

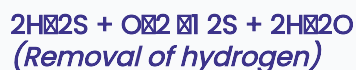
- Concept of Oxidation–Reduction, Redox Reactions and Oxidation Number
- Types of Redox Reactions, Balancing of Redox Reactions and Electrode Processes
- Quick Reference Table
- Common Mistakes and Misconceptions
- Glossary

Concept of Oxidation–Reduction, Redox Reactions and Oxidation Number

Oxidation

Oxidation is defined as the addition of oxygen or an electronegative element to a substance, or the removal of hydrogen from a substance.

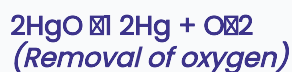
Examples:



Reduction

Reduction is defined as the removal of oxygen or an electronegative element from a substance, or the addition of hydrogen to a substance.

Examples:



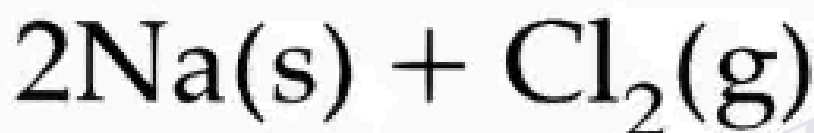


Redox Reaction

A redox reaction is a chemical reaction that involves both oxidation and reduction processes occurring simultaneously.

Redox Reaction in Terms of Electron Transfer

loss of $2e^-$



gain of

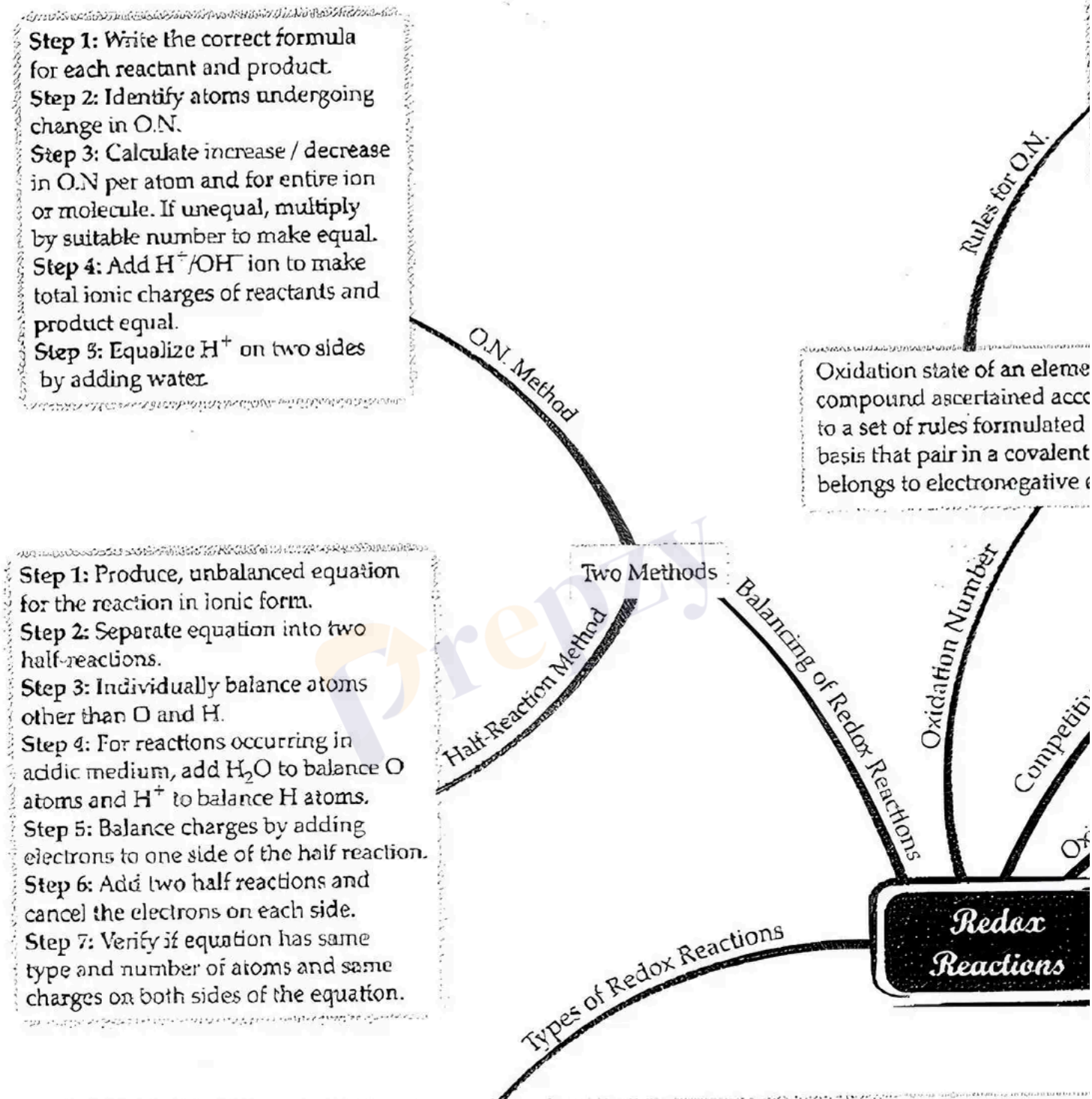
(Oxidation half reaction)



(Reduction half reaction)

In this reaction, sodium (Na) loses electrons to form Na^+ , which is oxidation, and chlorine (Cl_2) gains electrons to form Cl^- , which is reduction. This is a redox reaction.

According to the electronic concept, the half-reaction involving loss of electrons is called the oxidation half-reaction.



This diagram summarizes redox reactions, oxidation numbers, types of redox reactions, electron transfer, and more.

Oxidising Agent

The species that oxidises another substance and itself gets reduced is called the oxidising agent. It accepts electrons.

Reducing Agent

The species that reduces another substance and itself gets oxidised is called the reducing agent. It donates electrons.

Oxidation Number

Oxidation number is the charge assigned to an element in a compound based on a set of rules assuming electronegativity.

According to the oxidation number concept:

- Oxidation is an increase in oxidation number.
- Reduction is a decrease in oxidation number.
- Oxidising agent increases the oxidation number of an element.
- Reducing agent decreases the oxidation number of an element.
- Redox reactions involve changes in oxidation numbers.

Rules for Oxidation Numbers

- Elements in their free state have oxidation number zero (e.g., H_2 , O_2 , P_4 , Na, Mg).
- For monoatomic ions, oxidation number equals the ion charge (e.g., Na^+ , Mg^{2+} , Fe^{3+} , Cl^- , O^{2-}).
- Oxygen usually has oxidation number -2 except in peroxides (-1) and superoxides ($-1/2$), and when bonded to fluorine (positive values).
- Hydrogen usually has oxidation number $+1$ except when bonded to metals in binary compounds where it is -1 .
- Fluorine always has oxidation number -1 .
- Other halogens have oxidation number -1 when in halide ions but positive values when combined with oxygen.
- The sum of oxidation numbers in a neutral compound is zero.
- The sum of oxidation numbers in a polyatomic ion equals the ion charge.

Examples of Oxidation Numbers in Ions

Na^+ : $+1$

Mg^{2+} : $+2$

Fe^{3+} : $+3$

Cl^- : -1

O^{2-} : -2

Mnemonics

Concept: Redox reaction

Mnemonic: OIL RIG

Interpretation: Oxidation Is Losing electrons, Reduction Is Gaining electrons.

Solved Examples

Example 1: Identify oxidation and reduction in the reaction: $2\text{HgO} \rightarrow 2\text{Hg} + \text{O}_2$

Solution:

Step 1: Assign oxidation numbers.

Hg in HgO: +2, O in HgO: -2

Hg in Hg: 0, O in O_2 : 0

Step 2: Compare oxidation numbers.

Hg changes from +2 to 0 (reduction).

O changes from -2 to 0 (oxidation).

Step 3: Conclusion: HgO undergoes redox reaction with Hg reduced and O oxidized.

Example 2: Write oxidation and reduction half reactions for $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$

Solution:

Oxidation half reaction: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$

Reduction half reaction: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$

Overall reaction: $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$

Practice Set

Conceptual Questions

- Define oxidation and reduction with examples.
- Explain the role of oxidising and reducing agents in a redox reaction.

Application-based Question

- In the reaction $2\text{FeCl}_3 + \text{H}_2 \rightarrow 2\text{FeCl}_2 + 2\text{HCl}$, identify the oxidising and reducing agents and justify your answer.

Answer Key

Conceptual Questions

- Oxidation is the addition of oxygen or removal of hydrogen; reduction is the removal of oxygen or addition of hydrogen. Example: CH_4
- Oxidising agent accepts electrons and gets reduced; reducing agent donates electrons and gets oxidised.

Application-based Question

- Oxidising agent: FeCl_3 (Fe is reduced from +3 to +2). Reducing agent: H_2 (H is oxidised from 0 to +1).

Types of Redox Reactions, Balancing of Redox Reactions and Electrode Processes

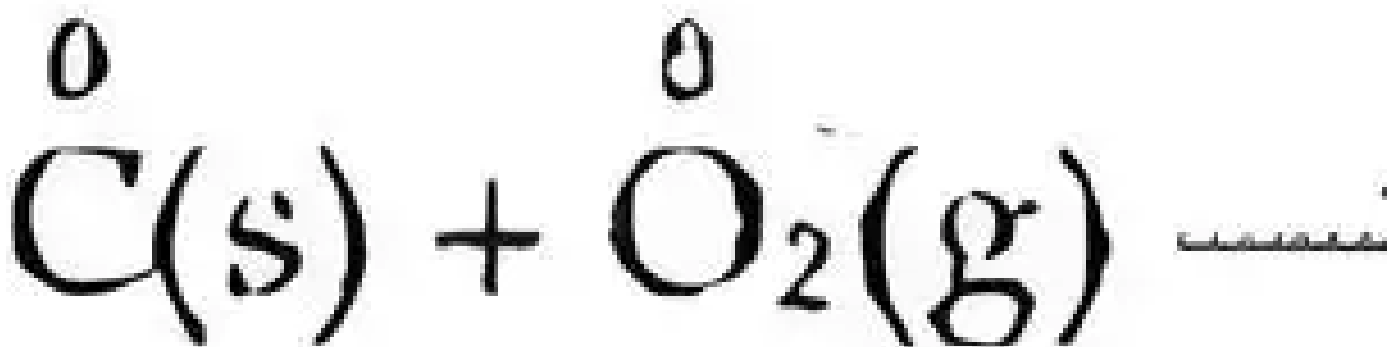
Types of Redox Reactions

1. Combination Reactions

These reactions involve the combination of two or more atoms or molecules to form a single compound.

Example:

e.g.,



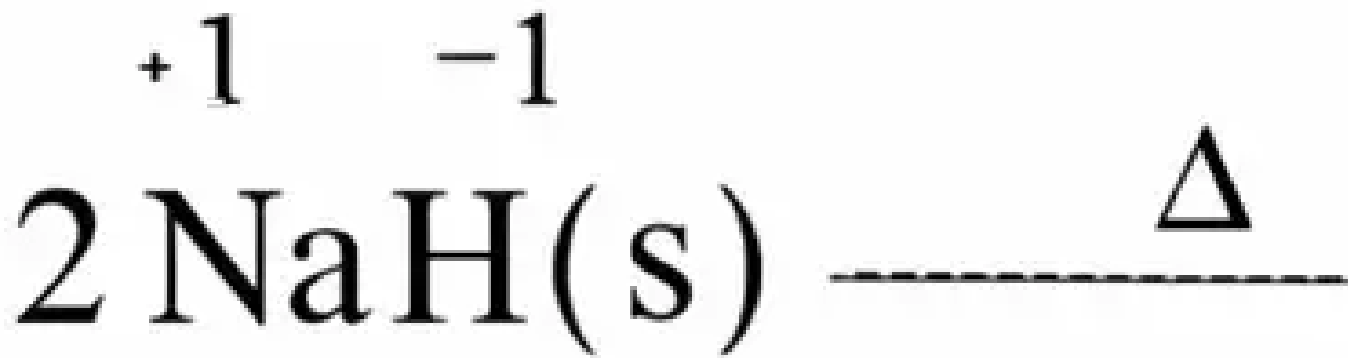
Carbon (C) reacts with oxygen (O_2) on heating to form carbon dioxide (CO_2).

2. Decomposition Reactions

These reactions involve the breakdown of a compound into two or more components, at least one of which is in

Example:

e.g.,



Sodium hydride (NaH) decomposes on heating to sodium metal (Na) and hydrogen gas (H₂).

3. Displacement Reactions

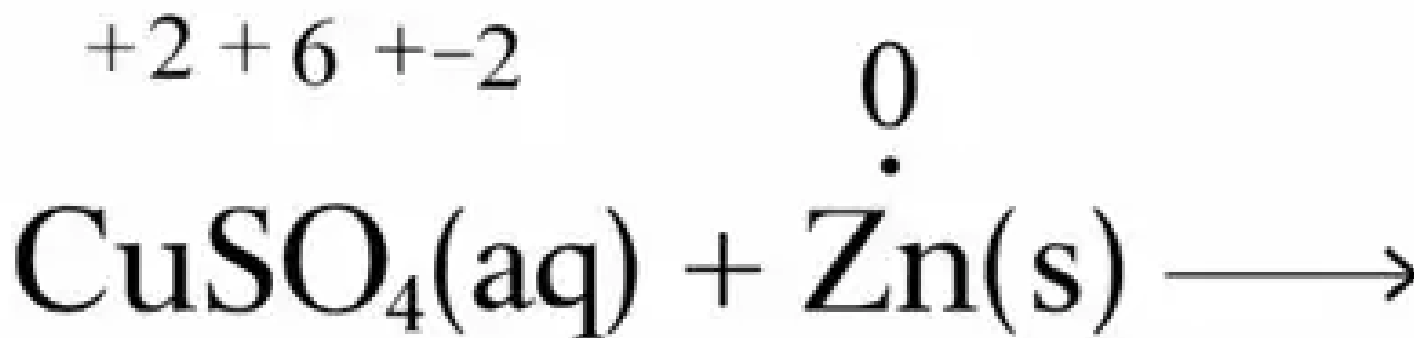
These reactions involve the replacement of an atom or ion by another atom or ion from a compound.

General form: $X + YZ \rightarrow XZ + Y$

Types:

(i) Metal Displacement Reactions

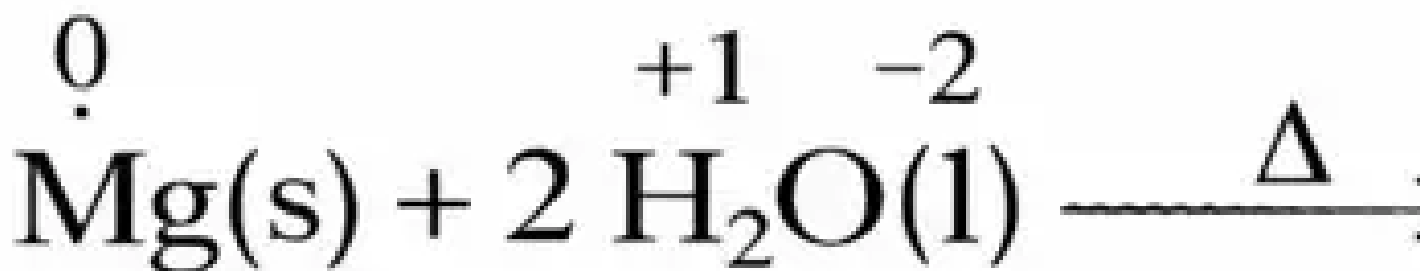
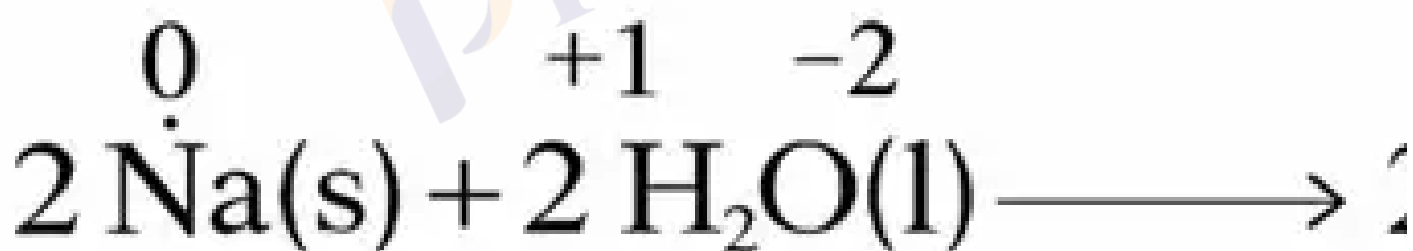
Example:



Zinc displaces copper from copper sulfate solution.

(ii) Non-metal Displacement Reactions

Example:



Sodium and magnesium react with water to produce metal hydroxides and hydrogen gas.

4. Disproportionation Reactions

In these reactions, an element in one oxidation state is simultaneously oxidised and reduced.

Examples:



Hydrogen peroxide decomposes into water and oxygen; phosphorus reacts with hydroxide and water to form phosphate.

Balancing of Redox Reactions

1. Oxidation Number Method

- Write correct formulas for reactants and products.
- Assign oxidation numbers and identify changes.
- Calculate total increase and decrease in oxidation numbers and balance them.
- Add H^+ or OH^- ions as needed for acidic or basic medium.
- Balance hydrogen atoms by adding H_2O molecules.
- Check oxygen atoms and ensure balanced equation.

2. Half Reaction Method

- Write unbalanced ionic equation.
- Separate into oxidation and reduction half reactions.
- Balance atoms other than O and H.
- Balance O by adding H_2O and H by adding H^+ (acidic medium).
- Balance charges by adding electrons.

- Multiply half reactions to equalize electrons and add them.
- For basic medium, add OH^- ions to neutralize H^+ ions.

Electrode Processes

Direct Redox Reactions

Oxidation and reduction occur in the same container, e.g., electrolysis in an electrolytic cell.

Indirect Redox Reactions

Oxidation and reduction occur in separate containers connected by a salt bridge, e.g., galvanic or electrochemical cell.

Example: Daniel Cell

$\text{Zn} / \text{Zn}^{2+}$ and $\text{Cu}^{2+} / \text{Cu}$ half cells connected by salt bridge and external circuit.

Half Reactions:

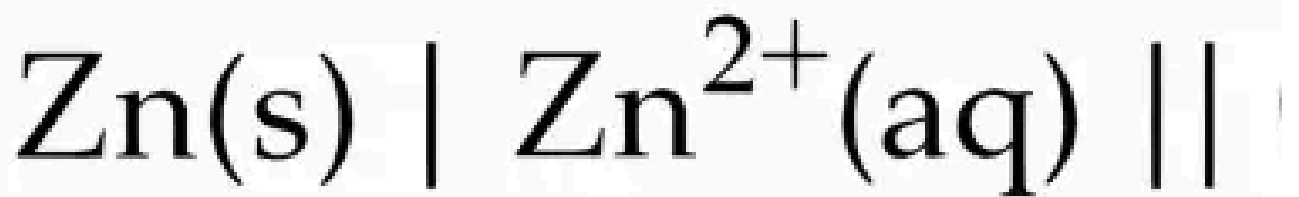
$\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ (oxidation at anode)

$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ (reduction at cathode)

Overall: $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$

Flow of electrons

Flow of current



Left
electrode
(Anode)

Salt
bridge

Oxidation
(Negative)

Electrons flow from zinc (anode) to copper (cathode) through the external circuit; salt bridge maintains charge balance.

Standard Electrode Potential (E^\ominus)

Potential of an electrode when all species are at unit concentration and gases at 1 atm pressure at 298 K.

Electrochemical Series

Arrangement of electrodes in decreasing order of standard reduction potentials.

Feasibility of Reaction

$$E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$$

Positive E_{cell} indicates feasible reaction.

Mnemonics

Concept: An Oil Rig Cat

Mnemonic: At the Anode, Oxidation Involves Loss of electrons; Reduction Involves Gaining electrons at the Cathode

Concept: Electrons flow From Anode To Cathode

Mnemonic: Fat Cat (LOAN: Left side, Oxidation, Anode, Negative)

Concept: Soak a jerk

Mnemonic: SOAC a GERC (Strongest Oxidizing Agent at the Cathode Gains Electrons and is Reduced at the Cathode)

Concept: Activity Series of Metals

Mnemonic: Please Stop Calling Me A Cute Zebra I Like Her Call Smart Goat. Kangaroos Naturally Muck About In Zinc

Order: Calcium > Magnesium > Aluminium > (Carbon) > Zinc > Iron > Lead > Hydrogen > Copper > Silver > Gold

Concept: For Galvanic Cell

Mnemonic: LOAN (Loss of electrons | Oxidation Anode Negative)

Concept: Electrolytic Cell

Mnemonic: LOAP (Loss of electrons | Oxidation Anode Positive)

Practice Set

Conceptual Questions

- Describe the four types of redox reactions with examples.
- Explain the role of salt bridge in an electrochemical cell.

Application-based Question

- Balance the redox reaction of zinc metal reacting with copper sulfate solution using the half reaction method.

Numerical Question

- Calculate the cell potential for the Daniel cell given $E^{\circ}_{\text{Zn}^{2+}/\text{Zn}} = -0.76 \text{ V}$ and $E^{\circ}_{\text{Cu}^{2+}/\text{Cu}} = +0.34 \text{ V}$.

Answer Key

Conceptual Questions

- Combination: $A + B \rightarrow AB$ (e.g., $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$). Decomposition: $AB \rightarrow A + B$ (e.g., $\text{NaH} \rightarrow \text{Na} + \text{H}_2$). Displacement: $X + \text{YZ} \rightarrow \text{XZ} + \text{Y}$ (e.g., $\text{Zr} + \text{O}_2$).
- Salt bridge completes the circuit and maintains electrical neutrality by allowing ion flow without mixing solutions.

Application-based Question

- Oxidation half reaction: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$; Reduction half reaction: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$; Balanced overall: $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$.

Numerical Question

- $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{Cu}^{2+}/\text{Cu}} - E^{\circ}_{\text{Zn}^{2+}/\text{Zn}} = 0.34 \text{ V} - (-0.76 \text{ V}) = +1.10 \text{ V}$; Positive value indicates feasible reaction.

Quick Reference Table

Common Mistakes and Misconceptions

Glossary
