

Chapter	10 Electrochemistry	Class	1 st Year	Subject	Chemistry
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Give Short Answers of the Following Questions

1. How and why electrical double layer is formed?

Ans. An electrical double layer forms at the interface between a solid electrode and an electrolyte solution. It consists of two layers: a layer of charge on the electrode surface and a layer of opposite charge in the solution. This structure arises due to electrostatic interactions between the electrode and the ions in the solution.

2. Why electrode potential of Cu is called reduction potential?

Ans. The electrode potential of Cu is called a reduction potential because it represents the tendency of Cu^{2+} ions to gain electrons (reduction) and form solid copper. A positive value indicates a strong tendency to gain electrons.

3. What are the advantages of salt bridge in a galvanic cell?

Ans. A salt bridge:

- Maintains electrical neutrality by allowing ion flow between half-cells.
- Prevents the solutions from mixing while completing the circuit.
- Enables continuous electron flow, allowing the cell to produce electricity.

4. How can we predict the feasibility of a chemical reaction using the cell voltage?

Ans. The feasibility of a reaction can be predicted by calculating the standard cell potential (E_{cell}°). A positive E_{cell}° indicates a spontaneous reaction, while a negative E_{cell}° suggests a non-spontaneous reaction.

5. During electrolysis of aqueous NaCl, why Na is not liberated at the cathode?

Ans. Sodium (Na) is not liberated at the cathode during the electrolysis of aqueous NaCl because water has a lower reduction potential than Na^{+} ions. Therefore, water is reduced to hydrogen gas instead of Na^{+} being reduced to Na metal.

6. Calculate the Ox. No. of chromium (Cr) in the following compounds:

i) CrCl_3

ii) $\text{Cr}_2(\text{SO}_4)_3$

iii) $\text{Cr}_2\text{O}_7^{2-}$

Ans. Oxidation Number of Chromium (Cr)

(i) CrCl_3 :

Given Data: Oxidation number of Cl = -1

Required: Oxidation number of Cr = x ?

Solution:

Oxidation No. of Cr + 3(Oxidation No. of Cl) = 0

$$x + 3(-1) = 0$$

$$x - 3 = 0$$

$$x = +3$$

Oxidation number of Cr is +3 in CrCl_3 .

(ii) $\text{Cr}_2(\text{SO}_4)_3$:

Given Data: Oxidation number of O = -2

Oxidation number of S = +6

Required: Oxidation number of Cr = x = ?

Solution:

2(Oxidation No. of Cr) + 3(Oxidation No. of S)

$$+ 12(\text{Oxidation No. of O}) = 0$$

$$2x + 3(+6) + 12(-2) = 0$$

$$2x = 24 - 18 = +6$$

$$x = +3$$

Oxidation number of Cr is +3 in $\text{Cr}_2(\text{SO}_4)_3$.

(iii) $\text{Cr}_2\text{O}_7^{2-}$:

Given Data: Oxidation number of O = -2

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Required: Oxidation number of Cr = x ?

Solution:

$$2(\text{Oxidation No. of Cr}) + 7(\text{Oxidation No. of O}) = -2$$

$$2x + 7(-2) = -2$$

$$2x - 14 = -2$$

$$2x = +14 - 2 = +12$$

$$x = +6$$

Oxidation number of Cr is +6 in $\text{Cr}_2\text{O}_7^{2-}$.

7. The order of decreasing reactivity of metals based on their position is $\text{K} > \text{Mg} > \text{Zn} > \text{Fe} > \text{Cu}$. Write balanced chemical equations for the reactions that would occur (if any) when:

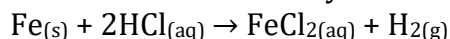
i) **Copper is added to a solution of magnesium sulphate.**

ii) **Iron is added to a dilute solution of hydrochloric acid.**

Ans. Copper is added to a solution of magnesium sulphate:

No reaction occurs because copper is less reactive than magnesium.

Iron is added to a dilute solution of hydrochloric acid:



8. Explain why some metals higher in the activity series can displace hydrogen from acids, while other lower in the series cannot.

Ans. Metals higher in the activity series (e.g., Zn, Mg) are more reactive and have a greater tendency to lose electrons (oxidize) compared to hydrogen. When reacted with acids, these metals displace H^+ ions (reducing them to H_2 gas) because their reduction potential is lower than hydrogen's.

Metals lower in the series (e.g., Cu, Ag) are less reactive than hydrogen. Their reduction potential is higher meaning they cannot supply electrons to reduce H^+ ions, so they do not displace hydrogen from acid.

In short, electrochemical reactivity (position in the series) determines whether a metal can reduce H^+ to H_2 .

9. Calculate the number of Faradays required to deposit 108 g of Ag^{1+} , 63.5 g of Cu^{2+} and 27 g of Al^{3+} .

Ans. Number of Faradays = Mass of substance (g)/equivalent mass of substance

The equivalent mass (E) is calculated as:

$$E = \frac{\text{Molar mass (M)}}{\text{Valency (n)}}$$

1. For 108 g of Ag^+ (Silver)

Molar mass (Ag) = 108 g/mol

Valency = 1

Equivalent mass (E) = $\frac{108}{1} = 108$ g/equivalent

Number of Faradays = $\frac{108g}{108g/equivalent} = 1F$

2. For 63.5 g of Cu^{2+} (Copper)

Molar mass (Cu) = 63.5 g/mol

Valency = 2

Equivalent mass (E) = $\frac{63.5g}{2} = 31.75$ g/equivalent

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$$\text{Number of Faradays} = \frac{63.5g}{31.75g/\text{equivalent}} = 2F$$

3. For 27 g of Al^{3+} (Aluminum)

$$\text{Molar mass (Al)} = 27 \text{ g/mol}$$

$$\text{Valency} = 3$$

$$\text{Equivalent mass (E)} = \frac{27}{3} \text{ g/equivalent}$$

$$\text{Number of Faradays} = \frac{27g}{9g/\text{equivalent}} = 3F$$

Final Answer:

- Ag^+ (108 g): 1 Faradays
- Cu^{2+} (63.5 g): 2 Faradays
- Al^{3+} (27 g): 3 Faradays

(Each Faradays = 96,485 C/mol, but here we only calculate the number of Faradays.)

10. In an electrolysis experiment, a current of 0.500 A was passed through a solution of AgNO_3 for 30.0 minutes. The mass of silver deposited on the cathode was found to be 0.503 g. Given that the molar mass of silver is $107.87 \text{ g mol}^{-1}$ and the charge on a silver ion is +1. Calculate the value of Avogadro's number (N_A) from this data.

Ans. Given Data:

$$\text{Current (I)} = 0.500 \text{ A}$$

$$\text{Time (t)} = 30.0 \text{ minutes} = 1800 \text{ seconds}$$

$$\text{Mass of Ag deposited (m)} = 0.503 \text{ g}$$

$$\text{Molar mass of Ag (M)} = 107.87 \text{ g mol}^{-1}$$

$$\text{Change on } \text{Ag}^+ \text{ ion (n)} = 1$$

Step 1: Calculate the total charge passed (Q)

$$Q = I \times t = 0.500 \text{ A} \times 1800 \text{ s} = 900 \text{ C}$$

Step 2: Determine moles of Ag deposited

$$\text{Moles of Ag} = \frac{m}{M} = \frac{0.503g}{107.87 \text{ g mol}^{-1}} = 4.663 \times 10^{-3} \text{ mol}$$

Step 3: Related charge to moles of electrons (Faradays law)

Since $\text{Ag}^+ + e^- \rightarrow \text{Ag}$, 1 mole of Ag requires 1 mole of electrons (1 Faraday = 96,485 C).

Total charge required for $4.663 \times 10^{-3} \text{ mol Ag}$.

$$Q_{\text{theory}} = \text{Moles of Ag} \times F = 4.663 \times 10^{-3} \times 96,485 \text{ C}$$

But the actual charge used (Q) = 900 C, so:

$$F = \frac{Q}{\text{Moles of Ag}} = \frac{900}{4.633 \times 10^{-3}} = 1.930 \times 10^5 \text{ C mol}^{-1}$$

Step 4: Calculate Avogadro's number (N_A)

Since $F = N_A \times e$ (where $e = 1.602 \times 10^{-19} \text{ C}$):

$$N_A = \frac{F}{e} = \frac{1.930 \times 10^5}{1.602 \times 10^{-19}} = 1.2 \times 10^{24}$$

Final Answer: $N_A = 1.20 \times 10^{24} \text{ mol}^{-1}$

(Note: Due to error in statement of question the answer comes 1.20×10^{24} molecules, however it should be 6.02×10^{23}).

11. A cell is set up with a standard nickel electrode ($\text{Ni}_{(\text{aq})}^{2+} + 2e^- \rightleftharpoons \text{Ni}_{(\text{s})}$ $E^\circ = -0.25 \text{ V}$) and a standard cobalt electrode ($\text{Co}_{(\text{aq})}^{2+} + 2e^- \rightleftharpoons \text{Co}_{(\text{s})}$ $E^\circ = -0.28 \text{ V}$).

i) Identify which metal will be the anode and which will be the cathode. Justify your

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answer.

- ii) Write the balanced overall cell reaction.
 iii) Calculate the standard cell potential (E_{cell}°).

Ans. Solution:

i) **Identifying Anode and Cathode**

• **Standard Reduction Potentials (E°):**

- $\text{Ni}^{2+} + 2e^{-} \rightleftharpoons \text{Ni}_{(s)} \quad E^{\circ} = -0.25 \text{ V}$
- $\text{Co}^{2+} + 2e^{-} \rightleftharpoons \text{Co}_{(s)} \quad E^{\circ} = -0.28 \text{ V}$

• **Comparison:**

- Nickel has a less negative (higher) reduction potential than cobalt.
- **Cathode (Reduction):** The electrode with the higher reduction potential (Ni^{2+}/Ni).
- **Anode (Oxidation):** The electrode with the lower reduction potential (Co^{2+}/Co), where oxidation occurs.

Reason:

Since, $E_{(\text{Ni}^{2+}/\text{Ni})}^{\circ} > E_{(\text{Co}^{2+}/\text{Co})}^{\circ}$, Ni^{2+} gains electrons (reduction) at the cathode, while Co less electrons (oxidation) at the anode.

ii) **Balanced Overall Cell Reaction**

- **Oxidation at Anode (Co):** $\text{Co}_{(s)} \rightarrow \text{Co}^{2+} + 2e^{-}$
- **Reduction at Cathode (Ni):** $\text{Ni}_{(aq)}^{2+} + 2e^{-} \rightleftharpoons \text{Ni}_{(s)}$
- **Overall Reaction:** $\text{Ni}_{(aq)}^{2+} + 2e^{-} \rightarrow \text{Ni}_{(s)}$

iii) **Standard Cell Potential (E°_{cell})**

- $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$
- $E_{\text{cell}}^{\circ} = (-0.25 \text{ V}) - (-0.28 \text{ V}) = +0.03 \text{ V}$
- The positive E_{cell}° confirms the reaction is spontaneous.
- Cobalt (anode) dissolves, while nickel (cathode) deposits.

Final Answer:

- i) **Anode:** $\text{Co}_{(s)}$ (oxidation) ii) $\text{Ni}_{(aq)}^{2+} + \text{Co}_{(s)} \rightarrow \text{Ni}_{(s)} + \text{Co}^{2+}$ iii) $E_{\text{cell}}^{\circ} = +0.03 \text{ V}$
Cathode: $\text{Ni}_{(aq)}^{2+}$ (reduction)