NCERT Solutions for Class 12 Chemistry

Chapter 2 – Electrochemistry

Intext Questions with Solutions of Class 12 Chemistry Chapter 2 – Electrochemistry

2.1

How would you determine the standard electrode potential of the system $Mg^{2+}|Mg$?

📌 ***Key Points to Note from Question***

What to Conclude:

- Use a standard hydrogen electrode to measure E_{cell}.
- The negative cell potential indicates that magnesium has a lower reduction potential than hydrogen.

Ans – The cell's electromagnetic field (EMF) can be measured with the angle of deflection using a voltmeter. A cell can be placed with Mg/$MgSO_4$ (1 M) as the primary electrode and the conventional hydrogen electrode Pt, H, (1 atm) H^+/(l M) as the secondary electrode. The shift in direction indicates that e^{-1} s travel from the magnesium electrode to the hydrogen electrode, meaning that reduction occurs on the hydrogen electrode. In contrast, oxidation occurs on the magnesium electrode. The structure of the cell can therefore be shown in the following manner:

$$Mg|Mg^{2+}(1\ M)||H^+\ (1\ M)|H_2,(1\ atm),Pt$$

$$E^o_{cell} = E^o_{H^+/\frac{1}{2}H_2} - E^o_{Mg^{2+}/Mg}$$

$$\text{Put } E^o_{H^+/\frac{1}{2}H_2} = 0$$

$$\therefore E^o_{Mg^{2+}/Mg} = -E^o_{cell}$$

2.2

Can you store copper sulphate solutions in a zinc pot?

Ans – Cu is readily removed from the $CuSO_4$ mixture in the subsequent process because it is less sensitive than zinc:
$Zn\ (s) + CuSO_4\ (aq) \rightarrow ZnSO_4\ (aq) + Cu\ (s)$

When expressed in terms of emf, it will be as follows:
$E^0_{cell} = E^0_{Cu2+/Cu} - E^0_{Zn2+/Zn}$
$= 0.34V - (-0.76V) = 1.10V$

The presence of involuntary responses is indicated by a positive E^0_{cell} measurement. The uniqueness of $CuSO_4$ will be lost if it is kept in a Zn container because Zinc will interact with copper.

2.3

Consult the table of standard electrode potentials and suggest three substances that can oxidize ferrous ions under suitable conditions.

Ans – The oxidation value of Fe^{2+} ions to Fe^{3+} ions can be depicted below:
$Fe^{2+} \rightarrow {}^{3+} + e^-$; $E^0_{OX} = -0{\cdot}77\ V$
Particular compounds that are positioned over iron in an electrochemical sequence or that can oxidize Fe^{2+} ions to Fe^{3+} ions that can take up electrons generated during oxidation are capable of doing so. Meanwhile, in the acidic medium, the substances include $Cl_2(g)$, $Br_2(g)$ & $Cr_2O_2{}^{-7}$ ions.

2.4

Calculate the potential of hydrogen electrode in contact with a solution whose pH is 10.

📌 ***Key Points to Note from Question***

Given Data:

- pH of the solution: **10**
- Temperature (T): **298 K**

Key Concept:

- Potential of a Hydrogen Electrode.

What to Calculate:

- Substitute pH = 10 into the formula to calculate E.

Ans – For a hydrogen electrode,

$$H^+ + e^- \rightarrow \frac{1}{2}H_2$$

It is given that pH = 10.
Using Nernst equation we get,

$$E_{H - \frac{1}{2}H_2} = E^0_{H + \frac{1}{2}H_2} - \frac{0.0591}{n}\log\frac{1}{[H^+]}$$

$$= 0 - \frac{0.0591}{1}\log\frac{1}{10^{-10}}$$

$$= -0.0591 \times 10$$

$$= -0.591\ V$$

Therefore, the potential of the hydrogen electrode in contact with a solution at pH 10 is 0.591 V.

2.5

Calculate the emf of the cell in which the following reaction takes place: $Ni_{(a)} + 2Ag^+(0.002M) \rightarrow Ni^+(0.160M) + 2Ag_{(a)}$
Given that = E°(cell)1.05 V.

📌 ***Key Points to Note from Question***

Given Data:

- Cell reaction,
- Standard emf of the cell (E°_{cell}): **1.05 V**
- Concentrations:
  - $[Ag^+]$ = 0.002 M
  - $[Ni^{2+}]$ = 0.160 M
- Temperature (T): **298 K**

Key Concept:

- Nernst Equation

What to Calculate:

- Reaction quotient (Q).
- Cell potential (E_{cell}) using the Nernst equation.

Ans – Using Nernst formula,

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

$$= 1.05\ V - \frac{0.0591}{2}\log\frac{0.160}{(0.002)^2}$$

$$= 1.05 = \frac{0.0591}{2}\log(4 \times 10^4)$$

$$= 1.05 - \frac{0.0591}{2}(4.6021)$$

$$= 1.05 - 0.14\ V = 0.91\ V$$

2.6

0.236V at 298K. calculate the standard Gibbs energy and the equilibrium constant of the cell reaction.

📌 ***Key Points to Note from Question***

Given Data:

- Standard cell potential (E°_{cell}): **0.236 V**
- Temperature (T): **298 K**
- Faraday's constant (F): **96500 C/mol**
- Gas constant (R): **8.314 J/mol·K**

Key Concept:

- Standard Gibbs Free Energy Formula.

What to Calculate:

- Standard Gibbs free energy.
- Equilibrium constant of the cell reaction.

Ans –

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

For the given cell, $n = 2$

$$\Delta_r G^o = -nF\,E^o_{cell}$$

$$= -2 \times 96500 \times 0.236$$

$$= -45.55 \text{ kJ mol}^{-1}$$

Also, $\Delta_r G^o = -2.303\, RT \log K_C$

$$\Rightarrow \log K_C = \frac{-\Delta_r G^o}{2.303RT}$$

$$= \frac{-45.55}{2.303 \times 8.314 \times 10^{-3} \times 298} = 7.983$$

$$\Rightarrow K_C = \text{antilog}\,(7.983)$$

$$= 9.616 \times 10^7$$

2.7

Why does the conductivity of solution decrease with dilution?

Ans – The quantity of ions available per unit volume is correlated with a solution's conductance. When they are diluted, the solution's particular conductive property or equivalent conductivity also lowers.

2.8

Suggest a way to determine the value Λ° of water.

📌 ***Key Points to Note from Question***

Key Concept:

- Application of Kohlrausch's Law.

What to Calculate:

- Use experimental or literature values of $\lambda^0(H^+)$ and $\lambda^0(OH^-)$.
- Add them to find the limiting molar conductivity of water.

Ans – We can figure out Λm° for water by applying Kohl Rausch's law.
Irrespective of the type of ion concerned, the Kohlrausch Law states that every ion contributes a particular quantity to the electrolyte's comparable conductivity when splitting takes place at endless dilution. The quantity of corresponding conductance at infinite dilution for every

electrolyte is the combined value of the individual contributions of its ionic constituents (cations followed by anions).

$\Lambda m^\circ = \Lambda m^\circ (HCl) + \Lambda m^\circ (NaOH) - \Lambda m^\circ (NaCl)$

Since HCl, NaOH, and NaCl are powerful electrolytes and fully split up, their Λ° constants are well-known. The magnitude of Λm° for water molecules may be obtained by entering the corresponding value in the formula mentioned earlier.

2.9

The molar conductivity of 0.025 mol L^{-1} methanoic acid is 46.1 S cm^2 mol^{-1}. Calculate its degree of dissociation and dissociation constant Given $\lambda^0(H^+)$ = 349.6 cm^2 and $\lambda^\circ(HCOO^-)$ = 54.6 cm^2 mol^{-1}

📌 ***Key Points to Note from Question***

Given Data:

- Concentration of methanoic acid (C): **0.025 mol/L**
- Molar conductivity (Λ_m): **46.1 S cm^2 mol^{-1}**
- Limiting molar conductivities

Key Concept:

- Degree of Dissociation formula.
- Dissociation Constant formula.

What to Calculate:

- Limiting molar conductivity.
- Degree of dissociation (α).
- Dissociation Constant (K_a)

Ans –

$$\Lambda_m^o(HCOOH) = \lambda^o(H^+) + \lambda^o(HCOO^-)$$

$$= 349.6 + 54.6 = 404.2 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\Lambda_m^C = 46.1 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\therefore \alpha = \frac{\Lambda_m^C}{\Lambda_m^o} = \frac{46.1}{404.2} = 0.114$$

	$HCOOH \rightleftharpoons$	$HCOO^{-1}$ +	H^+
Initial conc,	c	0	0
At equi,	$c(1-\alpha)$	$c\alpha$	$c\alpha$

$$\therefore K_a = \frac{c\alpha.\, c\alpha}{c(1-\alpha)} = \frac{c\alpha^2}{1-\alpha}$$

$$= \frac{0.025 \times (0.114)^2}{1 - 0.114} = 3.67 \times 10^{-4}$$

2.10

If a current of 0.5 ampere flows through a metallic wire for 2 hours, then how many electrons would flow through the wire?

📌 ***Key Points to Note from Question***

Given Data:

- Current (I): **0.5 A**
- Time (t): **2 hours = 7200 seconds**

Key Concept:

- Charge (Q) formula.
- Number of electrons formula.

What to Calculate:

- Total charge (Q) passing through the wire.
- Number of electrons (n_e) flowing through the wire.

Ans –

$Q = It$

$= 0.5 \times (2 \times 60 \times 60) = 3600C$

$1\ F \Rightarrow 96500C \Rightarrow 1$ mole of e^{-1} s

$\therefore 6.02 \times 10^{23}\ e^{-1}$ s

$\therefore$ 3600C is equivalent to the flow of e^{-1} s

$$= \frac{6.02 \times 10^{23}}{96500} \times 3600$$

$$= 2.246 \times 10^{22}\ e^{-1}s$$

2.11

Suggest a list of metals that are extracted electrolytically.

Ans – Electrolysis can be used to obtain extremely reactive elements with significant -ve E° coefficients that can also function as potent reducing substances. Electrolytic reduction is the name of the method that is used. For example, elements like magnesium, sodium, calcium, potassium, etc.

2.12

Consider the reaction: $CrO_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$. What is the quantity of electricity in coulombs needed to reduce 1mol of $Cr_2O_7^{2-}$

📌 ***Key Points to Note from Question***

Given Data:

- Chemical reaction.
- Number of electrons transferred per mole of $Cr_2O_7^{2-}$: **6 moles of e^-**

Key Concept:

- Quantity of Electricity (Q) Formula

What to Calculate:

- Total charge (Q) required to reduce 1 mole of $Cr_2O_7^{2-}$.

Ans – For the mentioned reaction to occur, 1 mol of $Cr_2O_7^{2-}$ requires 6 F = 6 × 96500 = 579000 C of electricity.

Therefore, 579000 C of electricity are necessary for the reduction of $Cr_2O_7^{2-}$ to Cr^{3+}

2.13

Write the chemistry of recharging the lead storage battery, highlighting all the materials that are involved during recharging

📌 ***Key Points to Note from Question***

Key Concept:

- Electrochemical Reaction During Recharging.
- Half-Reactions at Electrodes.
- Overall Reaction During Recharging.

What to Highlight:

- Reactions at both electrodes.
- Overall chemical process restoring the battery's capacity.

Ans – An anode made of lead, a cathode of a network of lead filled with lead dioxide (PbO_2), and an electrolyte of 38% water in a solution of H_2SO_4 make up a lead storage cell. The following effect occurs while the battery is being used:

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

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

Overall Reaction: $Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$

The opposite process occurs when the battery is charged, meaning that $PbSO_4$ that has been accumulated on the conducting surfaces is changed back into Pb and PbO_2, while H_2SO_4 is produced.

2.14

Suggest two materials other than hydrogen that can be used as fuels in fuel cells.

Ans – Methane (CH_4) & methanol (CH_3OH) can be implemented as fuel source as a perfect alternate to hydrogen components in the fuel storage cells.

2.15

Explain how rusting of iron is envisaged as setting up of an electro chemical cell.

📌 ***Key Points to Note from Question***

Key Concept:

- Rusting as an Electrochemical Process.
- Electrochemical Reactions.
- Formation of Rust.
- Role of Water and Oxygen

What to Explain:

- The anode and cathode regions on the iron surface.
- How rust forms as a product of electrochemical reactions.
- Role of water and oxygen in the process.

Ans – The water available on the above layer of the surface in the iron gets dissolved in the acidic oxides of air including CO_2, and SO_2 to produce acids that spits apart to offer H^+ ions:

$H_2O + CO_2 \rightarrow H_2CO_3 \rightleftharpoons 2H^+ + CO_3^{2-}$

Iron Fe^{2+} emits e^{-1} to produce Fe^{3+} when H^+ is present. Thus, this serves as an anode:

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

After that, the e^{-1} dissipated moves across the metal to a location where reduction occurs and H^+ ions use these electrons together with oxygen that is dispersed serves as a cathode as a result:

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

Overall reaction will be:

$2Fe\ (s) + O_2\ (g) + 4H^+\ (aq) \rightarrow 2Fe^{2+}\ (aq) + 2H_2O\ (l)$

Consequently, the outer layer develops an electrochemical cell. Rust, or hydrated ferric oxide, is created when water combines with ferrous ions that have been oxidized by oxygen in the atmosphere to ferric ions ($Fe_2O_3.XH_2O$).

Exercise Questions with Solutions of Class 12 Chemistry Chapter 2 – Electrochemistry

2.1

Arrange the following metals in the order in which they displace each other from the solution of their salts: Al, Cu, Fe, Mg and Zn

Ans – Mg, Al, Zn, Fe, Cu

2.2

Given the standard electrode potentials,
$K^+/K = 2.93V$
$Ag^+/Ag = 0.80V$
$Hg^{2+}/Hg = 0.79V$
$Mg^{2+}/Mg = -2.73V$
$Cl^{3+}/Cr = 0.74V$
Arrange these metals in their increasing order of reducing power.

📌 ***Key Points to Note from Question***

Key Concept:

- **Reducing Power:**
  - Metals with **lower standard electrode potential (E°)** are stronger reducing agents because they are more likely to lose electrons and get oxidized.
  - Reducing power increases as E° becomes **more negative**.
- **Trend Based on E°:**
  - K^+: **Most negative** → Strongest reducing agent.
  - Ag^+: **Most positive** → Weakest reducing agent.

Ans – The reducing power increases with the oxidation potential because it can be readily oxidized. Ag < Hg < Cr < Mg < K will experience the rising order of diminishing energy as a result.

2.3

Depict the galvanic cell in which the reaction,
$2Ag^+ (aq) \rightarrow Zn^{2+} (aq) + 2Ag(s)$ takes place. Further show:
(i) Which of the electrode is negatively charged?
(ii) The carriers of the current in the cell.
(iii) Individual reaction at each electrode.

📌 ***Key Points to Note from Question***

Key Concept:

- **Electrode Charges:**
  - The **anode (Zn)** is negatively charged as oxidation occurs (loss of electrons).
  - The **cathode (Ag)** is positively charged as reduction occurs (gain of electrons).
- **Current Carriers:**
  - **In the external circuit:** Flow of **electrons** from the anode to the cathode.
  - **In the electrolyte:** Flow of **ions**:
    - Zn^{2+} ions move into the solution at the anode.
    - NO^{3-} ions move toward the anode and Ag^{+} ions move toward the cathode.

Ans – (i) There will be a negative charge on the anode, which is the zinc electrode.

(ii) Current will flow from silver to copper in the external circuit.

(iii) At anode: $Zn\ (s) \rightarrow Zn^{2+}\ (aq) + 2e^{-}$

At cathode: $2Ag^{+}\ (aq) + 2e^{-} \rightarrow 2Ag\ (s)$

2.4

Calculate the standard cell potentials of galvanic cell in which the following reactions take place

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

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

Calculate the Δ_rG^0 and equilibrium constant of the reactions

📌 ***Key Points to Note from Question***

Key Concept:

- Standard Cell Potential Formula
- Gibbs Free Energy Formula
- Equilibrium Constant Formula

What to Calculate:

- Standard cell potential (E°cell) for both reactions.
- Gibbs free energy change ($\Delta_rG°$) for both reactions.
- Equilibrium constant (K) for both reactions.

Ans – (i) Calculation of E^0_{cell}

$$E^o_{cell} = E^o_{cathod} - E^o_{anode}$$
$$= -0.40 - (-0.74) = +0.34V$$

Calculation of ΔG^O

$$\Delta G^o = -nFE^o_{cell} = -(6\ \text{mol}) \times (96500\ \text{C mol}^{-1})(0.34\ \text{V})$$
$$= -196860\ \text{CV} = -196860\ \text{J} = -196.86\ \text{kJ}$$

Calculation of Equilibrium Constant K_c

$$\Delta G^o = -2.303\ RT \log K_c$$
$$\log K_c = (-)\frac{(-)\Delta G^o}{2.303\ RT}$$
$$= (-)\frac{(-)196860}{2.303 \times 8.314 \times 298} = 34.501$$
$$\Rightarrow K_c = \text{Antilog}\,(34.501) = 3.17 \times 10^{34}$$

(ii) Calculation of E^0_{cell}

$$E^o_{cell} = E^o_{cathode} - E^o_{anode}$$
$$= (0.80 - 0.77) = 0.03\ \text{V}$$

Calculation of ΔG^O

$$\Delta G^o = -nFE^o_{cell}$$
$$= -(1\ \text{mol}) \times (96500\ \text{C mol}^{-1})(0.03\ \text{V})$$
$$= -2895\ \text{CV} = -2895\ \text{J} = -2.895\ \text{kJ}$$

Calculation of Equilibrium Constant K_c

$$\Delta G^o = -2.303\ RT \log K_c$$
$$\log K_c = (-)\frac{(-)\Delta G^o}{2.303\ RT}$$
$$= (-)\frac{(-)2895}{2.303 \times 8.314 \times 298} = 0.5074$$
$$\Rightarrow K_c = \text{Antilog}\,(0.5074) = 3.22$$

2.5

Write the Nernst equation and emf of the following cells at 298K
(i) Mg(s) | Mg^{2+} (0.001M) || Cu^{2+} (0.0001M) Cu(s)
(ii) Fe(s) | Fe^{2+} (0.001M) || H^+ (1M) H_2(g)(1bar) Pt(S)
(iii) Sn(s) | Sn^{2+}(0.050M) || H^+ (0.020M) | H_2(g)(1bar) Pt(S)
(iv) Pt(s) | Br_2(l) | Br^- (0.010M) || H^+ (0.030M) H_2(g)(1bar) Pt(S)

📌 ***Key Points to Note from Question***

Given Data:

- **Temperature**: 298 K.
- Use the **Nernst equation** to calculate the emf for each cell.

Key Concept:

- Nernst Equation for a Cell

Ans – (i) Cell reaction:

$$\mathrm{Mg + Cu^{2+} \rightarrow Mg^{2+} + Cu\ (n = 2)}$$

Nernst equation:

$$E_{cell} = E^{o}_{cell} - \frac{0.0591}{2}\log\frac{[Mg^{2+}]}{[Cu^{2+}]}$$

$$\therefore E_{cell} = 0.34 - (-2.37) - \frac{0.0591}{2}\log\frac{10^{-3}}{10^{-4}}$$

$$= 2.71 - 0.02955 = 2.68\ V$$

(ii) Cell reaction:

$$\mathrm{Fe + 2H^{+} \rightarrow Fe^{2+} + H_2\ (n = 2)}$$

Nernst equation:

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

$$\therefore E_{cell} = 0 - (-0.44) - \frac{0.0591}{2}\log\frac{10^{-3}}{(1)^2}$$

$$= 0.44 - \frac{0.0591}{2} \times (-3)$$

$$= 0.44 + 0.0887 = 0.5287\ V.$$

(iii) Cell reaction:

$Sn + 2H^+ \rightarrow Sn^{2+} + H_2\ (n = 2)$

Nernst equation:

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

$$E_{cell} = E^o_{cell} - \frac{0.0591}{2}\log\frac{0.05}{(0.02)^2}$$

$$= 0 - (-0.14) - \frac{0.0591}{2}\log\frac{0.05}{(0.02)^2}$$

$$= 0.14 - \frac{0.0591}{2}\log 125$$

$$= 0.14 - \frac{0.0591}{2}(2.0969) = 0.078\ V$$

(iv) Cell reaction:

$2Br^{-1} + 2H^+ \rightarrow Br_2 + H_2\ (n = 2)$

Nernst equation:

$$E_{cell} = E^o_{cell} - \frac{0.0591}{2}\log\frac{1}{[Br^-]^2[H^+]^2}$$

$$= (0 - 1.08) - \frac{0.0591}{2}\log\frac{1}{(0.01)^2\,(0.03)^2}$$

$$= -1.08 - \frac{0.0591}{2}\log(1.111 \times 10^7)$$

$$= -1.08 - 0.208 = -1.288\ V$$

2.6

In the button cells widely used in watches and other devices the following reaction takes place:
$Zn(s) + Ag_2O(s) + H_2O(l) \rightarrow Zn^{2+}(aq) + 2Ag(s) + 2OH^-(aq)$
Determine Δ_rG° and E° for the reaction.

📌 ***Key Points to Note from Question***

Given Data:

- Given chemical reaction

Key Concept:

- E°cell = E°cathode − E°anode

- $\Delta_r G^\circ = -nFE^\circ cell$

What to Calculate:

- Standard cell potential.
- Standard Gibbs free energy change.

Ans – We know

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

$Ag_2O + H_2O + 2e^- \rightarrow 2Ag + 2OH^-$, $E^\circ = 0.344$ V

In the specified reaction, Zn undergoes oxidation whereas Ag_2O undergoes reduction.

$E^\circ_{cell} = 0.344 + 0.76 = 1.104$ V

$\Delta G^\circ = nFE^\circ_{cell} = -2 \times 96500 \times 1.104$ J

$\Delta G^\circ = -2.13 \times 10^5$ J

2.7

Define conductivity and molar conductivity for the solution of an electrolyte. Discuss their variation with concentration.

📌 ***Key Points to Note from Question***

Key Concept:

- **Conductivity:** Measures a solution's ability to conduct electricity.
- **Molar Conductivity:** Conductivity per mole of electrolyte in solution.

What to Calculate:

- **Variation with Concentration:**
  - For **strong electrolytes:** Λ_m decreases slightly with concentration.
  - For **weak electrolytes:** Λ_m increases significantly with dilution.

Ans – When an agent has a dimension of 1 cm & a cross-sectional measurement of 1 cm, its conductivity is determined as its conductance.

The molar conductivity of a fluid at a dilution (V) is the conductance of every ion generated from a single mole of the electrolyte present in the solution when the electrodes are spaced one centimetre apart and the area of their cross-section is sufficient to contain the whole substance between them Vcm3. The majority of highly incorporated indicators for conductivity is Λm.

As the electrolyte saturation decreases, or as dispersion takes place, a solution's conductivity decreases (for equally weak as well as strong electrolytes). The reason for this is that the quantity of ions per cubic inch of the mixture falls as it becomes dissolved. The molar conductivity of a mixture rises as the quantity of the electrolyte decreases upon diluting. This happens because there is a lower number of ions per cubic inch of the substance when it is dissolved. As a solution's electrolyte saturation decreases, its molar conductivity rises. The reason for this is that dilution raises the number of ions and their adaptability. As the quantity of concentration inches closer to zero, the molar conductivity is depicted as the restricting molar conductivity.

2.8

The conductivity of 0.20M solution of KCl at 298K is 0.0248 S cm^{-1}. Calculate its molar conductivity.

📌 ***Key Points to Note from Question***

Given Data:

- Conductivity: 0.0248 S cm^{-1}
- Concentration: 0.20 M

What to Calculate:

- Molar conductivity

Ans –

$$\Lambda_m = \frac{k \times 1000}{\text{Molarity}}$$

$$= \frac{0.2485\ \text{cm}^{-1} \times 1000\ \text{cm}^3\text{L}^{-1}}{0.20\ \text{mol L}^{-1}} = 124\ \text{Scm}^2\text{mol}^{-1}$$

2.9

The resistance of a conductivity cell containing 0.001M KCl solution at 298K is 1500Ω. What is the cell constant if conductivity of 0.001M KCl solution at 298K is 0.146 × 10^{-3} S cm^{-1}?

📌 ***Key Points to Note from Question***

Given Data:

- Conductivity: 0.146 × 10^{-3} S cm^{-1}
- Resistance: 1500 Ω

What to Calculate:

- Cell constant

Ans –

$$\text{Cell constant} = \frac{\text{Conductivity}}{\text{Conductance}}$$

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

$$= 0.146 \times 10^{-3}\ \text{Scm}^{-1} \times 1500\Omega$$

$$= 0.218\ \text{cm}^{-1}$$

2.10

The conductivity of NaCl at 298K has been determined at different concentrations and the results are given below:

concentration/M	0.001	0.010	0.020	0.050	0.100
$10^2 \times K/S\ m^{-1}$	1.237	11.85	23.15	55.53	106.74

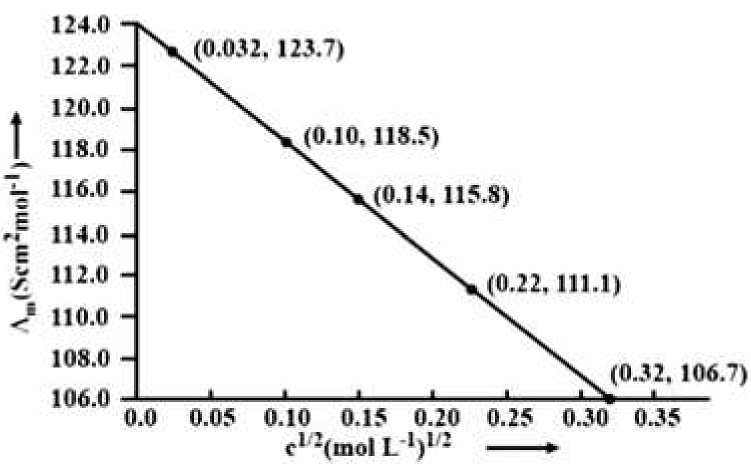
for all concentrations and draw a plot between Λ_m and $C^{1/2}$

Find the value of $\Lambda_m°$

Ans – Using the unit conversion factor shown below,

$$\frac{1\ \text{Scm}^{-1}}{100\ \text{Sm}^{-1}} = 1$$

Concentration (M)	$K(Sm^{-1})$	$K(Scm^{-1})$	$\Lambda_m = \frac{\mathbf{1000 \times k}}{\mathbf{Molarity}}(Scm^2mol^{-1})$	$C^{\frac{1}{2}}(M^{\frac{1}{2}})$
10^{-3}	1.237×10^{-2}	1.237×10^{-4}	$\frac{1000 \times 1.237 \times 10^{-4}}{10^{-3}} = 123.7$	0.0316
10^{-2}	11.85×10^{-2}	11.85×10^{-4}	$\frac{1000 \times 11.85 \times 10^{-4}}{10^{-2}} = 1118.5$	0.100
2×10^{-2}	23.15×10^{-2}	23.15×10^{-4}	$\frac{1000 \times 23.15 \times 10^{-4}}{2 \times 10^{-2}} = 115.8$	0.141
5×10^{-2}	55.53×10^{-2}	55.53×10^{-4}	$\frac{1000 \times 55.53 \times 10^{-4}}{5 \times 10^{-2}} = 111.1$	0.224
10^{-1}	106.74×10^{-2}	106.74×10^{-4}	$\frac{1000 \times 106.74 \times 10^{-4}}{10^{-1}} = 106.7$	0.316



Λ°cm = Intercept of $\Lambda_m$axis =124.0 S cm^2 mol^{-1}

The extrapolation to zero concentration produces this outcome.

2.11

Conductivity of 0.00241 M acetic acid is 7.896×10^{-5} S cm^{-1}. Calculate its. Calculate its molar conductivity. If Λ°m for acetic acid is 390.5 S cm^2 mol^{-1}, what is its dissociation constant?

📌 ***Key Points to Note from Question***

Given Data:

- Conductivity: 7.896×10^{-5} S cm^{-1}
- Concentration: 0.00241 M
- Λ^0_m (limiting molar conductivity): 390.5 S cm^2 mol^{-1}

What to Calculate:

- Molar conductivity.
- Degree of dissociation.
- Dissociation constant.

Ans –

$$A^c_m = \frac{k \times 1000}{\text{Molarity}}$$

$$= \frac{7.896 \times 10^{-5} \text{Scm}^{-1} \times 1000 \text{ cm}^3\text{L}^{-1}}{0.241 \times \text{mol L}^{-1}}$$

$$= 32.76 \text{ Scm}^2 \text{ mol}^{-1}$$

$$\alpha = \frac{A^c_m}{Am_o} = \frac{32.76}{390.5} = 8.4 \times 10^{-2}$$

$$k_a = \frac{C\alpha^2}{1-\alpha} = \frac{0.24 \times (8.4 \times 10^{-2})^2}{1 - 0.084}$$

$$= 1.86 \times 10^{-5}$$

2.12

How much charge is required for the following reductions:
(i) 1 mol of Al^{3+} to Al?
(ii) 1 mol of Cu^{2+} to Cu?
(iii) 1 mol of MnO_4^- to Mn^{2+}?

Ans – (i) Al^{3+} + 3e → Al is the process that occurs when aluminium ions lose 3 electrons to reach the elemental form of aluminium. For a decrease of 1mol, the required energy is Al^{3+} = 3F = 3 × 96500C = 289500C.

(ii) For it to exist in the copper (0) region, a cupric ion must shed 2 electrons. The following is the electrode response that is taking place: Cu^{2+} + $2e^-$ → Cu. The amount of energy needed to convert one mol of Cu^{2+} is equal to 2F = 2 × 96500 = 193000C.

(iii) 5 electrons are released to decrease manganese from its +7 oxidation condition to a +2 oxidation stage in the molecule MnO_4.

$MnO_4 \rightarrow Mn^{2+}$ i.e. $Mn^{7+} + 5e^- \rightarrow Mn^{2+}$

One mol of MnO_4 must be reduced to Mn^{2+} with an electrical charge value of 5F = 5 × 96500C = 4825000C.

2.13

How much electricity in terms of Faraday is required to produce.
(i) 20.0g of Ca from molten $CaCl_2$?
(ii) 40.0g of Al from molten Al_2O_3?

Ans – (i) $Ca^{2+} + 2e^- \rightarrow Ca$

One mol of calcium, or 40g of calcium, will call for = 2F electricity since the process above involves a pair of electrons, and 20g of calcium will need = 1F energy.

(ii) $Al^{3+} + 3e^- \rightarrow Al$

Since there are 3 electrons involved in the decrease of the aluminium ion previously mentioned, Thus, 27g of Al or one mole of Al will need 3F of power.
Additionally, 40g of aluminium will take 4.44F of power (3/27) times 40.

2.14

How much electricity is required in coulomb for the oxidation of
(i) 1 mol of H_2O to O_2?
(ii) 1 mol of FeO to Fe_2O_3?

Ans – (i) The following reaction happen when water oxidises:

$$H_2O \rightarrow H_2 + \frac{1}{2} O_2$$

Or

$$O^{2-} \rightarrow \frac{1}{2} O_2 + 2e^-$$

Two electrons are involved in the transaction. Thus, the required amount of electricity is equal to 2 F = 2 × 96500 C = 193000 C

(ii) The following reaction is how the oxidation process of FeO happen:

$FeO + \frac{1}{2}O_2 \rightarrow \frac{1}{2}Fe_2O_3$

Or

$Fe^{2+} \rightarrow Fe^{3+} + e^-$

The electron transfer involves a single electron unit; hence, the required quantity of electricity = 1 F = 96500C

2.15

A solution of $Ni(NO_3)_2$ is electrolyzed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode?

📌 ***Key Points to Note from Question***

Given Data:

- Current (I): 5 A
- Time (t): 20 min = 1200 s
- Molar mass of Ni: 58.69 g/mol
- Faraday's constant (F): 96500 C/mol

What to Calculate:

- Total charge (Q).
- Moles of electrons (n_e).
- Mass of Ni deposited.

Ans – The amount of electricity that was passed = (5A) x (20 x 60 sec) = 6000C

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

Thus, 2F, i.e. $2 \times 96500C$ of charge deposit

$= 1$ mole of Ni $= 58.7$ g

$\therefore$ 6000C of charge will deposit

$$= \frac{58.7 \times 6000}{2 \times 96500} = 1.825\text{g of Ni}$$

2.16

Three electrolytic cells, A, B, C containing solutions of $ZnSO_4$, $AgNO_3$ and $CuSO_4$, respectively are connected in series. A steady current of 1.5 amperes was passed through them until 45g of silver deposited at the cathode of cell B. How long did the current flow? What mass of copper and zinc were deposited?

📌 ***Key Points to Note from Question***

Given Data:

- Current (III): 1.5 A
- Mass of silver (Ag) deposited: 45 g
- Molar mass of Ag: 107.87 g/mol
- Molar mass of Cu: 63.55 g/mol
- Molar mass of Zn: 65.38 g/mol
- Faraday's constant (F): 96500 C/mol

What to Calculate:

- Total charge (Q) required for 45 g.
- Time (t) the current flowed.
- Mass of zinc and copper deposited in cells A and C, respectively.

Ans – Given

$I = 1.5\ \mathrm{A}, W = 1.45\ \mathrm{g\ of\ Ag}, E = 108, n = 1$

Using Faraday's 1st law of electrolysis $W = ZIt$

$\Rightarrow W = \frac{E}{nF} It$

$\Rightarrow t = \frac{1.45 \times 96500}{1.5 \times 108} = 863.77$ seconds.

Now for Cu, $W_1 = 1.45$ g of Ag, $E_1 = 108$, $E_2 = 31.75$

From Faraday's second law of electrolysis $\frac{W_1}{W_2} = \frac{E_1}{E_2}$

$$\frac{1.45}{W_2} = \frac{108}{31.75}$$

$$\therefore W_2 = \frac{1.45 \times 31.75}{108}$$

$= 0.426$ g of Cu

For Zn, $W_1 = 1.45$ g of Ag, $E_1 = 108$, $E_2 = 32.65$

Using formula, $\frac{W_1}{W_2} = \frac{E_1}{E_2}$

$$\frac{1.45}{W_2} = \frac{108}{32.65}$$

$$\therefore W_2 = \frac{1.45 \times 32.65}{108}$$

$= 0.438$ of Zn.

2.17

Using the standard electrode potentials given in Table 2.1, predict if the reaction between the following is feasible:

(i) Fe^{3+} (aq) and I^- (aq)

(ii) Ag^+ (aq) and Cu (s)

(iii) Fe^{3+} (aq) and Br^- (aq)

(iv) Ag (s) and Fe^{3+} (aq)

(v) Br_2 (g) and Fe^{2+} (aq)

Ans – A positive emf indicates the feasibility of the cell reaction.

(i) Fe^{3+} (aq) + I^- (aq) → Fe^{2+} (aq) + $\frac{1}{2}$ I_2 (g)

$\therefore E^\circ_{cell} = 0.77 - 0.54 = 0.23V$ (Feasible)

(ii) $2Ag^+$ (aq) + Cu (s) → 2Ag (s) + Cu^{2+} (aq)

$\therefore E^\circ_{cell} = 0.80 - 0.34 = 0.46V$ (Feasible)

(iii) Fe^{3+} (aq) + Br^- (aq) → Fe^{2+} (aq) + $\frac{1}{2}$ Br_2 (g)

$\therefore E^\circ_{cell} = 0.77 - 1.09 = -0.32V$ (Not feasible)

(iv) Ag (s) + Fe^{3+} (aq) → Ag^+ (aq) + Fe^{2+} (aq)

$\therefore E^\circ_{cell} = 0.77 - 0.80 = -0.03V$ (Not feasible)

(v) $\frac{1}{2}$ Br_2 (g) + Fe^{2+} (aq) → Br^- (aq) + Fe^{3+} (aq)

$\therefore E^\circ_{cell} = 1.09 - 0.77 = 0.32V$ (Feasible)

2.18

Predict the products of electrolysis in each of the following.
(i) An aqueous solution of $AgNO_3$ with silver electrodes.
(ii) An aqueous solution of $AgNO_3$ with platinum electrodes.
(iii) A dilute solution of H_2SO_4 with platinum electrodes.
(iv) An aqueous solution of $CuCl_2$ with platinum electrodes.

Ans – (i) $AgNO_3$ (s) → Ag^+ (aq) + NO_3^- (aq)

$H_2O \leftrightarrows H^+ + OH^-$

At the cathode, Ag^+ ions have a smaller release value than H^+ ions. Consequently, instead of settling down as H^+ ions, Ag ions will be implanted as Ag.

At the anode: Ag melts and creates ions in the mixture once NO_3^- ions hit the Ag anode.

Ag (s) → Ag^+ (aq) + e^-

(ii) At the cathode, Ag^+ ions have a smaller release value than H^+ ions. Consequently, instead of settling down as H^+ ions, Ag ions will be formed as Ag.

At anode: OH ions have a smaller discharging energy than NO_3^- ions due to the anode's inability to be attacked. Consequently, NO_3^- ions will dissolve and emit O_2 once OH has been eliminated initially.

OH^- → OH + e^-
4OH → 2H2O (l) + O_2 (g)

(iii) H_2SO_4 → $2H^+$ + SO_4^{2-}

$H_2O \rightleftharpoons H + OH^-$

At cathode: H^+ + e^- → H
H + H → H_2 (g)

At anode: OH^- → OH + e^-
4OH → $2H_2O$ + O_2 (g)
Thus, hydrogen gas is released at the cathode and oxygen gas at the anode.

(iv) $CuCl_2$ → Cu^{2+} + $2Cl^-$

$H_2O \rightleftharpoons H + OH^-$

Copper metal will be deposited at the cathode as an outcome of Cu^{2+} ions diminishing to greater H^+ ions.

Cu^{2+} + $2e^-$ → Cu

At the anode, OH^- ions that remain in reaction will be liberated after Cl^- ions.

$Cl^- \rightarrow Cl + e^-$

$Cl + Cl \rightarrow Cl_2$ (g)

As a result, Cl_2 vapor will be released at the anode & copper metal will settle on the cathode.