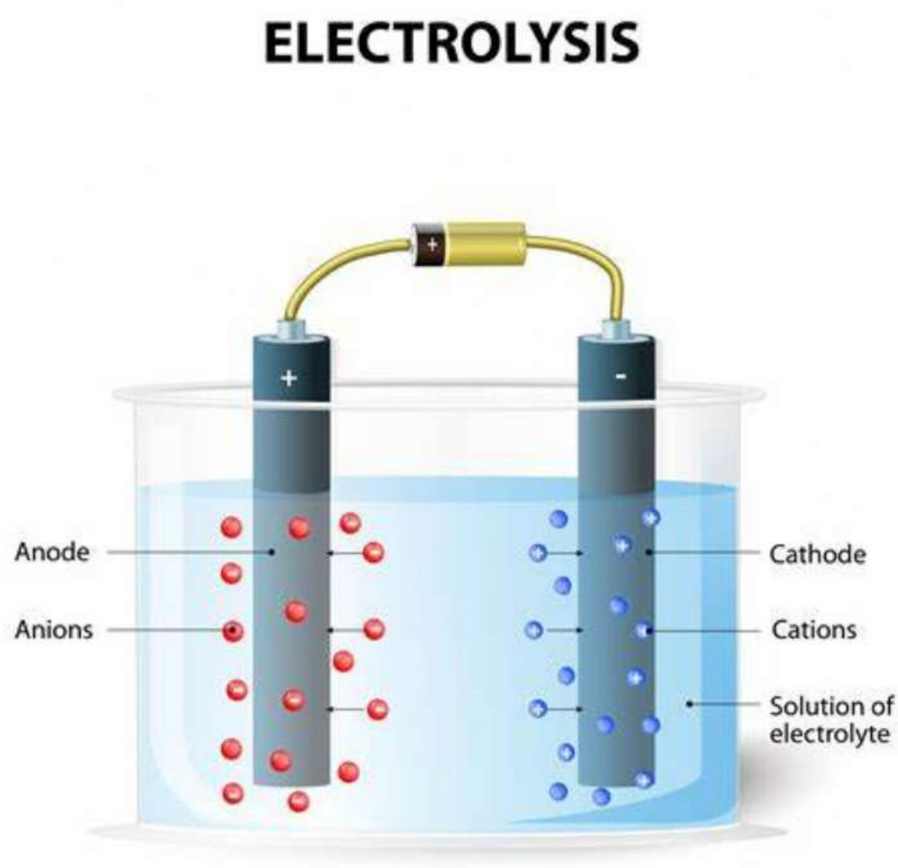
Electrolysis of Molten Compounds and Aqueous Solutions

Electrolysis

Definition: Electrolysis is a process by which electrical energy is used to cause a non-spontaneous chemical reaction. It typically involves the breaking down of compounds into their elements.

Setup: Requires a power source, two electrodes (anode and cathode), and an electrolyte (the substance to be electrolyzed).



Electrodes:

Anode: Positive electrode where oxidation occurs (loss of electrons).

Cathode: Negative electrode where reduction occurs (gain of electrons).

Electrolysis of Molten Compounds

Process:

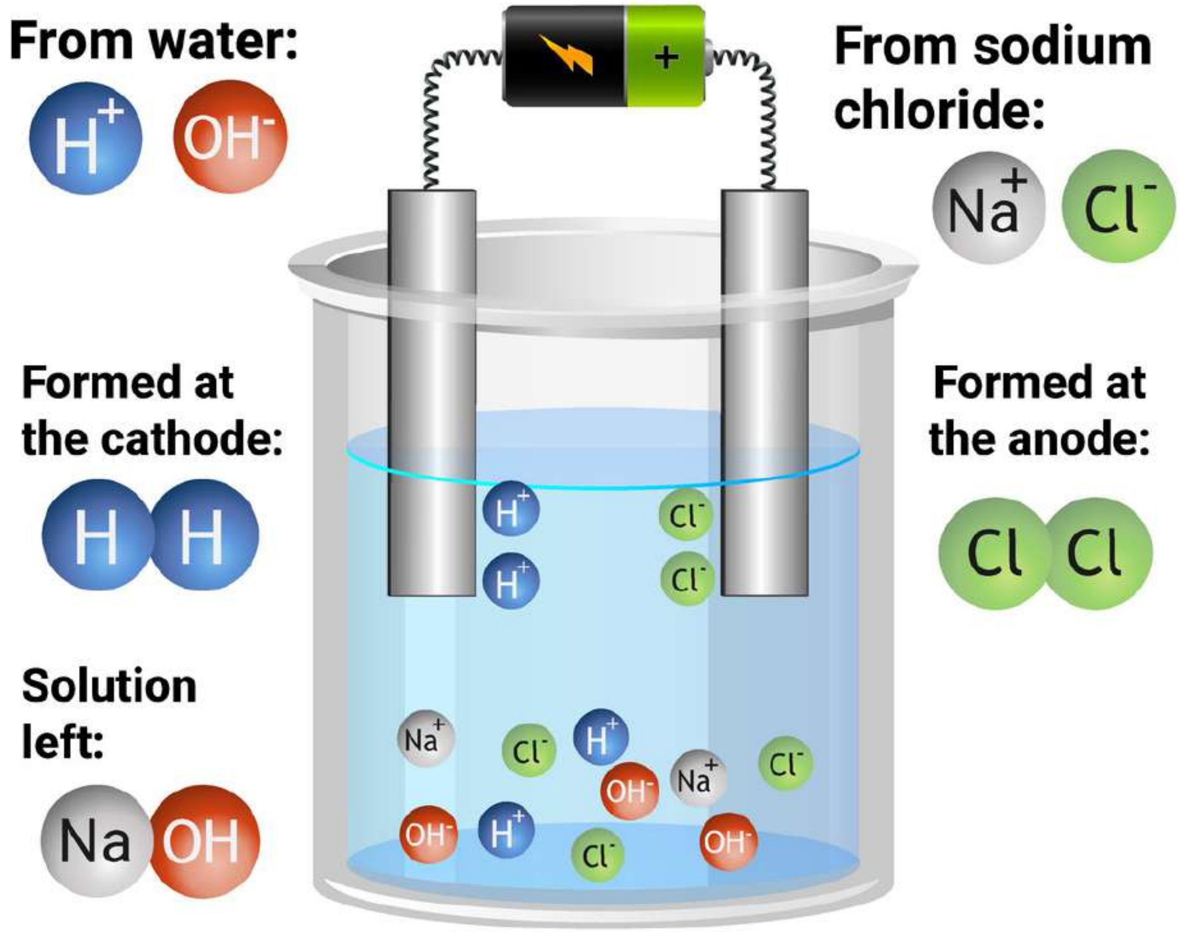
When ionic compounds are melted, their ions are free to move.

Passing an electric current through the molten compound allows cations to move to the cathode and anions to the anode.

Example: Electrolysis of molten sodium chloride (NaCl).

Reaction at the Cathode: ￼ (sodium metal forms).

Reaction at the Anode: ￼ (chlorine gas forms).



Applications: Used to obtain pure metals from ionic compounds, such as aluminum from aluminum oxide.

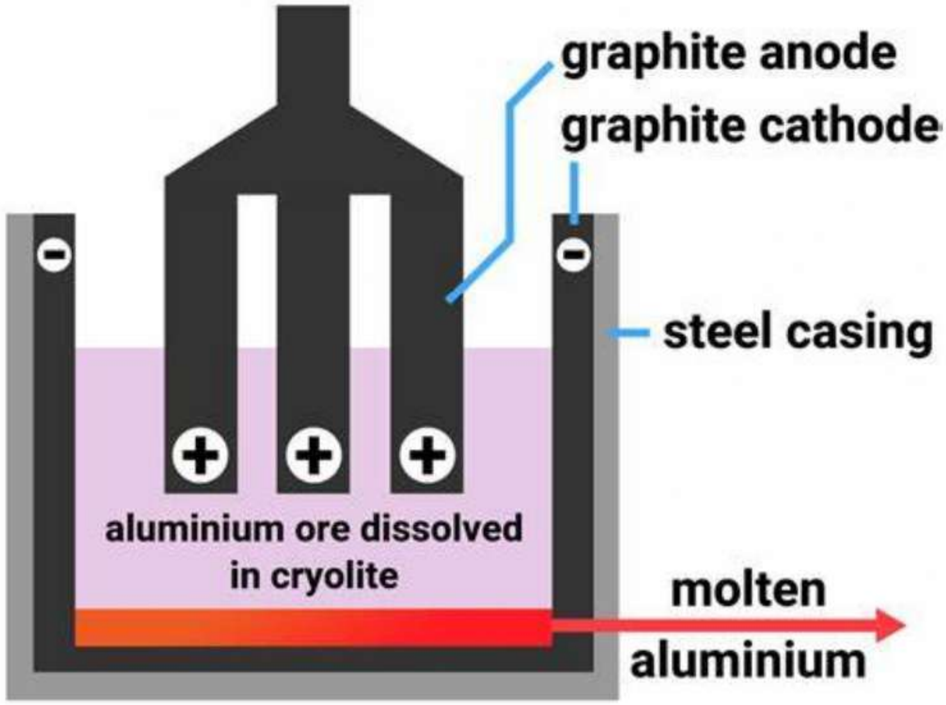
Electrolysis of Aqueous Solutions

Additional Factors: In aqueous solutions, water can also undergo electrolysis, producing hydrogen and oxygen gases.
Electrolysis of Copper(II) Sulfate ($CuSO_4$):
Cathode Reaction: ￼ (copper metal deposits).
Anode Reaction: ￼ (oxygen gas is released).
Selective Discharge: Depending on ion concentration and electrode type, different ions may be preferentially discharged.

Applications of Electrolysis

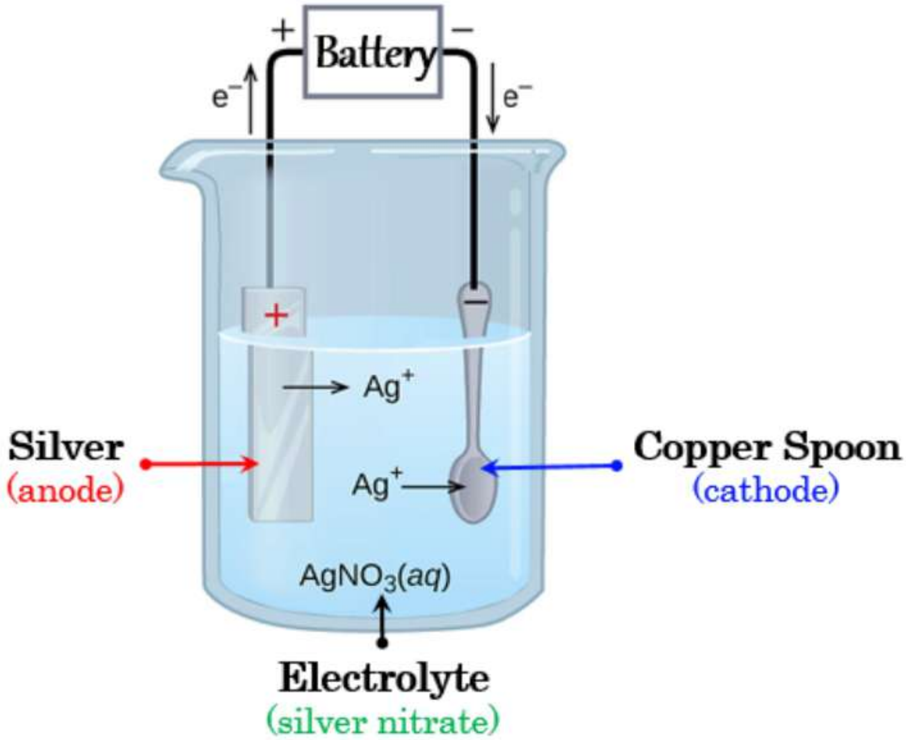
1. Extraction of Metals
Purpose: Electrolysis is used to extract reactive metals from their ores, especially metals that cannot be extracted by traditional methods like reduction with carbon.
Example: Aluminum Extraction from Bauxite (Al_2O_3)

Process:
Bauxite is purified to obtain aluminum oxide.
The aluminum oxide is then electrolyzed in molten cryolite to reduce melting points and save energy.
Aluminum ions move to the cathode, where they are reduced to aluminum metal, and oxygen ions move to the anode.



Reaction:
At the Cathode: ￼ (aluminum metal).
At the Anode: ￼ (oxygen gas).
Other Metals: Electrolysis is also used for metals like magnesium, sodium, and lithium.

2. Electroplating
Purpose: Electroplating is a process of coating a metal surface with a thin layer of another metal to improve appearance, prevent corrosion, or enhance surface properties.
Process:
The item to be plated is used as the cathode, and the metal to be plated onto it is used as the anode.
An electrolyte containing ions of the plating metal is used.
When electric current passes through, the metal ions are reduced at the cathode, coating the item.
Example:Silver Electroplating
Cathode: Object to be plated (e.g., a spoon).
Anode: Silver rod.
Electrolyte: Silver nitrate ($AgNO_3$) solution.
At the Cathode: ￼ (silver metal deposits on the object).
At the Anode: Silver dissolves to release ￼ ions, maintaining the electrolyte concentration.



Applications: Used for jewelry, cutlery, electronic components, and in corrosion-resistant coatings.

Electrochemical Cells and Electrode Reactions

Electrochemical Cells

Definition: Electrochemical cells convert chemical energy into electrical energy using redox (oxidation-reduction) reactions.

Types:

Galvanic (Voltaic) Cells: Produce electricity from spontaneous redox reactions.

Electrolytic Cells: Use electricity to drive non-spontaneous reactions (e.g., electrolysis).

Galvanic Cells

Structure: Consists of two electrodes (anode and cathode) in separate solutions connected by a salt bridge.

Anode: Where oxidation occurs, releasing electrons.

Cathode: Where reduction occurs, accepting electrons.

Example: Zinc-Copper Galvanic Cell

Anode: Zinc electrode in zinc sulfate solution, where zinc oxidizes.

Reaction: .

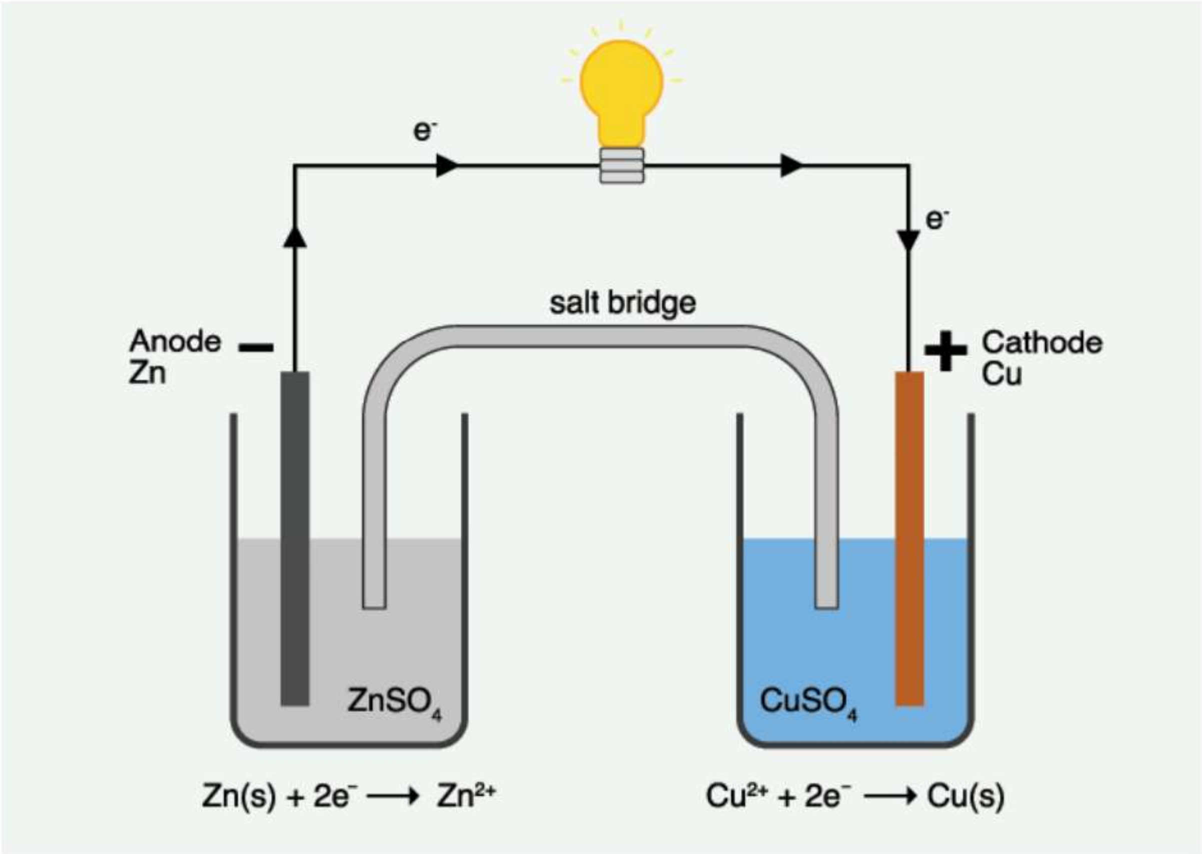
Cathode: Copper electrode in copper sulfate solution, where copper ions are reduced.

Salt Bridge: Completes the circuit and maintains charge balance by allowing ions to move between the two solutions.

Output: Generates a potential difference (voltage) that can power external circuits.

Electrode Reactions

Redox Reactions: Redox reactions are key to understanding electrochemical cells, as they involve electron transfer.



Oxidation: Loss of electrons at the anode.

Reduction: Gain of electrons at the cathode.

Applications of Galvanic Cells:

Batteries and cells are common examples, where chemical energy is converted into electrical energy, as seen in dry cells and lithium-ion batteries.

Conclusion

Understanding the interaction between electricity and chemistry provides insight into how electrical energy can drive chemical processes and vice versa. Electrolysis is essential for extracting metals, refining materials, and creating electroplated products, while electrochemical cells harness chemical reactions to generate electricity, fueling countless devices and applications. Mastery of these principles is foundational for advanced studies in electrochemistry, industrial applications, and the development of sustainable energy sources.