UNIT-3: ELECTROCHEMISTRY

Electrochemistry

Electro (Electrical Energy)

Chemistry (Chemical Energy)

Definition:

A branch of chemistry that deals with the conversion of electrical energy into chemical energy and vice versa is known as electrochemistry.

Thus, electrochemistry deals with cell reactions.

Cell

1. Electrochemical Cell

A device that converts chemical energy into electrical energy.

Examples: Dry cell, mercury cell.

2. Electrolytic Cell

A device through which chemical reactions are carried out with the help of electrical energy.

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Oxidation and Reduction:

1. Oxidation is the process where an atom, ion, or molecule loses electrons. In this process, the oxidation state of the substance increases.

2. Reduction is the opposite process, where an atom, ion, or molecule gains electrons. The oxidation state of the substance decreases.

These two processes always occur simultaneously in a redox (reduction-oxidation) reaction. One substance is oxidized (loses electrons), and the other is reduced (gains electrons).

Oxidation Potential and Reduction Potential:

1. Oxidation Potential refers to the tendency of a substance to lose electrons (i.e., to be oxidized). A higher oxidation potential means a substance is more likely to lose electrons and undergo oxidation.

2. Reduction Potential (also called electrode potential) is the tendency of a substance to gain electrons (i.e., to be reduced). It is measured under standard conditions (25°C, 1 M concentration for solutions, and 1 atm pressure for gases) and is usually reported in volts (V).

Examples:

- $Zn^{2+} + 2e^- \rightarrow Zn$ has a reduction potential of -0.76 V, meaning zinc is more likely to be oxidized than reduced.

- $Cu^{2+} + 2e^- \rightarrow Cu$ has a reduction potential of +0.34 V, meaning copper is more likely to be reduced than oxidized.

Difference between metallic and electrolytic conductor:

Metallic Conductors:

1. Conduction occurs due to the flow of free electrons in the material.
2. Metallic conductors are typically made of metals such as copper, aluminum, silver, and gold.
3. The conduction medium is solid metal.
4. Direction of Current: The current flows through metallic conductors due to the movement of electrons from the negative terminal to the positive terminal.

Electrolytic Conductors:

1. Nature of Conduction: Conduction occurs due to the movement of ions (positively charged cations and negatively charged anions) in a solution or molten state.
2. Material Type: Electrolytic conductors are usually ionic compounds dissolved in water or molten salts (e.g., NaCl in water, sulfuric acid).

3. Conduction Medium: The conduction medium is a liquid or molten ionic solution.

4. Direction of Current: In electrolytic conduction, positive ions move towards the cathode (negative terminal), and negative ions move towards the anode (positive terminal).

Here's a clear and expanded explanation of the content from the pictures with added details for a better understanding.

Electrochemistry

Electrochemistry is a branch of chemistry that studies the relationship between electrical energy and chemical reactions. It primarily focuses on how electrical energy can be used to drive chemical reactions and vice versa.

In electrochemistry:

Electrical Energy can convert into Chemical Energy.

Chemical Energy can convert into Electrical Energy.

This area of chemistry deals with different types of cells that can either produce or use electricity, such as electrochemical cells.

Types of Cells in Electrochemistry

There are two main types of cells in electrochemistry:

1. Electrochemical Cell (Galvanic/Daniell Cell)

Converts chemical energy into electrical energy.

Used in batteries where a chemical reaction produces electricity.

Examples: Dry cell, Mercury cell.

2. Electrolytic Cell

Uses electrical energy to drive a chemical reaction.

Commonly used in processes like electroplating, refining metals, and water splitting.

Example reaction:

The Electrochemical Cell (Daniell Cell)

A Daniell Cell is an example of an electrochemical cell where chemical reactions generate electricity.

Components of a Daniell Cell:

Anode (-): Zinc (Zn) – where oxidation occurs (loss of electrons).

Reaction at Anode:

Cathode (+): Copper (Cu) – where reduction occurs (gain of electrons).

Reaction at Cathode:

Salt Bridge: A U-shaped tube filled with an inert electrolyte like KCl or KNO_3.

Role of the Salt Bridge:

Maintains electrical neutrality by allowing the flow of ions between the two half-cells.

Prevents the build-up of charges that would stop the cell from functioning.

In this cell, zinc undergoes oxidation (loses electrons), and copper ions undergo reduction (gain electrons), creating a flow of electrons from zinc to copper, which generates an electric current.

Key Concepts in Electrochemistry

1. Conductors and Insulators

Conductor: A substance that allows electricity to pass through it.

Examples: Silver (Ag), Sodium Chloride (NaCl).

Insulator: A substance that does not allow electricity to pass through it.

Examples: Wood, Rubber.

2. Ionization

The process of breaking down a substance into charged particles called ions (cations and anions).

Example: When NaCl dissolves in water, it separates into Na^+ (cation) and Cl^- (anion).

3. Electrolytic Dissociation

The process where electrolytes break down into ions when dissolved in water or melted.

Essential for conducting electricity in solutions, as ions carry the current.

Arrhenius Theory of Electrolytic Dissociation

Electrolytes are substances that can conduct electricity in an aqueous solution or molten state by dissociating into ions. Based on their ability to dissociate, electrolytes are classified into two categories:

1. Strong Electrolytes

Completely dissociate into ions in solution, allowing a strong current to pass.

Examples: Strong acids (e.g., HCl), Strong bases (e.g., NaOH), Salts (e.g., NaCl).

2. Weak Electrolytes

Partially dissociate into ions, resulting in a weaker current.

Examples: Weak acids (e.g., CH_3COOH or acetic acid), Weak bases (e.g., NH_4OH).

In summary, strong electrolytes are highly effective in conducting electricity, while weak electrolytes only conduct a limited amount of current due to partial ionization.

Applications of Electrochemistry

1. Batteries: Use electrochemical cells to store and provide electrical energy for devices.

2. Electroplating: Electrolytic cells are used to coat metals with a thin layer of another metal, like gold or silver.

3. Water Splitting: Electrolytic cells break down water into hydrogen and oxygen gases, an important process in producing clean fuels.

Electrochemistry plays a crucial role in modern technology, enabling the development of batteries, fuel cells, metal refining, and various industrial processes. Understanding these fundamental principles can help in harnessing electrical and chemical energy efficiently.

The differences between strong electrolytes and weak electrolytes:

1. Definition

Strong Electrolyte: A substance that completely dissociates into ions when dissolved in water or melted, resulting in high electrical conductivity.

Weak Electrolyte: A substance that only partially dissociates into ions in water or molten state, resulting in low electrical conductivity.

2. Degree of Dissociation

Strong Electrolyte: Nearly 100% of the molecules dissociate into ions.

Weak Electrolyte: Only a small fraction of molecules dissociate into ions, while the rest remain as undissociated molecules.

3. Examples

Strong Electrolytes: Strong acids (like HCl, H_2SO_4), strong bases (like NaOH, KOH), and salts (like NaCl, KNO_3).

Weak Electrolytes: Weak acids (like CH_3COOH, HF) and weak bases (like NH_4OH).

4. Electrical Conductivity

Strong Electrolyte: Conducts electricity very well due to the high concentration of ions.

Weak Electrolyte: Conducts electricity poorly because it produces fewer ions in solution.

5. Ion Concentration in Solution

Strong Electrolyte: High ion concentration because almost all molecules are ionized.

Weak Electrolyte: Low ion concentration as only a portion of the molecules ionize.

6. pH Level (for Acids and Bases)

Strong Electrolyte (Acids/Bases): Strong acids have a very low pH (closer to 1), and strong bases have a high pH (closer to 14).

Weak Electrolyte (Acids/Bases): Weak acids have a pH slightly below 7, and weak bases have a pH slightly above 7.

7. Reversible Ionization

Strong Electrolyte: Ionization is not reversible, as the molecules dissociate completely.

Weak Electrolyte: Ionization is reversible, and an equilibrium is established between ions and undissociated molecules.

8. Effect on Colligative Properties

Strong Electrolyte: Causes significant changes in colligative properties (e.g., boiling point elevation, freezing point depression) because it produces more ions.

Weak Electrolyte: Has a lesser effect on colligative properties due to fewer ions.

9. Examples in Everyday Life

Strong Electrolyte: Table salt (NaCl) in water, which fully dissociates to help conduct electricity.

Weak Electrolyte: Vinegar (acetic acid solution) in water, which only partially dissociates and is a poor conductor.

10. Use in Electrochemical Cells

Strong Electrolyte: Often used in electrochemical cells and batteries because they ensure efficient ion flow and good conductivity.

Weak Electrolyte: Less commonly used in electrochemical cells as it provides less efficient conductivity.

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1. Electrolytes and Their Dissociation:

- Strong electrolytes (like NaCl, KOH, H2SO4) dissociate completely in water, meaning almost 100% of the molecules break into ions

- Weak electrolytes (like CH3COOH, NH4OH) only partially dissociate, with only a fraction of molecules breaking into ions

- Example: NaCl (strong) → Na+ + Cl- (complete dissociation)

- Example: CH3COOH (weak) ⇌ CH3COO- + H+ (partial dissociation)

2. Degree of Ionization (α):

- It's a crucial measurement in electrochemistry
- Expressed as a fraction or percentage
- Formula: α = (Number of molecules ionized)/(Total number of molecules taken)
- For strong electrolytes, $\alpha \approx 1$ (or 100%)
- For weak electrolytes, $\alpha < 1$ (less than 100%)

4. Factors Affecting Degree of Ionization:

a) Nature of Solute:

- Strong electrolytes: Have strong ionic or covalent bonds that break easily in solution
- Weak electrolytes: Have weaker bonds and reach an equilibrium between ionized and non-ionized forms
- The strength of the electrolyte depends on its chemical structure and bond types

b) Nature of Solvent:

- Polar solvents (like water) promote ionization better than non-polar solvents
- The dielectric constant of the solvent plays a crucial role
- Higher dielectric constant means better ionization
- Water is an excellent solvent due to its high dielectric constant and ability to form hydrogen bonds

c) Concentration of Solution:

- Dilution increases the degree of ionization
- This follows Le Chatelier's Principle
- At infinite dilution, weak electrolytes behave like strong electrolytes
- Higher concentration leads to more ion-ion interactions, which can reduce ionization
- This is known as the "dilution effect"

Additional Important Points:

1. Temperature Effects:

- Generally, increasing temperature increases ionization

- This is because more energy is available to break bonds

- However, the effect varies for different electrolytes

2. Common Ion Effect:

- The presence of a common ion decreases ionization

- This is also based on Le Chatelier's Principle

- Example: Adding CH_3COONa to CH_3COOH reduces the ionization of acetic acid

3. Practical Applications:

- Understanding ionization is crucial in:
  * Buffer solutions
  * pH calculations
  * Chemical analysis
  * Industrial processes
  * Biological systems

4. Conductivity Relationship:

- The degree of ionization directly affects electrical conductivity

- More ionization means better conductivity

- This principle is used in conductometric titrations

5. Measurement Methods:

- Degree of ionization can be measured by:

  * Conductivity measurements

  * pH measurements (for acids and bases)

  * Colligative property measurements

  * Spectroscopic methods

(9) Temperature:-

Nigher is the temperafore highes will be the degree of ionisation because molecular velocities are increased and the force of attraction between the ton decreases

Electrolysis

Electro Electroicity Lysis breaking of bond

Electrolysis! The process of compositi chemical decomposition of an electr olyte by the passage of electrit current through it is known as electrolysip. + Electrolytic cell.

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ELECTROCHEMICAL SERIES

The Electrochemical Series is a ranking of elements (usually metals) and their ions in order of their standard electrode potentials (E°). Here's a comprehensive breakdown:

1. Standard Electrode Potential (E°)

- Measured in volts (V)
- More positive E° = better oxidizing agent
- More negative E° = better reducing agent

Example values (from most negative to most positive):

```
$Li^+/Li$    E° = -3.04 V
$K^+/K$     E° = -2.92 V
$Na^+/Na$   E° = -2.71 V
$Zn^{2+}/Zn$  E° = -0.76 V
$Fe^{2+}/Fe$  E° = -0.44 V
$Cu^{2+}/Cu$  E° = +0.34 V
$Ag^+/Ag$   E° = +0.80 V
```

3. Half-Reactions

For any metal M with charge n+:

- Reduction: M^n+ + ne^- → M

- Oxidation: M → M^n+ + ne^-

Example with Zinc:

- Reduction: $Zn^{2+} + 2e^- \rightarrow Zn$
- Oxidation: $Zn \rightarrow Zn^{2+} + 2e^-$

4. Practical Applications:

Let's work through a complete example of copper and zinc:

When you put zinc metal in copper sulfate solution:

```
Zn(s) + Cu²⁺(aq) → Zn²⁺(aq) + Cu(s)
```

This reaction occurs spontaneously because:

- Zn has E° = -0.76 V

- Cu has E° = +0.34 V
- The difference (ΔE°) = +0.34 – (-0.76) = +1.10 V
- Positive ΔE° means spontaneous reaction

5. Cell Potential Calculation:

E°cell = E°cathode – E°anode

Example for Zn-Cu cell:

```

E°cell = E°Cu²⁺/Cu – E°Zn²⁺/Zn

E°cell = (+0.34 V) – (-0.76 V)

E°cell = +1.10 V
```

5. Predicting Reactions:

- A metal will displace any ion below it in the series
- Example: Zinc will displace copper ions but not potassium ions

Practical Example:

Iron nail in copper sulfate solution:

```

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

$E°cell = (+0.34\ V) – (-0.44\ V) = +0.78\ V$

```

This reaction occurs because:

- Fe is higher in the series (more negative)
- Positive E°cell indicates spontaneous reaction
- You'll observe copper metal plating on the iron nail

RE-ACTIVITY SERIES

The Activity Series is a list of metals (and hydrogen) arranged in order of decreasing reactivity based on their ability to displace other elements from compounds.

Here's the complete Activity Series from most reactive (top) to least reactive (bottom):

` ` `

Most Reactive

Lithium (Li)
Potassium (K)
Barium (Ba)
Calcium (Ca)
Sodium (Na)
Magnesium (Mg)
Aluminum (Al)
Zinc (Zn)
Chromium (Cr)
Iron (Fe)
Nickel (Ni)
Tin (Sn)
Lead (Pb)

[Hydrogen (H)]

Copper (Cu)

Mercury (Hg)

Silver (Ag)

Platinum (Pt)

Gold (Au)

Least Reactive

```

Key Rules and Applications:

1. Displacement Reactions:

- Any metal will displace all metals below it in the series

Example:

```

$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

(Zinc displaces copper from copper sulfate)

```
```

2. Reaction with Water:

- Metals above hydrogen react with cold water

Example:

```

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

(Sodium reacts vigorously with water)

```

3. Reaction with Steam:

- Metals from Mg to Pb react with steam

Example:

```

$Mg + H_2O(steam) \rightarrow MgO + H_2$

```

4. Reaction with Acids:

- Metals above hydrogen react with acids

Example:

```
Zn + 2HCl → ZnCl₂ + H₂
```

6. Practical Applications:

a) Protecting Metals (Sacrificial Protection):

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More reactive metal protects less reactive metal
Example: Zinc coating on iron (galvanization)
```

b) Metal Extraction:

```
More reactive metals require electrolysis
Less reactive metals can be extracted by reduction
```

c) Storage Conditions:

```

Reactive metals (Na, K) – stored in oil

Moderately reactive (Fe) – can be stored in air

Less reactive (Cu, Ag) – can be found as native metals

```

Common Reactions to Remember:

1. With Oxygen:

```

$4Na + O_2 \rightarrow 2Na_2O$

$2Mg + O_2 \rightarrow 2MgO$

$4Al + 3O_2 \rightarrow 2Al_2O_3$

```

2. With Water:

```

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

$Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$

```

3. With Acids:

```

$Zn + 2HCl \rightarrow ZnCl_2 + H_2$

$Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$

```

Important Notes:

1. The more reactive a metal:
   - More readily it loses electrons
   - More easily it forms compounds
   - More difficult to extract from its compounds

2. The less reactive a metal:

- More stable in pure form
- Less likely to corrode
- Easier to extract from compounds

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Nernst Equation

The Nernst equation is fundamental in electrochemistry for calculating cell potential under non-standard conditions.

1. Basic Form of Nernst Equation:

```

E = E° - (RT/nF)ln(Q)


Simplified form at 298K (25°C):

E = E° - (0.0592/n)log_10(Q)
```

2. Components Breakdown:

```

E = Actual cell potential (volts)

E° = Standard cell potential (volts)

R = Gas constant (8.314 J/mol·K)

T = Temperature (Kelvin)

N = Number of electrons transferred

F = Faraday constant (96,485 C/mol)

Q = Reaction quotient
```

3. Calculating Reaction Quotient (Q):

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For a reaction: $aA + bB \rightarrow cC + dD$

$Q = [C]^c[D]^d/[A]^a[B]^b$

Note:

- Use molar concentrations for solutions
- Pure solids and liquids = 1
- For gases, use partial pressures

```

## 4. Detailed Examples:

Example 1: Simple Cell

```

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

Given:

$E° = 1.10V$

$[Zn^{2+}] = 0.15M$

$[Cu^{2+}] = 0.05M$

$N = 2$

$Q = [Zn^{2+}]/[Cu^{2+}] = 0.15/0.05 = 3$

$E = 1.10 – (0.0592/2)\log(3)$

$E = 1.10 – 0.0296\log(3)$

$E = 1.10 – 0.0142$

$E = 1.086V$

` ` `

Example 2: pH-Dependent Reaction

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$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

$E = E° - (0.0592/5)\log([Mn^{2+}]/[MnO_4^-][H^+]^8)$

` ` `

5. Temperature Effects:

At different temperatures:

` ` `

$E = E° - (RT/nF)\ln(Q)$

For $T \neq 298K$:

Factor = $(RT/F) = (8.314 \times T)/(96{,}485)$

` ` `

6. Common Applications:

a) pH Measurements:

` ` `

For hydrogen electrode:

$E = E° - (0.0592)\log(1/[H^+])$

$E = E° + 0.0592 \times pH$

` ` `

b) Concentration Cells:

` ` `

$E = (0.0592/n)\log(C_1/C_2)$

Where C_1 and C_2 are concentrations

```
```

7. Practical Tips:

a) Signs Matter:

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- Negative log term increases potential

- Positive log term decreases potential

```

b) Concentration Effects:

```

- Higher product concentration → Lower E

- Higher reactant concentration → Higher E

```

8. Step-by-Step Problem Solving:

1. Write balanced equation
2. Identify E° for the cell
3. Count electrons transferred (n)
4. Write expression for Q
5. Insert values into Nernst equation
6. Calculate final answer

Example Problem:

```

$Fe^{3+} + Sn^{2+} \rightleftharpoons Fe^{2+} + Sn^{4+}$

Given:

$E° = 0.77V$

$[Fe^{3+}] = 0.2M$

$[Fe^{2+}] = 0.1M$

$[Sn^{2+}] = 0.3M$

$[Sn^{4+}] = 0.1M$
```

Step 1: $Q = [Fe^{2+}][Sn^{4+}]/[Fe^{3+}][Sn^{2+}]$

Step 2: $Q = (0.1)(0.1)/(0.2)(0.3) = 0.167$

Step 3: $E = 0.77 – (0.0592/1)\log(0.167)$

Step 4: $E = 0.77 + 0.0534$

Step 5: $E = 0.823V$

```

9. Important Notes:

1. Always check units:

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- Concentrations in mol/L

- Pressures in atm

- Temperature in Kelvin
```

2. Common Mistakes:
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```
- Forgetting to convert temperature to Kelvin
- Using natural log instead of log$_{10}$
- Incorrect reaction quotient setup
```

Here’s a detailed explanation of primary and secondary batteries, along with their differences, and an explanation of a labeled dry cell diagram.

Battery

Batteries are devices that store chemical energy and convert it into electrical energy. They are used to power various devices, from small electronics to vehicles. Batteries are generally divided into two main types:

1. Primary Battery: Non-rechargeable, single-use batteries.

2. Secondary Battery: Rechargeable batteries that can be used multiple times.

Primary Battery

Definition: A primary battery is a type of battery that is designed to be used once and then discarded. It cannot be recharged.

Chemical Reaction: The chemical reactions in a primary battery are irreversible. Once the chemicals are depleted, the battery stops working.

Examples: Alkaline batteries, zinc-carbon batteries, and lithium batteries (used in small electronics like remote controls and flashlights).

Key Features of Primary Batteries:

Single-use, non-rechargeable.

Generally have a longer shelf life, as they do not lose charge quickly when not in use.

Often used in low-drain devices or devices used infrequently.

Secondary Battery

Definition: A secondary battery is a type of battery that can be recharged and used multiple times. It is designed to be connected to an external power source to restore its charge.

Chemical Reaction: The chemical reactions in a secondary battery are reversible, allowing it to be recharged and used repeatedly.

Examples: Lithium-ion batteries (used in mobile phones and laptops), lead-acid batteries (used in cars), and nickel-cadmium batteries.

Key Features of Secondary Batteries:

Rechargeable and reusable.

Ideal for high-drain applications.

Can be used in devices that require frequent charging, like smartphones, laptops, and electric vehicles.

Differences Between Primary and Secondary Batteries

Dry Cell (Primary Battery) and Its Diagram

A common type of primary battery is the dry cell, specifically the zinc-carbon dry cell. Below is a breakdown of the labeled parts and their functions.

Diagram of a Dry Cell:

[Diagram here would include the following parts, labeled]

1. Outer Zinc Container (Anode): Acts as the negative terminal and also serves as the container for the battery's contents. It participates in the electrochemical reaction by releasing electrons.

2. Graphite Rod (Cathode): Located at the center, the graphite rod serves as the positive terminal but does not participate in the reaction. It collects electrons from the chemical reactions and delivers them to the external circuit.

3. Electrolyte Paste: Consists of ammonium chloride (NH_4Cl) and zinc chloride ($ZnCl_2$), which help conduct ions within the cell.

4. Manganese Dioxide (MnO_2) and Carbon Mixture: Surrounds the graphite rod and reacts with the electrolyte to facilitate the flow of electrons.

5. Separator: A porous material that separates the electrolyte paste from the zinc container to prevent direct contact while allowing ion movement.

Explanation of Each Part

1. Outer Zinc Container (Anode):

The zinc container serves as the anode (negative terminal). During the reaction, zinc oxidizes and releases electrons.

The zinc gradually corrodes as the battery discharges, which is why primary batteries cannot be recharged.

2. Graphite Rod (Cathode):

The graphite rod acts as the positive terminal of the battery.

It doesn't participate in the reaction directly but collects electrons from the chemical reaction and provides them to the external circuit.

3. Electrolyte Paste:

Made up of ammonium chloride (NH_4Cl) and zinc chloride ($ZnCl_2$).

The electrolyte enables the flow of ions, which completes the internal circuit and sustains the chemical reaction.

4. Manganese Dioxide (MnO_2) and Carbon Mixture:

Manganese dioxide acts as a depolarizer, preventing the buildup of hydrogen gas, which can interfere with the battery's performance.

It aids In sustaining the flow of electrons by maintaining an efficient reaction at the cathode.

5. Separator:

This porous material prevents direct contact between the zinc container (anode) and the electrolyte paste, which would short-circuit the battery.

It allows the movement of ions through the cell, ensuring the smooth functioning of the chemical reaction.

Working of a Dry Cell

When the dry cell is connected to an external circuit, a chemical reaction occurs:

1. Oxidation at the Anode (Zinc):

Zinc (Zn) atoms at the anode lose electrons (oxidation) and become zinc ions (Zn^{2+}).

2. Reduction at the Cathode (Manganese Dioxide):

The electrons flow through the external circuit to the cathode, where manganese dioxide (MnO_2) reacts with ammonium ions (NH_4^+) from the electrolyte and accepts electrons.

3. Ion Flow in the Electrolyte:

As electrons flow from the anode to the cathode, ammonium ions (NH_4^+) from the electrolyte neutralize the excess electrons at the cathode, maintaining charge balance.

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Q. State the arrenhius theory of electrolytic disociation.

The Arrhenius theory of electrolyte dissociation, proposed by Swedish scientist Svante Arrhenius, explains how certain substances conduct electricity when dissolved in water. According to this theory, electrolytes dissociate into ions when they are dissolved in water, enabling them to conduct electricity. Here's a breakdown of each point in simple terms:

1. Definition of Electrolytes:

Explanation: According to Arrhenius, an electrolyte is a substance that can conduct electricity when dissolved in water or in a molten state.

Example: Salts (like $NaCl$), acids (like HCl), and bases (like $NaOH$) are examples of electrolytes because they can split into ions in water.

Why This Matters: Only substances that dissociate into ions can conduct electricity in a solution.

2. Ionization Process:

Explanation: When an electrolyte dissolves in water, it undergoes ionization or dissociation, meaning it splits into positively and negatively charged ions.

Example: When sodium chloride (NaCl) dissolves, it separates into Na^+ (sodium ions) and Cl^- (chloride ions).

Why This Matters: These ions are what allow the solution to conduct electricity. The positive ions move towards the negative electrode, while the negative ions move towards the positive electrode.

3. Types of Electrolytes:

Explanation: Arrhenius classified electrolytes into strong and weak based on their degree of dissociation.

Strong Electrolytes: Completely dissociate in water, producing a large number of ions. Examples include HCl, NaOH, and KCl.

Weak Electrolytes: Partially dissociate in water, producing fewer ions. Examples include acetic acid (CH_3COOH) and ammonia (NH_3).

Why This Matters: Strong electrolytes are better conductors of electricity than weak electrolytes because they produce more ions in the solution.

4. Conductivity in Solution:

Explanation: Only ions can conduct electricity in a solution because they are charged particles that can move freely.

Why This Matters: This movement of ions towards opposite electrodes creates an electric current, allowing the solution to conduct electricity.

5. Water as a Solvent:

Explanation: According to Arrhenius, water plays a key role in the dissociation of electrolytes. The polar nature of water molecules (having a positive and negative side) helps separate the ions in the electrolyte.

Example: Water molecules surround Na^+ and Cl^- ions when NaCl dissolves, preventing them from recombining.

Why This Matters: Without water (or a similar polar solvent), many electrolytes would not dissociate and therefore wouldn't conduct electricity.

6. Electrolytic Dissociation is a Reversible Process:

Explanation: For weak electrolytes, the dissociation process is reversible, meaning the ions can recombine to form the original molecule.

Example: Acetic acid (CH_3COOH) dissociates into CH_3COO^- and H^+ ions, but these ions can also recombine to form CH_3COOH.

Why This Matters: This reversibility affects the number of free ions available in the solution, impacting conductivity.

7. Degree of Ionization Depends on Concentration and Temperature:

Explanation: The extent to which an electrolyte dissociates (degree of ionization) depends on factors like the concentration of the solution and the temperature.

Concentration: In dilute solutions, electrolytes tend to ionize more completely.

Temperature: Higher temperatures increase the energy of molecules, making them more likely to dissociate.

Why This Matters: These factors affect the conductivity of the solution. Higher ionization means better conductivity.

Defination-

Electroplating is a process that uses electrical current to deposit a layer of metal onto a conductive surface, often for the purposes of enhancing appearance, preventing corrosion, or improving weat resistance. It is commonly used in industries such as electronics, jewelry, automotive, and manufacturing.

Process of Electroplating

1. Preparation of the Object (Substrate)

The item to be plated (usually made of metal or plastic with a conductive coating) must first be cleaned to remove any dirt, oil, rist, or oxidation. This is typically done using chemical cleaners, abrasives, or ultrasonic taths

2. Electrolyte Solution:

An electrolyte solution is prepared, which contains a metal salt of the metal to be plated, along with a suitable acid or bose to facilitate the electroplating reaction. For example, if plating with gold, a solution of gold chloride or gold cyanide may be used

3. Setup of Electroplating Bath:

The item to be plated is connected to the negative terminal (cathode) of a power supply, and the metal electrode (anode), which is the source of the plating metal, is connected to the positive terminal. The electrolyte solution is placed in a bath, and both the object and anode are submerged.

4. Electroplating Reaction

When current is applied, metal cations in the electrolyte are reduced at the cathode (the object being plated), forming a metal coating. The metal from the anode dissolves into the electrolyte to replace the metal being deposited on the object

5. Finishing Touches

After plating, the item may be polished, buffed, or coated with additional lagers to achieve the desired appearance or functionality

Benefits of Electroplating

Corrosion Resistance:

Electroplating can provide a protective coating that prevents rust and corrosion on metals, extending the lifespan of components exposed to harsh environments

3. Aesthetic Appeat

Electroplating can enhance the visual appearance of products, especially in jewelry and decorative items, by providing a shing, attractive surface with a lager of precious metals like gold, silver, or platinum.

4. Improved Durability:

The plating process can increase the hardness and wear resistance of the surface, making it ideal for components subject to frequent friction, such as machine parts or automotive parts

5. Reduced Friction

Electroplated coatings, such as those made of chrome or nickel, can help reduce friction, making moving parts

In machinery operate more smoothly.

4. Cost-Effective

Electroplating allows for the application of precious metals, such as gold or silver, without the need for using large amounts of expensive material. This makes it more economical compared to solid metal manufacturing

6 Electrical Conductivity

Electroplating is used in electronics to improve the conductivity and performance of electrical components like connectors, circuits, and semiconductors.

6. Customization:

Electroplating can be used to customize or modify the properties of a material. For example, a product can be plated with metals like copper for conductivity, silver for better oxidation resistance, or gold for luxueg and corrosion resistance.