D. $\mathrm{Ni}^{2\oplus}$

Answer: C

Solution:

In an electrochemical cell notation, the substances on the left side of the double vertical line represent the anode (where oxidation occurs), and those on the right side represent the cathode (where reduction occurs). The cell notation $\mathrm{Ni}_{(s)}\left|\mathrm{Ni}_{(\mathrm{IM})}^{2\oplus}\right|\left|\mathrm{Ag}_{(\mathrm{IM})}^{\oplus}\right|\mathrm{Ag}_{(s)}$ indicates that nickel metal (Ni) is at the anode and is oxidized to $\mathrm{Ni}^{2\oplus}$ ions, while $\mathrm{Ag}^{\oplus}$ ions are reduced to silver metal (Ag) at the cathode.

Oxidation involves losing electrons, meaning the substance that is oxidized is the reducing agent. Here, nickel (Ni) is the substance that loses electrons:

$\mathrm{Ni} \rightarrow \mathrm{Ni}^{2\oplus} + 2e^-$

Since nickel (Ni) is losing electrons, it is the reducing agent in this cell.

Thus, the correct answer is:

Option C

Ni

Question139

How many Faraday of electricity is required to deposit $0.8 \mathrm{~g}$ of calcium at cathode by the electrolysis of CaCl_2 ?

(Molar mass of $\mathrm{Ca}=40 \mathrm{~g} \mathrm{~mol}^{-1}$)

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Options:

A. 4 F

B. 0.04 F

C. 2.5 F

D. 2 F

Answer: B

Solution:

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

40 g of Ca requires 2 F of electricity

$\therefore$ 0.8 g of Ca requires $=\frac{2 \mathrm{~F} \times 0.8 \mathrm{~g}}{40 \mathrm{~g}}=0.04$ F of electricity.

Question140

What is the conductivity of $0.02 \mathrm{~M} \mathrm{KCl}$ solution if cell constant is $1.29 \mathrm{~cm}^{-1}$ with resistance 645 Ω ?

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Options:

A. $5.0 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}$

B. $2.0 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}$

C. $8.3 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}$

D. $2.5 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}$

Answer: B

Solution:

$$
\begin{aligned}
& \mathrm{C}=0.02 \mathrm{M}, \text { cell constant }=1.29 \mathrm{~cm}^{-1} \\
& \mathrm{R}_{\text {solution }}=645 \Omega, \mathrm{k}=? \\
& \mathrm{k}=\frac{\text { cell constant }}{\mathrm{R}_{\text {solution }}}=\frac{1.29 \mathrm{~cm}^{-1}}{645 \Omega} \\
& \therefore \mathrm{k}=2.0 \times 10^{-3} \Omega^{-1} \mathrm{~cm}^{-1}
\end{aligned}
$$

Question141

Electrical conductance due to all the ions in $1 \mathrm{~cm}^{3}$ of given solution is called as

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Options:

A. Molar conductivity

B. Resistivity

C. Conductivity

D. Electrical conductance

Answer: C

Solution:

The correct answer is Option C: Conductivity.

Electrical conductivity is a measure of the ability of a solution to conduct electric current. It is determined by the presence of ions in the solution, as these ions act as charge carriers. The definition you've provided, which refers to the conductance due to all the ions in 1 cm^3 of a given solution, precisely matches the definition of conductivity. Conductivity is usually denoted by the symbol σ and is the reciprocal of resistivity (ρ).

To clarify further:

- **Molar conductivity** (Option A) refers to the conductivity of a solution containing one mole of electrolyte dissolved in a specified volume of solvent, usually expressed in units of $S \cdot cm^2/mol$. It emphasizes the conductive contribution of the electrolyte on a per-mole basis.

- **Resistivity** (Option B) is the reciprocal of conductivity. It measures the resistance of a solution to the flow of an electric current, typically denoted by ρ and measured in ohm-meters ($\Omega \cdot m$).

- **Conductivity** (Option C) is the correct answer as per the definition provided in the question. It's a bulk property of the solution and measures how easily electric current can flow through it, generally given in Siemens per meter (S/m).

- **Electrical conductance** (Option D) is closely related to conductivity but is not the exact answer to the definition provided. Conductance measures the ease with which an electric current passes through an object (not specifically in 1 cm^3), denoted by G and measured in Siemens (S).

Therefore, conductivity most accurately fits the description of the electrical conductance due to all ions in 1 cm^3 of a given solution.

Question142

What is the resistance of $0.01 \mathrm{~M~KCl}$ solution if its conductivity is $200 \mathrm{~ohm}^{-1} \mathrm{cn}^{-1}$ and cell constant is $1 \mathrm{~cm}^{-1}$?

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Options:

A. 1×10^{-3} ohm

B. 5×10^{-3} ohm

C. 4×10^{-3} ohm

D. 2×10^{-3} ohm

Answer: B

Solution:

To calculate the resistance of the 0.01 M KCl solution, we can use the relationship between conductivity (κ), cell constant (l/A), and resistance (R). The formula is:

$R = \frac{l}{A\kappa}$

Here:

- κ is the conductivity of the solution.
- $\frac{l}{A}$ is the cell constant.

Given:

- $\kappa = 200 \text{ ohm}^{-1} \text{ cm}^{-1}$
- Cell constant $\frac{l}{A} = 1 \text{ cm}^{-1}$

Now, substituting the given values into the formula:

$R = \frac{1 \text{ cm}^{-1}}{200 \text{ ohm}^{-1} \text{ cm}^{-1}}$

$R = \frac{1}{200} \text{ ohm}$

$R = 0.005 \text{ ohm}$

Therefore, the resistance of the 0.01 M KCl solution is 5×10^{-3} ohm.

The correct option is:

Option B 5×10^{-3} ohm

Question143

Molar conductivity of 0.04 M $BaCl_2$ solution is 230 $\Omega^{-1} \text{ cm}^2 \text{ mol}^{-1}$ at 27°C. What is it's conductivity?

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Options:

A. $2.3 \times 10^{-3} \ \Omega^{-1} \text{ cm}^{-1}$

B. $9.2 \times 10^{-3} \ \Omega^{-1} \text{ cm}^{-1}$

C. $6.9 \times 10^{-3} \ \Omega^{-1} \text{ cm}^{-1}$

D. $4.6 \times 10^{-3} \ \Omega^{-1} \text{ cm}^{-1}$

Answer: B

Solution:

To determine the conductivity (κ) of the 0.04 M $BaCl_2$ solution, we use the relationship between molar conductivity (Λ_m) and conductivity (κ). The formula is:

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

Where:

- Λ_m is the molar conductivity
- κ is the conductivity
- c is the concentration of the solution in mol/L

Given:

- $\Lambda_m = 230\ \Omega^{-1}\ \mathrm{cm}^2\ \mathrm{mol}^{-1}$
- c = 0.04 M

We need to find κ. Rearrange the formula to solve for κ:

$\kappa = \Lambda_m \times c$

Substitute the given values:

$\kappa = 230\ \Omega^{-1}\ \mathrm{cm}^2\ \mathrm{mol}^{-1} \times 0.04\ \mathrm{mol\ L}^{-1}$

$\kappa = 9.2\ \Omega^{-1}\ \mathrm{cm}^{-1}$

So the correct answer is:

Option B

$9.2 \times 10^{-3}\ \Omega^{-1}\ \mathrm{cm}^{-1}$

Question144

Which from following electrolytes, molar conductivity is determined using Kohlrausch theory?

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Options:

A. KCl

B. Na_2SO_4

C. CH_3COOH

D. HCl

Answer: C

Solution:

The molar conductivity of strong electrolytes such as KCl and HCl is typically determined directly, but for weak electrolytes like CH_3COOH (acetic acid), the molar conductivity at infinite dilution cannot be measured directly due to the incomplete dissociation of the weak electrolyte. Instead, we use Kohlrausch's Law of Independent Migration of Ions to estimate it.

Kohlrausch's Law states that the limiting molar conductivity of an electrolyte can be expressed as the sum of the individual ionic contributions. Mathematically, this can be written as:

$$\Lambda_m^0 = \lambda_+^0 + \lambda_-^0$$

Where:

- Λ_m^0 is the limiting molar conductivity.
- λ_+^0 is the limiting molar conductivity of the cation.
- λ_-^0 is the limiting molar conductivity of the anion.

Using this theory, we can calculate the molar conductivity of weak electrolytes by combining the molar conductivities of the ions from strong electrolytes that exhibit complete dissociation.

Therefore, for the given options, Kohlrausch's theory is primarily used to determine the molar conductivity of:

Option C: CH_3COOH

Question145

Which among the following statements is true for Galvanic cell?

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Options:

A. The anode is negative and cathode is positive.

B. Electrical energy is converted into chemical energy.

C. The anode is positive and cathode is negative.

D. Oxidation takes place at positive electrode and reduction takes place at negative electrode.

Answer: A

Solution:

In a Galvanic cell :

- 1. Oxidation occurs at the anode, which is the negative electrode.
- 2. Reduction occurs at the cathode, which is the positive electrode.
- 3. The cell operates by converting chemical energy into electrical energy.

Correct answer :

Option A : The anode is negative and cathode is positive.

Question146

What will be the concentration of solution of electrolyte if it's molar conductivity and conductivity are respectively 230 Ω^{-1} cm ^{2} mol^{-1} and 0.0115 Ω^{-1} cm^{-1} at 298 K?

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Options:

A. 0.04 M

B. 0.03 M

C. 0.01 M

D. 0.05 M

Answer: D

Solution:

$$\wedge = 230 \Omega^{-1}\ \mathrm{cm}^2\ \mathrm{mol}^{-1}, \mathrm{k} = 0.0115 \Omega^{-1}\ \mathrm{cm}^{-1}$$

$$\wedge = \frac{1000\mathrm{k}}{\mathrm{c}}$$

$$\therefore \mathrm{c} = \frac{1000\mathrm{k}}{\hat{N}} = \frac{1000\ \mathrm{cm}^3\ \mathrm{L}^{-1} \times 0.0115 \Omega^{-1}\ \mathrm{cm}^{-1}}{230 \Omega^{-1}\ \mathrm{cm}^2\ \mathrm{mol}^{-1}}$$

$$= 0.05\ \mathrm{mol}\ \mathrm{L}^{-1}$$

Question147

What is the number of moles of electrons passed when current of 5 ampere is passed through a solution of $FeCl_3$ for 20 minutes?

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Options:

A. 6.25×10^{-2}

B. 1.56×10^{-2}

C. 3.12×10^{-2}

D. 4.25×10^{-2}

Answer: A

Solution:

$$\begin{aligned} I = 5\ \mathrm{A}, \mathrm{t} = 20\ \mathrm{min} = 20 \times 60 = 1200\ \mathrm{s} \\ Q = I \times t = 5\ \mathrm{A} \times 1200\ \mathrm{s} \\ = 6000\mathrm{C} \end{aligned}$$

Moles of electrons actually passed

$$\begin{aligned} &= \frac{\mathrm{Q(C)}}{965000\,(\mathrm{C/mole^-})} \\ &= \frac{6000}{96500} \\ &= 6.22 \times 10^{-2}\ \mathrm{mol\ e^-} \end{aligned}$$

Question148

Which among the following statements is true for conductivity?

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Options:

A. It is inversely proportional to resistivity.

B. It is inversely proportional to molar conductivity.

C. It is directly proportional to resistivity.

D. It is directly proportional to resistance.

Answer: A

Solution:

Conductivity is inversely proportional to resistivity i.e. $\mathrm{k}=\frac{1}{\rho}$

Question149

The molar conductivity of $0.4 \mathrm{~M} \mathrm{KCl}$ solution is $2.5 \times 10^{5} \Omega^{-1} \mathrm{~cm}^{2} \mathrm{~mol}^{-1}$. What is the resistivity of solution?

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Options:

A. 2.1×10^{2}

B. 2.5×10^{2}

C. 1×10^{-2}

D. 2.8×10^{-2}

Answer: C

Solution:

$\mathrm{c}=0.4 \mathrm{M}, \wedge=2.5 \times 10^{5} \Omega^{-1} \mathrm{~cm}^{2} \mathrm{~mol}^{-1}, \rho=\ ?$

(i) $\wedge = \frac{1000k}{C} \quad \therefore k = \frac{\wedge C}{1000} \quad \therefore k = \frac{2.5 \times 10^5 \times 0.4}{1000} = 100$

(ii) $k = \frac{1}{\rho} \quad \therefore \rho = \frac{1}{k} = \frac{1}{100} = 1 \times 10^{-2}$

Question150

What is the number of electrons passed through an electrolyte solution when 1 ampere current is passed for 16.1 minutes?

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Options:

A. 5.022×10^{24}

B. 3.011×10^{22}

C. 6.022×10^{21}

D. 2.022×10^{23}

Answer: C

Solution:

$I = 1\ A, t = 16.1\ min = 16.1 \times 60\ s$

$$\begin{aligned} \text{Moles of electrons actually passed } &= \frac{Q(C)}{96500\,(C/mole^{-}\cdot)} \\ &= \frac{I \times t}{96500} \\ &= \frac{1 \times 16.1 \times 60}{96500} \\ &= 0.01\ mol \end{aligned}$$

Now, $1\ mol = 6.022 \times 10^{23}$ electrons

$\therefore 0.01\ mol = 6.022 \times 10^{21}$ electrons

Question151

Which among the following aqueous salt solution is used in conductivity cell to determine cell constant?

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Options:

A. $AgNO_3$

B. $ZnSO_4$

C. KCl

D. $CuSO_4$

Answer: C

Solution:

The cell constant is determined using the 1 M, 0.1 M or 0.01 M KCl solutions.

Question152

What is the molar conductivity of $0.05 \mathrm{M}$ solution of sodium hydroxide, if it's conductivity is $0.0118 \mathrm{~S~cm}^{-1}$ at $298 \mathrm{~K}$?

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Options:

A. $236 \mathrm{~S~cm}^2 \mathrm{~mol}^{-1}$

B. $423 \mathrm{~S~cm}^2 \mathrm{~mol}^{-1}$

C. $354 \mathrm{~S~cm}^2 \mathrm{~mol}^{-1}$

D. $590 \mathrm{~S~cm}^2 \mathrm{~mol}^{-1}$

Answer: A

Solution:

$$\begin{aligned} &c = 0.05\mathrm{M}, k = 0.0118\ \mathrm{S\ cm^{-1}} \\ &\wedge = \frac{1000k}{c} \\ &= \frac{1000\ \mathrm{cm^3\ L^{-1}} \times 0.0118\ \mathrm{S\ cm^{-1}}}{0.05\ \mathrm{mol\ L^{-1}}} \\ &= 236\ \mathrm{S\ cm^2\ mol^{-1}} \end{aligned}$$

Question153

How many faraday of electricity is required to produce 10 g of calcium metal (molar mass = 40 g mol^{-1}) from calcium ions?

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Options:

A. 1.5 F

B. 2.0 F

C. 0.50 F

D. 1.0 F

Answer: C

Solution:

$Ca^{2+}_{(aq)} + 2e^{-} \longrightarrow Ca_{(s)}$

40 g of Ca requires 2 F of electricity

$\therefore$ 10 g of Ca $= \frac{2\times10}{40} = 0.5$ F of electricity

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Question154

The molar conductivity of 0.1 M $BaCl_2$ solution is 106 Ω^{-1} cm^2 mol^{-1} at 25°C. What is it's conductivity?

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Options:

A. $1.06 \times 10^{-2}\ \Omega^{-1}\ cm^{-1}$

B. $5.3 \times 10^{-3}\ \Omega^{-1}\ cm^{-1}$

C. $3.66 \times 10^{-3}\ \Omega^{-1}\ cm^{-1}$

D. $2.6 \times 10^{-2}\ \Omega^{-1}\ cm^{-1}$

Answer: A

Solution:

$\wedge = \frac{1000k}{c} \quad \therefore k = \frac{\wedge c}{1000}$

Now, $\wedge = 106\Omega^{-1}\ cm^2\ mol^{-1}, c = 0.1\ mol\ L^{-1}$

$$\therefore k = \frac{106\Omega^{-1}\ cm^2\ mol^{-1} \times 0.1\ mol\ L^{-1}}{1000\ cm^3\ L^{-1}}$$

$\therefore k = 1.06 \times 10^{-2}\Omega^{-1}\ cm^{-1}$

Question155

Which from following is NOT correct regarding electrolysis?

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Options:

A. It helps in refining of metals.

B. Conversion of electrical energy into chemical energy takes place.

C. It is useful in electroplating.

D. Metal is deposited at anode.

Answer: D

Solution:

The correct answer is :

Option D : Metal is deposited at anode.

This statement is NOT correct regarding electrolysis. In electrolysis, metal is deposited at the cathode, not the anode. The anode is where oxidation occurs, and the cathode is where reduction occurs.

Question156

A conductivity cell shows resistance of 600 ohm. If conductivity of 0.01 M KCl is 0.0015 Ω^{-1} cm^{-1}, what is cell constant?

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Options:

A. 0.60 cm^{-1}

B. 0.45 cm^{-1}

C. 0.90 cm^{-1}

D. 0.75 cm^{-1}

Answer: C

Solution:

$\mathrm{R}=600 \Omega, \mathrm{k}_{\mathrm{KCl}}=0.0015\ \Omega^{-} \mathrm{cm}^{-1}, \mathrm{c}=0.01\ \mathrm{M}$

$\text{Cell constant } =\mathrm{k}_{\mathrm{KCl}} \times \mathrm{R}_{\text{solution}} \quad =0.0015\ \Omega^{-1}\ \mathrm{cm}^{-1} \times 600 \Omega=0.90\ \mathrm{cm}^{-1}$

Question157

How many Faradays of electricity is required to produce 4.8 g of Mg at cathode in the electrolysis of molten $\mathrm{MgCl_2}$?(Molar mass of $\mathrm{Mg} = 24\ \mathrm{g/mol}$)

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Options:

A. 4 F

B. 0.4 F

C. 1 F

D. 10 F

Answer: B

Solution:

Electrode reaction:

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

To produce 24 g of Mg from molten $\mathrm{MgCl_2}$, electricity required $= 2$ F

To produce 4.8 g Mg from $\mathrm{MgCl_2}$, electricity required

$= \frac{(2F)\times 4.8\ \mathrm{g}}{24\ \mathrm{g}} = 0.4\ \mathrm{F}$

Question158

What is the SI unit for electrochemical equivalent?

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Options:

A. $\mathrm{Js^{-1}}$

B. JC^{-1}

C. kgC

D. kgC^{-1}

Answer: D

Solution:

The SI unit for electro-chemical equivalent is kgC^{-1}.

Electrochemical equivalent

$(Z)=\frac{\text{mass}}{1 \text{ coulomb of electrical charge}}=\frac{\mathrm{kg}}{\mathrm{C}}$

The electrochemical equivalent, sometimes abbreviated E_q or Z of a chemical element is the mass of that element transported by 1 coulomb of electrical charge. It is measured with a voltameter.

Question159

Number of electrons involved in the reaction, when 1 Faraday of electricity is passed through an electrolytic solution is

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Options:

A. 12×10^{46}

B. 96500

C. 6.022×10^{23}

D. 8×10^{16}

Answer: C

Solution:

Given,

Quantity of charge passed $(Q)=1 \mathrm{~F}$

$= 1 \times 96500\text{C} = 96500\text{C}$

Value of electron charge $(e^-) = 1.60217 \times 10^{-19}\text{C}$

We know that, $Q = ne^-$

$\Rightarrow \quad n = \frac{Q}{e^-} = \frac{96500\text{C}}{1.60217 \times 10^{-19}\text{C}} = 6.022 \times 10^{23}$

Question160

A solution of CuSO_4 is electrolysed using a current of 1.5 amperes for 10 minutes. What mass of Cu is deposited at cathode? [Atomic mass of $\text{Cu} = 63.7$]

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Options:

A. 0.637 g

B. 0.297 g

C. 0.150 g

D. 0.395 g

Answer: B

Solution:

The mass of copper (Cu) deposited at the cathode can be calculated using Faraday's laws of electrolysis. The formula to find the mass deposited is :

$m = \frac{Q \times M}{n \times F}$

where :

m is the mass of the substance deposited (in grams),

Q is the total electric charge passed through the solution (in coulombs),

M is the molar mass of the substance (in grams per mole),

n is the number of moles of electrons required to deposit one mole of the substance,

F is Faraday's constant (96500 C/mol).

First, calculate the total electric charge (Q) :

$Q = I \times t$

where :

I is the current (1.5 amperes),

t is the time (10 minutes = 600 seconds).

$Q = 1.5\,\mathrm{A} \times 600\,\mathrm{s} = 900\,\mathrm{C}$

For copper (Cu), the number of moles of electrons (n) required to deposit one mole is 2 (since the reaction is $\mathrm{Cu}^{2+} + 2e^{-} \rightarrow \mathrm{Cu}$).

Now, substitute the known values into the formula :

$m = \frac{900\text{ C } \times 63.7\text{ g/mol}}{2 \times 96500\text{ C/mol}}$

$m = \frac{57330}{193000}\,\mathrm{g}$

$m = 0.297\,\mathrm{g}$

Therefore, the mass of copper deposited at the cathode is 0.297 g.

Option B : 0.297 g

Question161

If resistivity of 0.8 M KCl solution is $2.5 \times 10^{-3} \Omega$ cm. Calculate molar conductivity solution?

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Options:

A. $2 \times 10^{5} \Omega^{-1}\,\mathrm{cm}^{2}\,\mathrm{mol}^{-1}$

B. $5 \times 10^{5} \Omega^{-1}\,\mathrm{cm}^{2}\,\mathrm{mol}^{-1}$

C. $4 \times 10^{5} \Omega^{-1}\,\mathrm{cm}^{2}\,\mathrm{mol}^{-1}$

D. $3 \times 10^{5} \Omega^{-1}\,\mathrm{cm}^{2}\,\mathrm{mol}^{-1}$

Answer: B

Solution:

Resistivity (ρ) $=2.5 \times 10^{-3} \Omega \mathrm{cm}$

$$M=0.8 \mathrm{~m}$$

We know that,

$$\begin{aligned}
k & =\frac{1}{\rho}=\frac{1}{2.5 \times 10^{-3}} \\
& =4 \times 10^2 \Omega \mathrm{cm}^{-1}
\end{aligned}$$

Also,

$$\text{molar conductivity } (\lambda_m)=k \times \frac{1000}{M}$$

$$\begin{aligned}
& =\frac{4 \times 10^2 \times 1000}{0.8} \\
\lambda_m & =5 \times 10^5 \Omega \mathrm{cm}^2 \mathrm{~mol}^{-1}
\end{aligned}$$

Question162

How many electrons are involved in the reaction, when $0.40 \mathrm{~F}$ of electricity is passed through an electrolytic solution?

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Options:

A. 6.642×10^{25}

B. 1.505×10^{24}

C. 6.022×10^{23}

D. 2.4088×10^{23}

Answer: D

Solution:

Given,

Quantity of charge (Q) passed $=0.40 \mathrm{~F}$

$$\begin{aligned}
& =0.40 \times 96500 \mathrm{C} \\
& =38600 \mathrm{C}
\end{aligned}$$

We know that, $Q=n e^{-}$

where, $n=$ number of electrons

$e^{-}=$ charge on electron $\left(1.6020 \times 10^{-19} \mathrm{C}\right)$

$\therefore n=\frac{Q}{e^{-}}=\frac{38600 \mathrm{C}}{1.6020 \times 10^{-19} \mathrm{C}}=2.4088 \times 10^{23}$

Question163

What will be the concentration of NaCl solution, if the molar conductivity and conductivity of NaCl solution is $124.3 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$ and $1.243 \times 10^{-4} \Omega^{-1} \mathrm{~cm}^2$ respectively?

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Options:

A. $0.1 \mathrm{~mol} \mathrm{~L}^{-1}$

B. $0.001 \mathrm{~mol} \mathrm{~L}^{-1}$

C. $0.01 \mathrm{~mol} \mathrm{~L}^{-1}$

D. $0.02 \mathrm{~mol} \mathrm{~L}^{-1}$

Answer: B

Solution:

Given,

Molar conductivity $\left(\lambda_{\mathrm{m}}\right)=124.3 \Omega^{-1} \mathrm{~cm}^2 \mathrm{~mol}^{-1}$

Conductivity $(\kappa)=1.243 \times 10^{-4} \Omega^{-1} \mathrm{~cm}^2$

We know that,

Molar conductivity $\left(\lambda_m\right)=\frac{\kappa \times 1000}{M}$

$\Rightarrow \quad M=\frac{\kappa \times 1000}{\lambda m}$

$$=\frac{1.243 \times 10^{-4} \times 1000}{124.3}$$
$$=0.001 \mathrm{~mol} \mathrm{~L}^{-1}$$

Question164

Which among the following is correct for electrolysis of brine solution?

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Options:

A. O_2 gas is liberated at cathode

B. Sodium metal is collected at anode

C. H_2 gas is liberated at cathode

D. Cl_2 gas is liberated at cathode

Answer: C

Solution:

During the electrolysis of brine (a concentrated solution of sodium chloride, NaCl), the following chemical reactions occur:

At the Cathode (Negative Electrode):

Reduction occurs, and water is reduced to form hydrogen gas:

$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

As a result, H_2 gas is liberated at the cathode.

At the Anode (Positive Electrode):

Oxidation occurs, and chloride ions are oxidized to form chlorine gas:

$2Cl^- - 2e^- \rightarrow Cl_2$

Thus, Cl_2 gas is liberated at the anode, not the cathode.

Therefore, the correct option is:

Option C: H_2 gas is liberated at the cathode.

Question165

In the electrolysis of aqueous sodium chloride with inert electrodes the products obtained at anode and cathode respectively are

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Options:

A. Cl_2 and Na

B. O_2 and Na

C. Cl_2 and H_2

D. Na and Cl_2

Answer: C

Solution:

In the electrolysis of aqueous sodium chloride with inert electrodes, the products obtained at anode and cathode respectively are Cl_2 and H_2.The solution contains four ions Na^+, Cl, H^+, OH. Following reaction takes place:

At cathode

$$\begin{aligned} & \mathrm{Na}^+(aq) + e^- \longrightarrow \mathrm{Na}(s); \\ & E_{\mathrm{Na}^+/\mathrm{Na}} = -2.71 \mathrm{~V} \\ & 2\mathrm{H}_2\mathrm{O}(\mathrm{l}) + 2e^- \longrightarrow \mathrm{H}_2(g) + 2\overline{\mathrm{O}}\mathrm{H}(aq); \\ & E^0_{\mathrm{H}_2\mathrm{O}/\mathrm{H}_2} = -0.83 \mathrm{~V} \end{aligned}$$

At anode

$$2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-; E^o_{Cl^-/Cl_2} = -1.36V$$

$$2H_2O_2(l) \longrightarrow O(g) + 4H^+(aq) + 4e^-; E^o_{H_2O/O_2} = -1.23V$$

$E^\circ_{\mathrm{H_2O/H_2}} > E^\circ_{\mathrm{Na^+/Na^-}}$. So, H_2 is formed at cathode. Whereas $E^\circ_{\mathrm{H_2O/O_2}} > \mathrm{E}^\circ_{\mathrm{Cl^-/Cl_2}}$. So, Cl_2 is formed at anode.

Question166

Calculate E.M.F. of following cell at 298 K
$\mathrm{Zn}(s)\left|\mathrm{ZnSO}_4(0.01\mathrm{M})\right|\left|\mathrm{CuSO}_4(1.0\mathrm{M})\right|\mathrm{Cu(s)}$ **if** E° **cell** $= 2.0$ V

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Options:

A. 2.0 V

B. 2.0592 V

C. 2.0296 V

D. 1.0508 V

Answer: B

Solution:

oxidation $\mathrm{Zn}(s) \longrightarrow \mathrm{Zn}^{2+}(\mathrm{aq})[0.01\mathrm{M}] + 2e^{-}$

Reduction $\mathrm{Cu}^{2+}(\mathrm{aq})[1.0]\mathrm{M} \longrightarrow \mathrm{Cu(s)}$

Using Nernst equation,

$$\begin{aligned} E_{\text{cell}} & = E_{\text{cell}}^{\circ} \quad - \frac{0.059}{2} \log\left[\frac{1}{100}\right] \\ & = 2 - \frac{0.59}{2} \times -2 = 2 + 0.059 \\ & = 2.059 \text{ V} \end{aligned}$$

Question167

The SI unit of electrochemical equivalent is

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Options:

A. $\mathrm{J\ s}^{-1}$

B. kg C^{-1}

C. kg ms^{-2}

D. kg m^{-1} s^{-2}

Answer: B

Solution:

The electrochemical equivalent of a substance is the mass liberated in electrolysis by 1 coulomb. It can be calculated by assuming Faraday's second law and using the experimental fact that the equivalent weight in grams of all substances is liberated by 96500 C .

Electrochemical equivalent $= \frac{\text{Equivalent weight}}{96500}$ gm per coulomb. Its unit is kg C^{-1}.

Question168

The molar conductivities at infinite dilution for sodium acetate, HCl and NaCl are $91 \mathrm{~S~cm}^2 \mathrm{~mol}^{-1}$, $425.9 \mathrm{~S~cm}^2 \mathrm{~mol}^{-1}$ and $12.6.4 \mathrm{~S~cm}^2 \mathrm{~mol}^{-1}$ respectively. The molar conductivity of acetic acid at infinite dilution is

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Options:

A. $390.5 \mathrm{~S~cm}^2 \mathrm{~mol}^{-1}$

B. $530.9 \mathrm{~S~cm}^2 \mathrm{~mol}^{-1}$

C. $300.5 \mathrm{~S~cm}^2 \mathrm{~mol}^{-1}$

D. $930.5 \mathrm{~S~cm}^2 \mathrm{~mol}^{-1}$

Answer: A

Solution:

Given,

$$\begin{aligned}\Lambda^{\circ}_{\mathrm{m(NaAC)}} &= 91 \mathrm{~s~cm}^2 \mathrm{~mol}^{-1}\\ \Lambda^{\circ}_{\mathrm{m(HCl)}} &= 425.9 \mathrm{~s~cm}^2 \mathrm{~mol}^{-1}\\ \Lambda^{\circ}_{\mathrm{m(NaCl)}} &= 126.4 \mathrm{~S~cm}^2 \mathrm{~mol}^{-1}\end{aligned}$$

Using Kohlrausch law, to find the molar conductivity of acetic acid.

$$\begin{aligned}\Lambda^{\circ}_{m(\mathrm{HAC})} &= \lambda^{\circ}_{\mathrm{H}^+} + \lambda^{\circ}_{\mathrm{AC}^-}\\ &= \lambda^{\circ}_{\mathrm{H}^+} + \lambda^{\circ}_{\mathrm{Cl}^-} + \lambda^{\circ}_{\mathrm{AC}^-} + \lambda^{\circ}_{\mathrm{Na}^+} - \lambda^{\circ}_{\mathrm{Cl}^-} - \lambda^{\circ}_{\mathrm{Na}^+}\\ &= \Lambda^{\circ}_{\mathrm{m(HCl)}} + \Lambda^{\circ}_{\mathrm{m(NaAC)}} - \Lambda^{\circ}_{\mathrm{m(NaCl)}}\\ &= (425.9 + 91.0 - 126.4)\mathrm{S~cm}^2 \mathrm{~mol}^{-1}\\ &= 390.5 \mathrm{~S~cm}^2 \mathrm{~mol}^{-1}\end{aligned}$$

Question169

Which of the following acts as oxidising agent in hydrogen-oxygen fuel cell ?

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Options:

A. H_2

B. KOH

C. O_2

D. C

Answer: B

Solution:

Hydrogen-oxygen fuel cell is an electrochemical cell that converts the chemical energy of hydrogen which is a fuel and oxygen which act as an oxidising agent into electricity through a pair of redox reaction. Thus, option (b) is a correct.

Question170

Standard hydrogen electrode (SHE) is a

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Options:

A. primary reference electrode

B. secondary reference electrode

C. metal - sparingly soluble salt electrode

D. metal - metal ion electrode

Answer: A

Solution:

Standard hydrogen electrode (SHE) is primaly reference electrode. SHE is represented by $\mathrm{Pt(s)}\,|\mathrm{H_2(\ g)}|\mathrm{H^+(aq)}$, is assigned a zero potential at all temperatures corresponding to reaction.

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

Question171

The conductivity of an electrolytic solution decreases on dilution due to

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Options:

A. decrease in number of ions per unit volume

B. increase in ionic mobility of ions

C. increase in percentage ionisation

D. increase in number of ions per unit volume

Answer: A

Solution:

Conductivity always decreases with decrease in concentration (i.e. with dilution) of both the strong and weak electrolytes. This is due to the fact that the number of ions that carry current in a volume of solution always decreases with decrease in concentration.

Question172

Two electrolytic cells are connected in series containing $\mathrm{CuSO_4}$ solution and molten $\mathrm{AlCl_3}$. If in electrolysis 0.4 moles of 'Cu' are deposited on cathode of first cell. The number of moles of 'Al' deposited on cathode of the second cell is

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Options:

A. 0.6 moles

B. 0.27 moles

C. 0.18 moles

D. 0.4 moles

Answer: B

Solution:

Given,

Number of moles of Cu deposited $= 0.4$ moles

According to Faraday's second law,

$$\frac{\text{weight of Cu deposited}}{\text{weight of Al deposited}} = \frac{\text{Eq wt. of Cu}}{\text{Eq wt. of Al}} \quad \text{.... (i)}$$

$\because$ No. of moles $= \frac{\text{weight}}{\text{molecular weight}}$

$\therefore$ Weight of Cu $= 0.4 \times 63.5$

Now, from Eq. (i),

$$=\frac{0.4\times 63.5}{\text{weight of Al deposited}}=\frac{\frac{63.5}{2}}{\frac{27}{3}}$$

$\therefore$ Weight of Al deposited $=\frac{0.4\times 63.5\times 9}{31.75}=7.2\text{ g}$

Now, number of moles of Al deposited $=\frac{7.2}{27}=0.27$ moles.

Question173

The resistance of $\frac{1}{10}M$ solution is 25×10^{3} ohm. What is the molar conductivity of solution? (cell constant $=1.25\text{ cm}^{-1}$)

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Options:

A. $3.5\text{ ohm}^{-1}\text{ cm}^{2}\text{ mol}^{-1}$

B. $5.00\text{ ohm}^{-1}\text{ cm}^{2}\text{ mol}^{-1}$

C. $2.50\text{ ohm}^{-1}\text{ cm}^{2}\text{ mol}^{-1}$

D. $2.00\text{ ohm}^{-1}\text{ cm}^{2}\text{ mol}^{-1}$

Answer: B

Solution:

Given, Molarity $=\frac{1}{10}M$

$R=2.5\times 10^{3}\Omega$

Cell constant $\left(\frac{l}{a}\right)=1.25\text{ cm}^{-1}$

Now, κ(conductivity) $=\frac{1}{\kappa}\times\frac{l}{a}$

$\therefore\kappa=\frac{1}{2.5\times 10^{3}}\times 1.25=0.5\times 10^{-3}\Omega^{-1}\text{ cm}^{-1}$

$\Rightarrow$ Molar conductivity $(\Lambda_{\mathrm{m}})=\frac{\kappa\times 1000}{M}$

$$\therefore \quad \Lambda_{\mathrm{m}}=\frac{0.5 \times 10^{-3} \times 1000}{1/10}$$

$$\Lambda_{\mathrm{m}}=5 \Omega^{-1} \mathrm{~cm}^{2} \mathrm{~mol}^{-1}$$

Question174

The correct representation of Nernst's equation for half-cell reaction $\mathrm{Cu}^{2+}(\mathrm{aq})+\mathrm{e}^{-} \longrightarrow \mathrm{Cu}^{+}(\mathrm{aq})$ is

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Options:

A.

$E_{\mathrm{Cu}^{+}}^{\circ}, \mathrm{Cu}^{2+}=E_{\mathrm{Cu}^{+}}^{\circ}, \mathrm{Cu}^{2+}-\frac{0.0592}{2} \log \frac{\left[\mathrm{Cu}^{+}\right]}{\left[\mathrm{Cu}^{2+}\right]}$

B.

$E_{\mathrm{Cu}^{+}, \mathrm{Cu}^{2+}}=E_{\mathrm{Cu}^{+}, \mathrm{Cu}^{2+}}^{\circ}-\frac{0.0592}{1} \log \frac{\left[\mathrm{Cu}^{+}\right]}{\left[\mathrm{Cu}^{2+}\right]}$

C.

$E_{\mathrm{Cu}^{+}, \mathrm{Cu}^{2+}}=E_{\mathrm{Cu}^{+}, \mathrm{Cu}^{2+}}-\frac{0.0592}{2} \log \frac{\left[\mathrm{Cu}^{+}\right]}{\left[\mathrm{Cu}^{2+}\right]}$

D.

$E_{\mathrm{Cu}^{+}, \mathrm{Cu}^{2+}}=E_{\mathrm{Cu}^{+}}^{\circ} ; \mathrm{Cu}^{2+}-\frac{0.0592}{1} \log \frac{\left[\mathrm{Cu}^{+}\right]}{\left[\mathrm{Cu}^{2+}\right]}$

Answer: B, D

Solution:

For the half-cell

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

The correct Nernst's equation is

$$E_{Cu^+,Cu^{2+}} = E_{Cu^+,Cu^{2+}} - \frac{0.0591}{1} \log \frac{[Cu^+]}{[Cu^{2+}]}$$
