Electrochemistry

Thursday, May 22, 2025 7:31 PM

Definition:

Electrochemistry is the branch of chemistry that deals with the study of interconversion of electrical and chemical energies inside the cell.

Electrical energy ⇌ Chemical energy

⇒ Electrochemical Cell *(Device/Instrument)*

Definition:

A device/instrument that converts chemical energy into electrical energy and vice versa is called an electrochemical cell.

Types of Electrochemical Cells:

1. Electrochemical Cell

   (Electric ⟶ Chemical):

   - A cell that uses electrical energy to drive a chemical reaction.

2. Galvanic Cell

   (Chemical ⟶ Electrical):

   - A cell that converts chemical energy into electrical energy.

⇒ Nature of Electrochemical Process

- Electrochemical processes are redox reactions where:
  - Chemical energy released by a spontaneous reaction is converted to electricity.

    OR
  - Electrical energy is used to drive a non-spontaneous reaction.
  - In either case, the process involves electron transfer from one substance to another.
  - Whether this process release or absorb it always involves the transfer of electrons from one substance to other.

⇒ Redox

Definition:

A reaction in which oxidation and reduction occur simultaneously is called a redox reaction.

Example:

$Na + Cl_2 \rightarrow NaCl$

⇒ Oxidation & Reduction

Oxidation:

- Addition of O_2
- Removal of H
- Loss of electron
- Increase in oxidation state (O.S.)

Reduction:

- Addition of H_2
- Removal of O
- Gain of electron
- Decrease in oxidation state (O.S.)

⇒ Spontaneous & Non-Spontaneous Reaction:

→ Spontaneous

Definition:

- A reaction that happens on its own is called spontaneous.
- Exothermic (releases energy).

→ Non-Spontaneous:

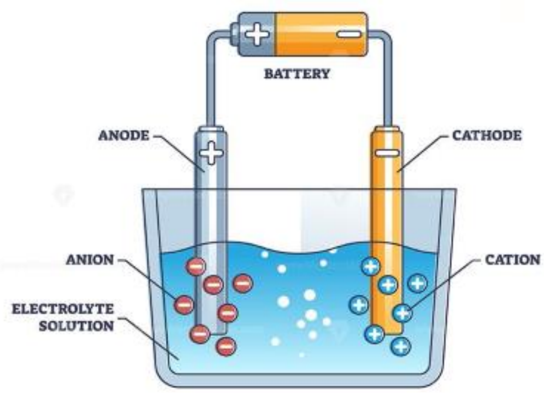
Definition:

- A reaction that does not occur on its own and need energy is called non-spontaneous.
- Endothermic (absorb energy).

⇒ Electrolytic Cell:

Components:

1. Container
2. Battery
3. 2 metallic rods
4. Electrolyte



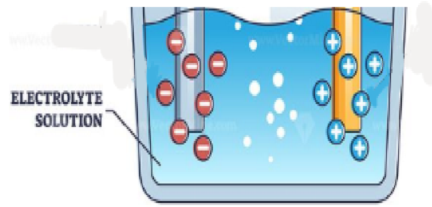
⇒ **Electrolyte:**

Definition:

Substance that conduct electricity in aqueous or molten form is called electrolyte.

OR

Substance which forms ions in water/molten ion.



Example:

$NaCl \rightarrow Na + Cl$

$KCL \rightarrow K + Cl$

▶ Electrolysis of molten electrolyte

⇒ Electrolysis of molten sodium chloride

Working:

- During electrolysis the sodium ions move towards cathode and chlorine ions move towards anode
- Sodium ions are reduced (gain of e-) at cathode to liquid sodium metal
- Cathode:

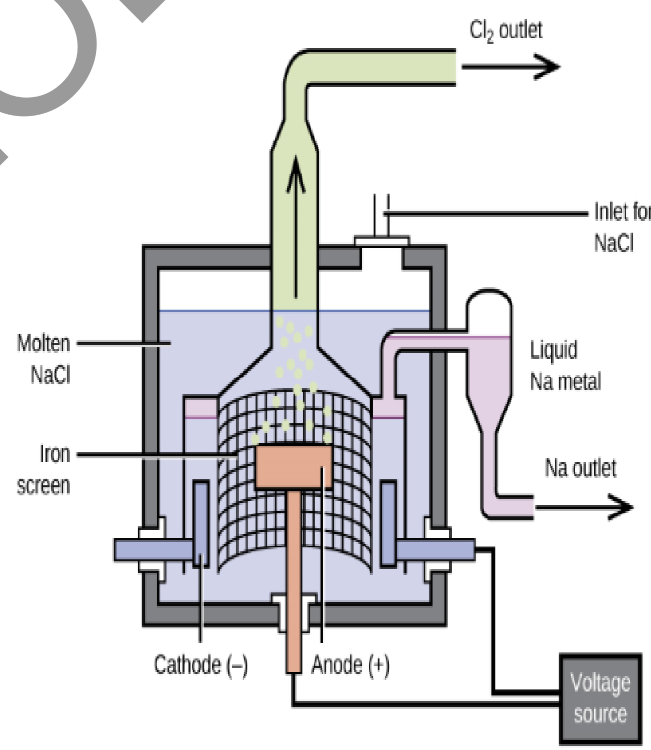
$$2Na^{+} + 2e^{-} \rightarrow 2Na$$

(reduction half reaction)

- Chlorine ions are oxidized (loss of e-) at anode to chlorine gas
- Anode:

$$2Cl^{-} \rightarrow Cl_2 + 2e^{-}$$

(oxidation half reaction)



- **Key points:**
- Sodium metal is produced on large scale by electrolysis of fused NaCl (molten NaCl) at 801°C
- The electrolytic cell called downs cell is used for production of Na metal.
- Cathode (–) is made up of iron (attached to -ve terminal)
- Anode (+) is made up of graphite (attached to +ve terminal)
- Na^{+} ions move towards cathode
- Cl^{-} ions move towards anode
- $NaCl \longrightarrow Na^{+} + Cl^{-}$

⇒ Electrolysis of molten lead II chloride (pbcl$_2$) using platinum(pt)/graphite

- Electrolysis of molten pbcl$_2$ using platinum will produce molten lead at cathode
- platinum electrodes are inert (doesn't react) acting as pathway for electron flow.

$$PbCl_2 \rightleftharpoons Pb^{2+} + 2Cl^-$$

- **Electrolyte:**

when lead chloride is heated to a molten state it dissociates ions

$$PbCl_2 \xrightleftharpoons{327^\circ C} Pb^{2+} + 2Cl^-$$

These ions are free to move and conducts electricity.

- **Cathode(-ve electrode):**

Pb+2 ions are attracted to cathode they will gain 2e-/reduced to form molten lead

$$Pb^{2+} + 2e^- \rightarrow Pb\ (l)$$

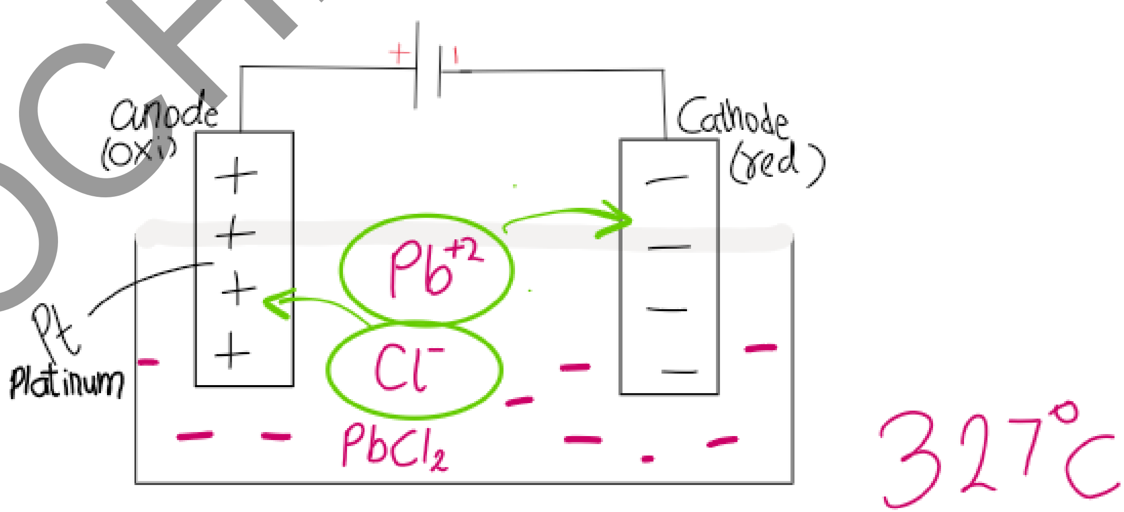
- **Anode(+ve electrode):**

Cl$^-$ ions are attracted to anode they will lose 1 e$^-$ each and are oxidized to form Cl$_2$ gas

- **Product:** lead (Pb),

chlorine gas (Cl$_2$)

- **Diagram:**



▶ Electrolysis of concentrated aqueous

⇒ Electrolysis of concentrated aqueous solution of NaCl

Explanation:

- **Electrolyte:**

Brine is used as an electrolyte
(A concentrated solution of NaCl in water.)

$$H_2O + NaCl \longrightarrow Brine$$

Ions:

$$H_2O + NaCl \rightleftharpoons OH^- + Cl^- + H^+ + Na^+$$

- **Cathode:**

H+ will move towards the cathode and will gain e^- because the reduction power of H+ is greater than Na+

$$2H + 2e^- \longrightarrow H_2$$

- **Anode:**

Cl^- will move towards cathode and lose e^- because the oxidation power of Cl^- is greater than OH^-

$$Cl^- \longrightarrow Cl_2 + 2e^-$$

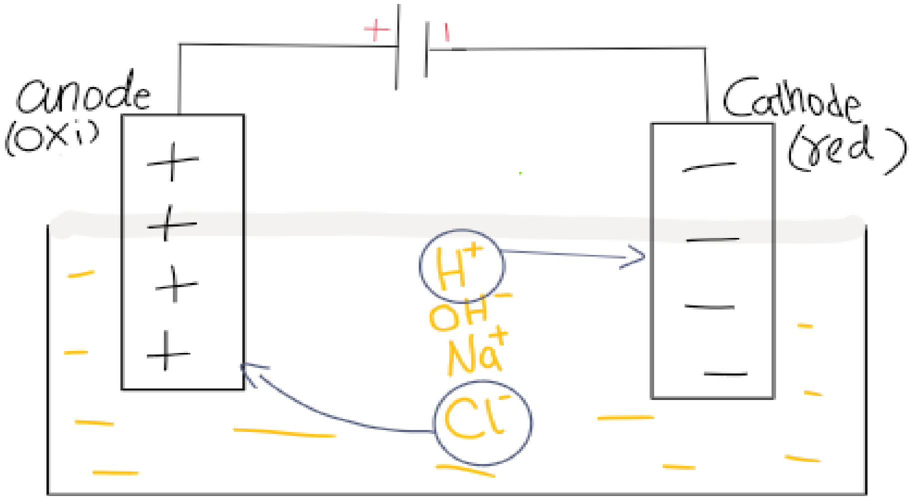
- **Products:**

hydrogen gas (H_2)
chlorine gas (Cl_2)
sodium hydroxide (NaOH)

Note:

- Electrolysis of aqueous solution of NaCl in water produces NaOH, Cl_2 and H_2 while, Electrolysis of molten NaCl produces Na and Cl_2

- **Diagram:**



- Overall reaction:

Anodic $2Cl^- \longrightarrow Cl_2 + 2e^-$

Cathodic $2H^+ + 2e^- \longrightarrow H_2$

$$2Cl^- + 2H^+ \longrightarrow Cl_2 + H_2$$

Electrolysis of Dil aqueous solution

A) Electrolysis of Dil aqueous H_2SO_4 using platinum electrode.

→ **Explanation:**

- **Electrode:**

The electrodes used are of platinum (inert-doesn't take part in the reaction)

- **Electrolyte:**

Dilute H_2SO_4 solution is used as an electrolyte

(A dilute solution of sulfuric acid in water)

From sulfuric acid:

$$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$$

from water :

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

Ions:

H^+, SO_4^{2-}, OH^-

- **Cathode:**

H^+ will move towards cathode and lose e^-

$$4H^+ + 4e^- \rightarrow 2H_2$$

- **Anode:**
- Competing ions: SO_4^{2-} and OH^-

OH^- will move towards the Anode and will gain e^- because the OH^- is unstable = more reactive while SO_4^{2-} is stable = less reactive

$$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$$

- **Products:**

hydrogen gas (H_2)
oxygen gas (O_2)

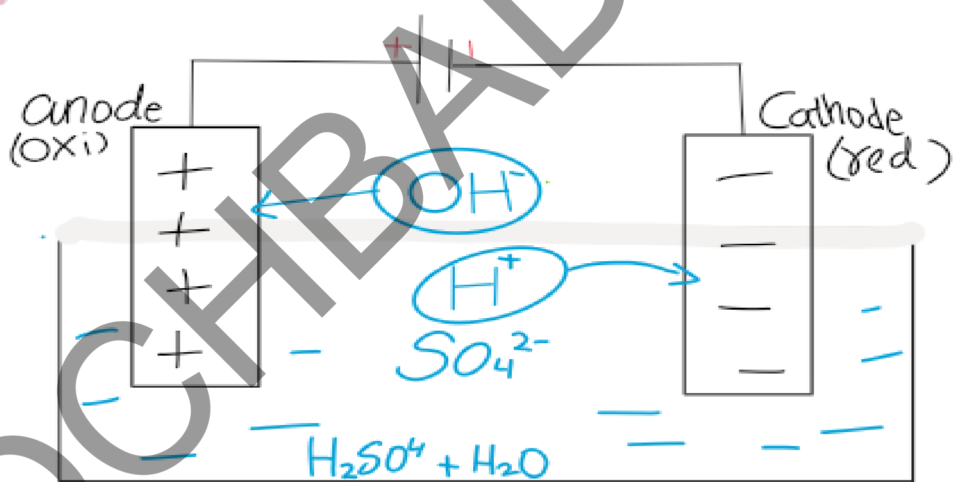
- **Overall equation:**

$$4H^+ + \cancel{4e^-} \rightarrow 2H_2$$

$$4OH^- \rightarrow 2H_2O + \cancel{4e^-} + O_2$$

$$\overline{4H^+ + 4OH^- \rightarrow 2H_2 + 2H_2O + O_2}$$

- **Diagram:**



B) Electrolysis of Cooper(II) sulphate using platinum electrode.

→ **Explanation:**

- **Electrode:**

The electrodes used are of platinum (inert-doesn't take part in the reaction)

- **Electrolyte:**

Dilute $CuSO_4$ solution is used as an electrolyte

(A dilute solution of copper sulphate in water)

From copper sulphate:

$$CuSO_4 \longleftrightarrow Cu^{2+} + SO_4^{-2}$$

from water :

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

Ions:

$$Cu^{2+} + SO_4^{2-} + H^+ + OH^-$$

- **Cathode:**

Copper(II) ions (Cu^{2+}) are attracted to the cathode and gain electrons, forming solid copper metal.

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

- **Anode:**

- Competing ions: SO_4^{2-} and OH^-

OH^- will move towards the Anode and will gain e^- because the OH^- is unstable = more reactive while SO_4^{2-} is stable = less reactive

$$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$$

- **Products:**

solid copper metal

Oxygen gas (O_2)

- **Color Change:**

- Copper

(blue) ⟶ transparent

Because their concentration in the solution decreases

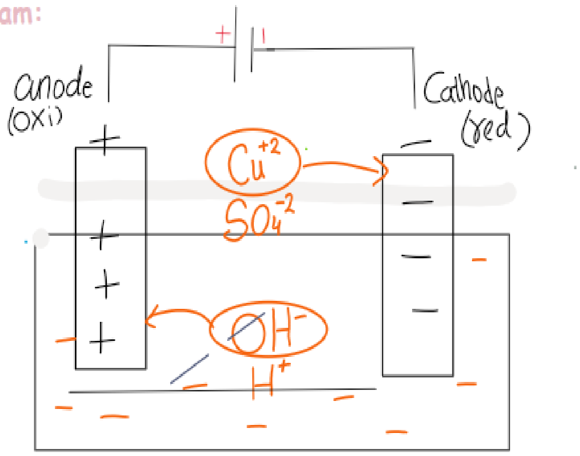
- Overall equation:

$$4OH^-_{(aq)} \rightarrow 2H_2O + 4e^- + O_{2(g)}$$

$$2Cu^{2+}_{(aq)} + 4e^- \rightarrow 2Cu_{(s)}$$

$$4OH^- + 2Cu^{2+} \rightarrow 2Cu + O_2 + 2H_2O$$

- Diagram:



C) Electrolysis of Cooper(II) sulphate using copper electrode.
Or
Purification of copper metal by electrolysis

If you use copper electrodes rather than platinum, the cathode reaction remains the same but the anode reaction differs

- Explanation:
- Electrode:

The electrodes used are of copper (active — they participate in the reaction)

Used for purification of copper metal

- Electrolyte:

$CuSO_4$ solution is used as an electrolyte

From copper sulphate:

$CuSO_4 \longleftrightarrow Cu^{2+} + SO_4^{-2}$

Note:

Impure copper → pure copper
Blistered copper
99 % — 99.5%
(zinc, iron, silver, gold)

Ions:

$Cu^{2+} + SO_4^{2-}$

- Process:
  - In this process impure copper bar act as anode and pure copper bar act as cathode
  - On passing electricity copper anode dissolves forming Cu^{+2} ions
  - Cu^{+2} ions moves towards cathode and the impurities (zinc, Fe, Au, Ag) will settle down as anode mud
  - Cu^{2+} will gain $2e^-$ and deposit at the cathode. Hence, impure copper is purified.
  - Purity increases from 99% → 99.5% pure copper.

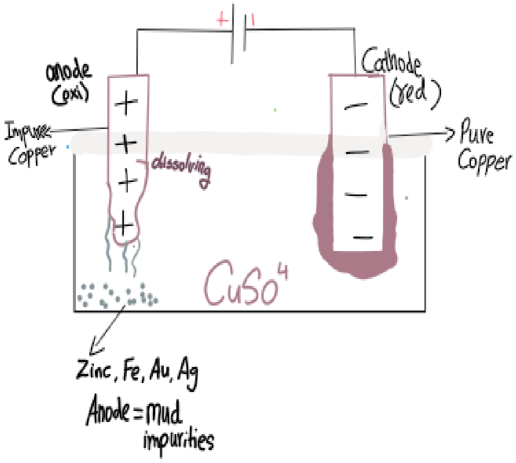
At anode: $Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^-$

At cathode: $Cu^{2+}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$

- **Products:**

  Purified copper metal

- **Diagram:**



Electrolysis of Halide compound

During the electrolysis of a halide compound (a compound containing halogen ions like chlorine, bromide, or iodide), the products depend on whether the solution is dilute or concentrated.

VIIA (halide ions end with -ide)

- $F \rightarrow F^-$
- $Cl \rightarrow Cl^-$
- $Br \rightarrow Br^-$
- $I \rightarrow I^-$
- $As \rightarrow$ (rare)

A) Electrolysis of dil halide solution

Anode:

Water is oxidized → Oxygen gas (O_2) is released.

Reason: Water competes with halide ions and wins due to low halide concentration.

Cathode:

Water is reduced → Hydrogen gas (H_2) is released.

Reason: Sodium ions are not discharged; water is easier to reduce.

Example: Electrolysis of dilute NaCl solution

- Anode: Oxygen gas (O_2)
- Cathode: Hydrogen gas (H_2)

Key idea: Water molecules are more likely to be discharged than halide ions because the concentration of halide ions is low.

A) Electrolysis of conc halide solution

- **Anode:** Halide ions (Cl^-, Br^-, I^-) are oxidized → Halogen gas (Cl_2, Br_2, I_2) is released.

  Reason: High concentration makes halide ions more likely to be discharged.

Key idea: Halide ions are now in high concentration, so they are discharged instead of wager.

- **Cathode:** Water is still reduced → Hydrogen gas (H_2) is released.
  Reason: Sodium ions are still not discharged.

- **Example:** Electrolysis of concentrated NaCl solution
- Anode: Chlorine gas (Cl_2)
- Cathode: Hydrogen gas (H_2)

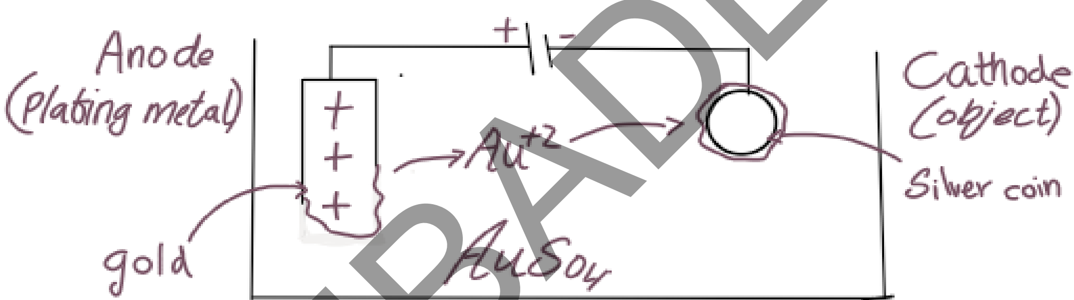
Key Rule to Remember:
Dilute Solution: Oxygen at the anode.
Concentrated Solution: Halogen at the anode.
(The cathode always produces hydrogen gas in both cases.)

Electroplating (deposition using the same electrode) (Important)

Definition: The electrolytic process used to deposit one metal on another metal is called Electroplating.

At ANODE → Plating metal

At CATHODE → Object



⇒ Electroplating Applications

- **Jewelry:** Adds shine and prevents corrosion (e.g. gold/silver plating).
- **Trophies & Medals:** Decorative coating for prestige.
- **Steel Objects:** Zinc plating protects from rust.
- **Car Parts:** Chrome plating for durability and appearance.
- **Electrical Components:** Gold/silver plating improves conductivity.
- **Kitchenware:** Silver plating enhances look and hygiene.

⇒ Zinc plating

→ Zinc plating is also called Galvanizing

- **Electrodes:**

→ Zn metal is taken at anode

→ Steel object is taken at cathode

- **Electrolyte:**

→ Electrolyte used for this process: potassium zinc cyanide complex $K_2[Zn(CN)_4]$

- **During Electrolysis:**

→ Zn from the anode enters the solution as Zn^{2+} ions

→ Zn^{2+} will move toward cathode (−) & deposited

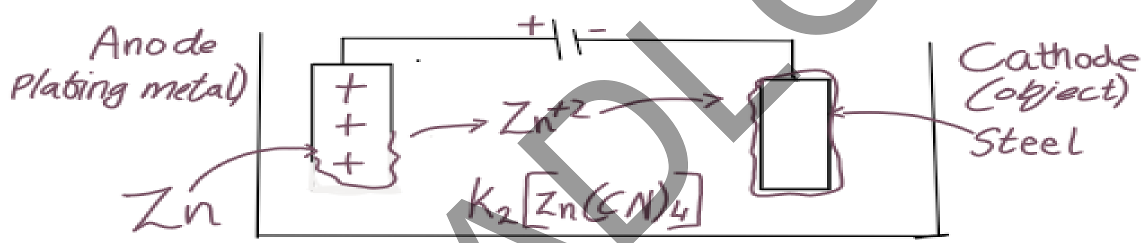
→ To prevent hydrolysis, NaCN is added in Electrolyte

- **reactions at the electrodes:**

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

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

- **Diagram:**



⇒ Tin plating

→ Tin (Sn) is also called stannous

- **Electrodes:**

→ Tin metal is taken at anode

→ Steel object is taken at cathode

- **Electrolyte:**

→ Electrolyte used for this process: stannous sulphate $SnSO_4$

- **During Electrolysis:**

→ When electricity starts flowing from the battery, Sn from the anode enters the solution as Sn^{2+} ions

→ Sn^{2+} will move toward cathode (−) & deposited

→ To prevent hydrolysis, H_2SO_4 is added in Electrolyte

- Reactions at electrodes:

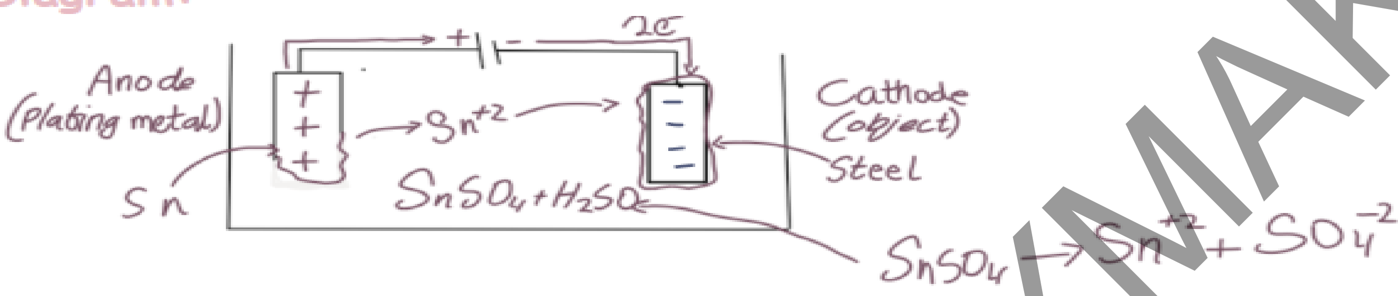
At anode:

$$Sn_{(s)} \rightarrow Sn^{2+}_{(aq)} + 2e^-$$

At cathode:

$$Sn^{2+}_{(aq)} + 2e^- \rightarrow Sn_{(s)}$$

- Diagram:



⇒ **Chromium plating**

→ Cr metal does not adhere strongly to steel, therefore steel is first electroplated with Cu or Ni metal.

Electrode:

→ For electroplating chromium,
The Cr metal is taken as the anode and
The steel object to be electroplated is taken as cathode.

Electrolyte:

→ $Cr(SO_4)_3$ (stannous sulphate) is used as the electrolyte.
To prevent its hydrolysis, a few drops of dilute H_2SO_4 are added.

Electrolysis:

→ When electricity starts flowing from the battery, Cr from the anode enters the solution as Cr^{3+} ions

→ Cr^{3+} will move toward cathode (−) & deposited

→ To prevent hydrolysis, H_2SO_4 is added in Electrolyte

Reactions at electrodes:

- At anode:

  - $$Cr_{(s)} \rightleftharpoons Cr^{3+}_{(aq)} + 3e^-$$

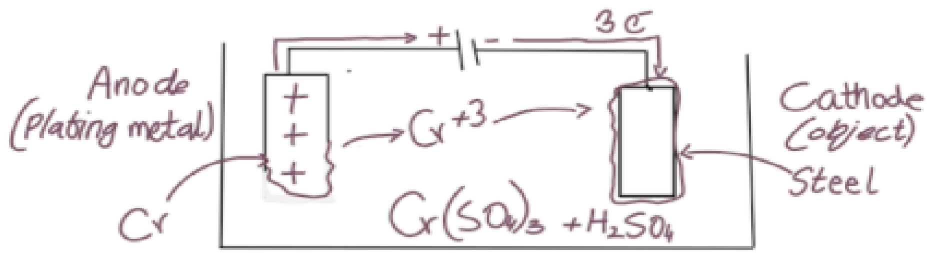
- At cathode:

$$Cr^{3+}_{(aq)} + 3e^- \rightleftharpoons Cr_{(s)}$$

- Uses:

Cr plated steel are used to make automobile parts

- Diagram:



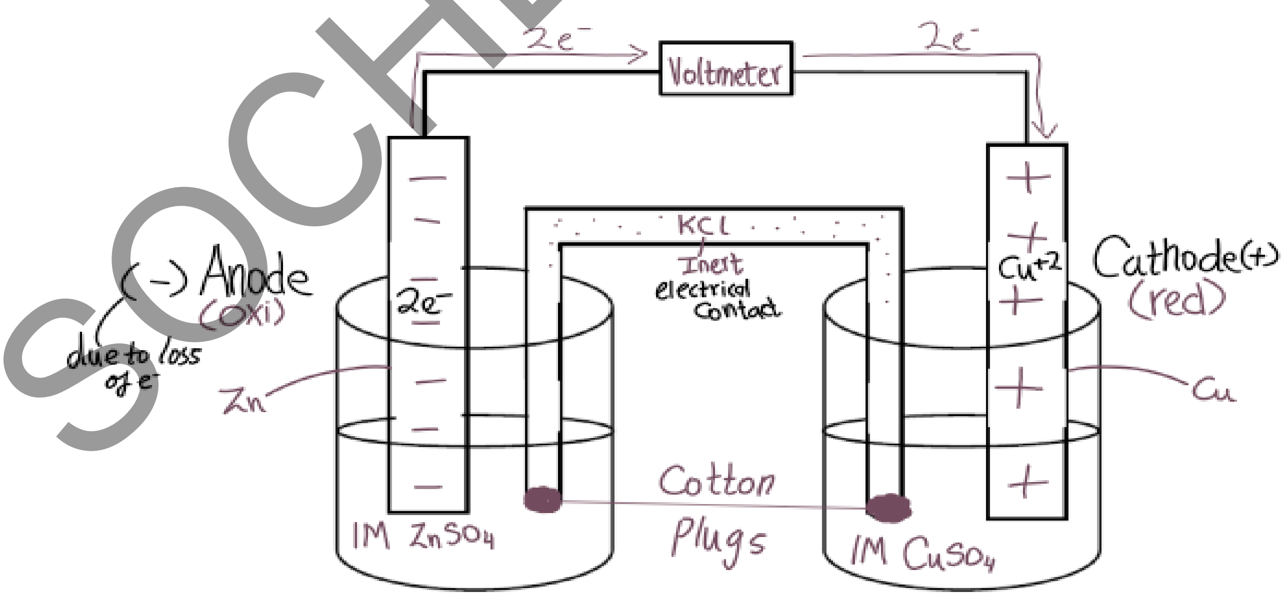
Galvanic cell

- Introduction:
  - Discovered by Frederick Daniel
  - Constructed the first voltaic cell using:
    - Zinc (Zn) as the anode
    - Copper (Cu) as the cathode
  - Named after Alessandro Volta → hence "voltaic cell"

- Other Names:
  - Daniel Cell
  - Voltaic Cell

- Spontaneous Reaction:
  - Involves a spontaneous redox reaction
  - Converts chemical energy → electrical energy

- Components:
  1. Anode: Zinc (Zn) in 1M $ZnSO_4$
  2. Cathode: Copper (Cu) in 1M $CuSO_4$
  3. Salt Bridge: KCl (maintains ionic contact between solutions)
  4. Voltmeter: Measures current in the solution



1) Anodic Half Cell

- Electrode used: Zinc (Zn)
- Electrolyte: 1M $ZnSO_4$ solution
- Reaction: $Zn \rightarrow Zn^{2+} + 2e^-$ (Zinc loses 2 electrons — oxidation)

2) Cathodic Half Cell

- Electrode used: Copper (Cu)
- Electrolyte: 1M $CuSO_4$ solution
- Reaction: $Cu^{2+} + 2e^- \rightarrow Cu$ (Copper gains 2 electrons — reduction)

Electron Flow & Measurement

- $2e^-$ are released by Zn at the anode which travel through the external circuit through voltmeter that measure the electrical current
- These $2e^-$ are accepted by Cu at the cathode.

Salt Bridge

- Maintains electrical neutrality and ion flow between the half cells.
- Contains an inert electrolyte (e.g., KCl).

Electrochemical series(important)

Definition:

A table showing arrangements of metals based on increasing reduction potentials is call a electrochemical series

Reduction potential =how easily a substance gains electrons (i.e., gets reduced).

Element	Electrode rection	E°(V)
Li	$Li + e^- \rightarrow Li$	−3.045
K	$K^+ + e^- \rightarrow K$	−2.925
Cs	$Cs^+ + e^- \rightarrow Cs$	−2.923
Ba	$Ba^{2+} + 2e^- \rightarrow Ba$	−2.906
Ca	$Ca^{2+} + 2e^- \rightarrow Ca$	−2.866
Na	$Na^+ + e^- \rightarrow Na$	−2.714
Mg	$Mg^{2+} + 2e^- \rightarrow Mg$	−2.363
Al	$Al^{3+} + 3e^- \rightarrow Al$	−1.662
H_2	$H_2O + 2e^- \rightarrow H_2 + 2OH^-$	−0.829
Zn	$Zn^{2+} + 2e^- \rightarrow Zn$	−0.763
Fe	$Fe^{2+} + 2e^- \rightarrow Fe$	−0.440
Cd	$Cd^{2+} + 2e^- \rightarrow Cd$	−0.403
Pb	$PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$	−0.310
Co	$Co^{2+} + 2e^- \rightarrow Co$	−0.280
Ni	$Ni^{2+} + 2e^- \rightarrow Ni$	−0.250
Sn	$Sn^{2+} + 2e^- \rightarrow Sn$	−0.136
Pb	$Pb^{2+} + 2e^- \rightarrow Pb$	−0.126
Fe	$Fe^{3+} + 3e^- \rightarrow Fe$	−0.036
H_2	$2H^+ + 2e^- \rightarrow H_2$(SHE)	0
Cu	$Cu^{2+} + e^- \rightarrow Cu^+$	+0.153
S	$S_4O_6^{2-} + 2e^- \rightarrow 2S_2O_3^{2-}$	+0.170
Cu	$Cu^{2+} + 2e^- \rightarrow Cu$	+0.337
I_2	$I_2 + 2e^- \rightarrow 2I^-$	+0.534
Fe	$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.77
Ag	$Ag^+ + e^- \rightarrow Ag$	+0.799
Hg	$Hg^{2+} + 2e^- \rightarrow Hg$	+0.854
Br_2	$Br_2 + 2e^- \rightarrow 2Br^-$	+1.066
O_2	$O_2 + 4H^+ + 2e^- \rightarrow 2H_2O$	+1.230
Cr	$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	+1.330
Cl_2	$Cl_2 + 2e^- \rightarrow 2Cl^-$	+1.359
Au	$Au^{3+} + 3e^- \rightarrow Au$	+1.498
Mn	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	+1.510
F_2	$F_2 + 2e^- \rightarrow 2F^-$	+2.870

Increase

Low Value
- ↳ Instable
- ↳ (−ve)
- ↳ easily lose e⁻ (oxidation)
- ↳ active metal
- ↳ act as anode

→ SHE 0.00
standard hydrogen electrode

Standard condition
- → 1 mole
- → 1 atm
- → 25°C

above H_2 low value
below H_2 high value

High value
- ↳ stable
- ↳ +ve
- ↳ less active
- ↳ easily gain e⁻ (reduction)
- ↳ act as cathode

Electrode potential:

Measure of the potential of a reaction that can occur at the electrode.

- It tells us how likely an electrode is to gain or lose electrons
- Measured in volts

Electrochemical series helps us to:

- Predict which substance will form the anode and cathode.
- Determine the reactivity of metals.
- Identify:
  - Oxidation half-cell (where electrons are lost)
  - Reduction half-cell (where electrons are gained)

Example:

$Na^+ \longrightarrow 2Na.$

At the Anode
Small | Negative Electrode
(−2.714 V)
Active metal
Oxidation
Easy to lose e⁻

$H^+ \longrightarrow H_2$ gas

At the cathode
Greater | positive electrode
(0.00 V)
Less active metal
reduction
Easy to gain e⁻

OH^-

→ Anode
→ Active metal
→ Oxidation
→ Lose of e^-
→ Small, -ve value
→ Product: O_2 gas and H_2O

Cl^-

→ Cathode
→ Less active
→ Reduction
→ Gain of e^-
→ More, +ve value
→ Product: Cl_2 gas

Battery, A source of electrical energy

Definition:

A Battery is galvanic cell or group of galvanized cell joined in series.

Galvanic → Galvanic → Galvanic

- It generate electric current by redox reaction.
- Electrons will move anode to cathode

- **Example:**
- Dry cell
- Storage cell
- Mercury battery

- Fuel cell

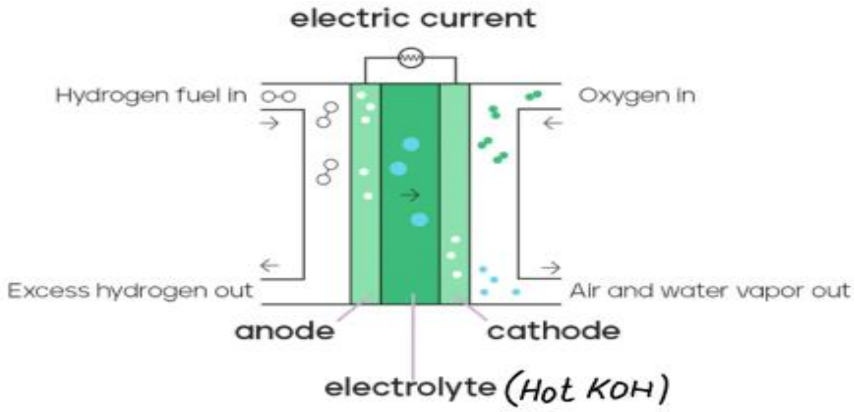
$2H_2 + O_2 \rightarrow 2H_2O$ + heat → electricity (Endothermic)

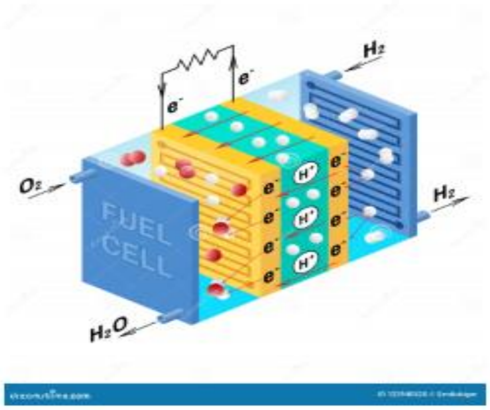
→ When hydrogen burns in presence of oxygen it produces water and heat.

→ Energy released from reaction is converted into electrical energy.

Compartment of fuel cell (inert):

1. Electrode of porous carbon anode
2. Electrode of porous carbon cathode
3. Middle compartment: hot aq. KOH solution





- Fuel cell has 3 compartments
- In anode compartment, H_2 gas reacts with OH^- ions and forms water & $4e^-$
- In cathode compartment, O_2 gas reacts with water and forms OH^- ions
- Middle compartment contains hot aq. KOH electrolyte
- Fuel cell generates 0.9V energy
- Fuel cell operates at high temperature
- Fuel cell forms evaporated water & if needed, it can be condensed!

⇒ Advantages:

→ Clean source of energy
→ Highly reliable, don't have moving parts
→ Eco-friendly, doesn't release pollution
→ Low-cost fuel
→ Long lifespan with little maintenance
→ H_2 and O_2 used in fuel are obtained from renewable energy sources

⇒ Disadvantages:

→ H_2 production issues
→ H_2 is highly inflammable, explosive in nature
→ Transport and storage of H_2 is difficult
→ Water produced by fuel cell is very expansive

Multiple choice question

EXERCISE

1. Multiple Choice Questions (MCQs)

i. Which of the following statements is not correct about the galvanic cell?
   (a) Cations are reduced at the cathode (b) Anions are oxidized at the anode
   (d) Electrons flow from cathode to anode (c) Oxidation occurs at the anode

ii. Which of the following is not true about the Daniel cell
   (a) Half-cell of an active metal acts as a cathode.
   (b) Half-cell contains an element in contact with its ions in aqueous solution.
   (c) A salt bridge connects the two half-cells.
   (d) A spontaneous oxidation-reduction reaction generates electricity.

iii. Which of the following does not involve an electrolytic process?
   (a) Refining of copper
   (b) Manufacture of sodium from NaCl
   (c) Electroplating of steel
   (d) Reduction of metal oxide by a reducing agent.

iv. Galvanizing is
   (a) Coating with Sn (b) Coating with Zn
   (c) Coating with Cr (d) Coating with Cu

v. Which of the following is true for the Nelson cell?
   (a) Sodium metal is produced at the anode
   (b) Chlorine gas is produced at anode
   (c) Hydrogen gas is produced at the anode
   (d) Sodium ions are not reduced at the cathode

vi. Electroplating involves:
   a) Using an electrolyte to oxidize a metal.
   b) Coating a metal object with a thin layer of metal through electrolysis.
   c) Melting a metal object to form a thin coating.
   d) Using heat to form a layer of metal on another metal object

vii. In a voltaic cell, the salt bridge serves the function of:
   a) Preventing electrons from flowing in the external circuit.
   b) Maintaining electrical neutrality by allowing ion movement.
   c) Storing excess energy from the cell.
   d) Generating additional voltage for the cell.

viii. If the voltage data shows metal A has a higher standard reduction potential metal B, it can be concluded that:
   a) Metal A is more reactive than metal B.
   b) Metal B is more reactive than metal A.
   c) Metal A and metal B have the same reactivity.
   d) Metal A is less likely to gain electrons than metal B

Exercise questions

Q i. **What is a fuel cell?**

Ans: *Fuel Cell:*

A fuel cell is a special type of galvanic cell which is based upon the reaction between oxygen and gaseous fuel hydrogen.

A fuel cell is a special type of galvanic cell which is based upon the reaction between oxygen and a gaseous fuel hydrogen. When hydrogen burns in air, an exothermic reaction occurs and a lot of chemical energy is released in the form of heat and light. In this reaction, hydrogen is oxidized to water.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g) + \text{heat}$$

Q ii. **Write chemical reactions that occur in Nelson's cell.**

Ans: Following reactions occur in the electrolytic cell Nelson's cell:

Brine ionizes to produce ions:

$$2NaCl(aq) \rightarrow 2Na^{+}(aq) + 2Cl^{-}(aq)$$

At anode (Oxidation):

$$2Cl^{-}(aq) \rightarrow Cl_2(g) + 2e^{-}$$

At cathode (Reduction):

$$4OH^{-}(aq) + 2H_2O(l) + 2e^{-} \rightarrow 2NaOH(aq)$$

Overall Reaction:

$$2NaCl(aq) + 2H_2O(l) \rightarrow H_2(g) + Cl_2(g) + 2NaOH(aq)$$

Q iii. **Why tin-plated steel is used to make food cans?**

Ans: Food cans are made from steel with a thin coating of tin to prevent corrosion.

Tin plated steel is used to make cans. Food and beverages industries use tin-plated steel cans; this is because the components of food beverages and the preservatives contain organic acids or their salts. They may form toxic substances by reacting with iron and acids and salts are corrosive. Tin plating is non-poisonous and prevents corrosion.

Q iv. Explain one example from daily life which involves an oxidation-reduction reaction.

Ans: *Redox in photography:*

A photographic film is basically an emulsion of silver bromide $(AgBr)$ in gelatin. When the film is exposed to light, silver bromide granules become activated. This activation depends upon the intensity of the light and the time of exposure.

When exposed film is placed in the developer solution that is actually a reducing agent. Hydroquinone which is used in this reducing agent is used as developer. In hydroquinone the activated granules of silver bromide are reduced to black metallic silver.

$$Ag^+ + 1e^- \rightarrow Ag(s)$$

Inactivated silver bromide is removed from the film by using a solvent called a fixer. Sodium thiosulphate is used for this purpose. The areas of the film exposed to high light appear darkest because they have the highest concentration of metallic silver. Thus photography involves oxidation-reduction reaction.

Q v. Define electrochemical series.

Ans: A table showing the arrangement of metals based on increasing reduction potentials is called an electrochemical series.

Q vi. Why does chlorine gas form at the anode during the electrolysis of molten lead (II) chloride?

Ans: Chlorine gas forms at the anode during the electrolysis of molten lead(II) chloride $(PbCl_2)$ because chloride ions (Cl^-) are oxidized. At the anode, chloride ions lose electrons to form chlorine gas (Cl_2) according to the reaction:

$$2Cl^- \rightarrow Cl_2 + 2e^-$$

This occurs because halide ions are more easily oxidized than lead ions, which are reduced at the cathode to form solid lead (Pb).

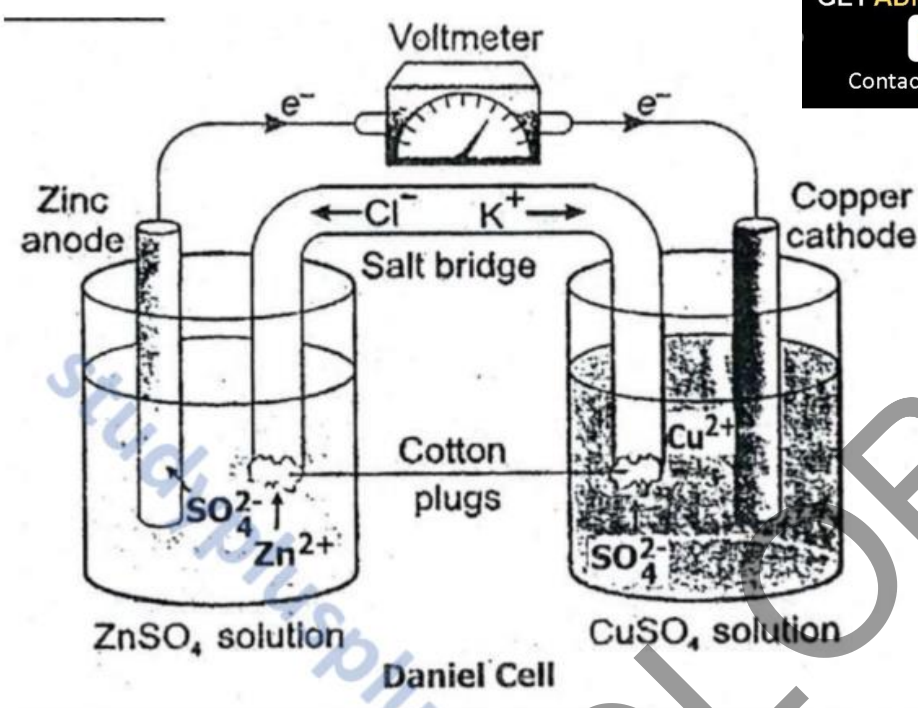
Q vii. How do hydrogen-oxygen fuel cells benefit the environment compared to gasoline engines?

Ans: Hydrogen-oxygen fuel cells benefit the environment compared to gasoline engines primarily because they produce zero emissions during operation. The only byproduct of a hydrogen fuel cell is water vapor (H_2O), whereas gasoline engines release harmful pollutants such as carbon dioxide (CO_2), nitrogen oxides (NO_x), and particulate matter, which contribute to air pollution and climate change.

Additionally, fuel cells are generally more efficient than gasoline engines, meaning less energy is wasted during operation. As hydrogen can be produced from renewable sources, fuel cells also offer the potential for a cleaner energy cycle, further reducing the environmental impact compared to fossil fuel-based engines.

Q viii. Sketch a Daniel cell with labelled components and indicate the direction of electron flow.

Ans: *Daniel Cell:*



Daniel Cell

Electrochemistry

SLO QUESTIONS

Q1: What is the Difference between electrolyte & non electrolyte ?

Property	Electrolyte	Non-Electrolyte
Definition	A substance that conducts	A substance that does not conduct
Electricity in Solution	electricity in solution	electricity in solution
Ion Formation	Dissociates into ions in water	Does not dissociate into ions
Examples	NaCl (salt), HCl (acid), KOH (base)	Sugar, glucose, ethanol
Conductivity in Water	High	None or very low
Type of Bonding	Usually ionic compounds or strong acids/bases	Usually covalent compounds
Effect on Bulb Test	Lights up the bulb (if strong)	Does not light the bulb

Q2: What is the Difference between electrolytic & galvanic cell ?

Ans)

Feature	Electrolytic Cell	Galvanic Cell (Voltaic Cell)
Energy Conversion	Converts electrical energy to chemical energy	Converts chemical energy to electrical energy
External Power Source	Required (battery or power supply is used)	Not required (it generates electricity itself)
Spontaneity of Reaction	Non-spontaneous (needs electricity to occur)	Spontaneous (occurs naturally)
Electrode Charges	Anode is positive, Cathode is negative	Anode is negative, Cathode is positive
Examples	Electrolysis of water, electroplating	Daniell cell, batteries

Q3: What happen if we remove battery from electrolytic cell?

Ans) If we remove the battery from an electrolytic cell, here's what happens:

- No Reaction Will Occur

Reason: The electrolytic cell works by using external electrical energy (from the battery) to drive a non-spontaneous reaction.
Without the battery, there is no energy source to force the ions to move and cause a redox (oxidation-reduction) reaction.

- No Electron Flow

Electrons will not flow from anode to cathode.
Electrodes become inactive, so no electrolysis takes place.

- No Product Formation

For example, in electrolysis of water, no hydrogen or oxygen gas will be produced without the battery.
I)Why Anode +ve & Cathode -ve charge in electrolytic cell?

Q4. Why Anode is +ve and Cathode is −ve in an Electrolytic Cell?

Ans)

◆ Anode = Positive (+)

- The battery pulls electrons away from the anode.
- This causes oxidation (loss of electrons).
- As electrons are lost, the anode becomes electron-deficient → positively charged.
- It attracts negative ions (anions) from the solution.

◆ Cathode = Negative (−)

- The battery pushes electrons into the cathode.
- This extra supply of electrons makes it negatively charged.
- Positive ions (cations) from the solution gain these electrons → reduction occurs.

Q5. Difference Between Spontaneous and Non-Spontaneous Reactions

Feature	Spontaneous Reaction	Non-Spontaneous Reaction
Definition	Happens naturally under normal conditions	Needs **external energy** to occur
Energy Requirement	No external energy needed	Requires external energy (e.g., heat, electricity, light)
Examples	Rusting of iron, combustion, mixing salt in water	Electrolysis of water, charging a battery
Common In	Galvanic cells, natural processes	Electrolytic cells, artificial processes

Q6: What is direction of flow of electrons in electrolytic cells?

Ans) Direction of Electron Flow in Electrolytic Cell:

- Electrons flow from the anode (+) to the battery, then from the battery to the cathode (−).
- So, external flow:

  Anode → Battery → Cathode.
- Internal flow:

  Cations → Cathode, Anions → Anode.

Q7: What safety precautions are needed during electrolysis of molten NaCl?

Ans)

- Use heat-resistant gloves and goggles
- Keep away from water – molten salt reacts strongly
- Work in a well-ventilated area (chlorine gas is harmful)
- Avoid touching the hot equipment – it's at very high temperature

. Q8: Why does solid NaCl not conduct electricity, but molten NaCl does?

Ans)

Solid NaCl does not conduct electricity, but molten NaCl does because:

- In solid NaCl, the ions are fixed and can't move
- In molten NaCl, the ions are free to move and carry current.

Q9: What are the electrode materials used in Down's cell?

Ans)

Materials used in Downs cell are:

- Anode: Graphite (carbon)
- Cathode: Iron

Q10: Explain the half-reactions occurring at anode and cathode during electrolysis of molten NaCl.

Ans)

Half-reactions occurring during electrolysis of molten NaCl:

- At anode (+):

$$2Cl^- \rightarrow Cl_2\ (gas) + 2e^-\ (Oxidation)$$

- At cathode (−):

$$Na^+ + e^- \rightarrow Na\ (liquid\ metal)\ (Reduction)$$

Q11: Why is sodium chloride melted before electrolysis?

Ans)

Sodium chloride is melted before electrolysis because only molten NaCl has free ions to conduct electricity. In solid form, the ions are locked and can't move

Q12: What products are formed during electrolysis of molten NaCl?

Ans)

Products formed during the electrolysis of molten NaCl are:

- Sodium metal (Na) at the cathode
- Chlorine gas (Cl_2) at the anode

Q13: Why does chlorine gas form at the anode?

Ans)

Chlorine gas form at the anode because Cl^- ions lose electrons (oxidation) and form Cl_2 gas.

Q14: Why is sodium metal produced in liquid form during this process?

Ans)

Sodium metal is produced in liquid form during this process because:

- Sodium's melting point is low (98°C)
- The electrolysis temperature is very high (~800°C), so sodium is liquid

Q15: What is the temperature required to melt sodium chloride for electrolysis?

Ans)

Temperature required to melt sodium chloride for electrolysis is 801°C

Q16: Write the overall redox reaction for electrolysis of molten NaCl.

Ans)

$$2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$$

Q17: Differentiate between oxidation and reduction half-reactions in this process

Ans)

Feature	Oxidation Half-Reaction	Reduction Half-Reaction
Occurs at	Anode (+)	Cathode (–)
Ion involved	Chloride ion (Cl^-)	Sodium ion (Na^+)
Reaction	$2Cl^- \rightarrow Cl_2 + 2e^-$	$Na^+ + e^- \rightarrow Na$
Electron transfer	Electrons are **lost**	Electrons are **gained**
Product formed	Chlorine gas (Cl_2)	Liquid sodium metal (Na)

18) What are different conditions by which electrochemical processes/reactions stop?

Answer:

1. Reactants are used up – no ions or molecules are left to sustain the reaction.
2. Products build up – gases or deposits on electrodes block further reaction.
3. Voltage drops too low – potential difference falls below the decomposition potential.
4. Circuit is broken – disconnection, electrode damage, or electrolyte loss stops current flow.

19) Why does an electrolytic cell require an external power source, but a galvanic cell does not

Answer:

- In a galvanic cell, the redox reaction is spontaneous and naturally produces electrical energy.
- In an electrolytic cell, the redox reaction is non-spontaneous and must be driven by an external power source to force electrons to flow in the opposite direction to their natural tendency.

Q20: Write the products of

- Fused NaCl electrolysis
- Brine electrolysis
- Dilute aqueous solution of NaCl electrolysis.

Answer:

- Fuse NaCl electrolysis – Products:

Na (sodium metal) and Cl_2 (chlorine gas)

- Brine electrolysis – Products:

H_2 (hydrogen gas), Cl_2 (chlorine gas), and NaOH (sodium hydroxide solution)

- Dilute aqueous solution of NaCl electrolysis – Products:

H_2 (hydrogen gas) and O_2 (oxygen gas)

Q21) Why does NaCl not participate in electrolysis of dilute aqueous solution of NaCl?

Answer

Because in dilute solution, water is preferentially discharged:

- At cathode, H^+ from water is reduced more easily than Na^+ (Na^+ has much more negative reduction potential).
- At anode, OH^- from water is oxidized more easily than Cl^- in dilute conditions.

Q22) Difference between Down cell and Nelson cell:

Feature	Down's Cell	Nelson Cell
Electrolyte	Molten NaCl	Concentrated NaCl(aq)
Main Products	Na and Cl_2	NaOH, H_2, and Cl_2
Temperature	High (molten)	Low (aqueous)
Purpose	Metal extraction	Caustic soda production

Q23: Why H^+ ions are reduced at cathode but Na^+ is not (in dilute NaCl solution electrolysis)?

Because H^+ ions have a higher reduction potential (0 V) and are easier to reduce than Na^+ ions (−2.71 V), so H^+ is discharged first.

Q24)Why does the anode of galvanic cell has negative charge and why anode of electrolytic cell has positive charge?

- Galvanic cell: Anode is negative because oxidation releases electrons which flow out into the circuit.
- Electrolytic cell: Anode is positive because it's connected to the positive terminal of an external power source, pulling electrons away to force oxidation.

Q25) What would happen if both electrodes in a galvanic cell were made of the same metal?

No potential difference is created, so no current flows.

Q26) What happens if we remove the salt bridge?

Electron flow stops because charges build up in each half-cell, preventing the reaction.

Q27) Why does a galvanic cell need a voltmeter?

To measure the EMF between the two electrodes without drawing significant current.

Q28) Why there is a cotton plug at the end of the salt bridge?

To prevent leakage of the salt solution while still allowing ion movement.

Electrochemical Series Mnemonics

- Ascending Order of Reactivity (Least Reactive to Most Reactive Metals):
- Mnemonic:
  "A little king cleverly bakes cakes. Naughty monkeys always hide zesty fresh cakes. Pretty cool naughty squirrels play funny happy cats singing cool inspiring funny artworks happily baking orange cream cakes and muffins forever."

Word	Symbol/Ion
Little	Li
King	K
Cleverly	Cs
Bakes	Ba
Cakes	Ca
Naughty	Na
Monkeys	Mg
Always	Al
Hide	H_2O
Zesty	Zn
Fresh	Fe^{2+}
Cakes	Cd
Pretty	$PbSO_4$
Cool	Co
Naughty	Ni
Squirrels	Sn
Play	Pb^{2+}
Funny	Fe^{3+}/Fe^{2+}
Happy	H_2
Cats	Cu^{+}
Singing	$S_2O_3^{2-}$
Cool	Cu^{2+}
Inspiring	I_2
Funny	Fe^{3+}/Fe^{2+}
Artworks	Ag
Happily	Hg
Baking	Br_2
Orange	O_2
Cream	$Cr_2O_7^{2-}$
Cakes	Cl_2
And	Au
Muffins	MnO_4^{-}
Forever	F_2