# Table of Contents

1. Introduction
2. Classical Idea of Redox Reactions-Oxidation and Reduction Reactions
3. Redox Reactions Electron Transfer
4. Competitive Electron Transfer Reactions
5. Oxidation Number
6. Oxidation Number of Important Compounds
7. Types of Redox Reactions
8. Balancing of Redox Reactions
9. Redox Reactions as the Basis for Titrations
10. Redox Reactions and Electrode Processes
11. Predicting Reducing and Oxidising Power

# Introduction

## Overview of Redox Reactions

- Chemistry studies matter and its transformations via reactions.
- Redox reactions involve transfer of electrons, changing matter from one form to another.
- Redox processes occur in physical, biological, industrial, pharmaceutical, metallurgical, and agricultural fields.
- Environmental phenomena like hydrogen fuel use and ozone depletion involve redox reactions.

## Solved Examples

- **Ex 1:** Identify if the reaction $2Mg + O_2 \rightarrow 2MgO$ is redox.
  Step 1: Mg gains oxygen → oxidation.
  Step 2: $O_2$ is reduced.
  Conclusion: Both oxidation and reduction occur simultaneously, so it is a redox reaction.

## Practice Set

- **Q1:** Which of the following is a redox reaction?
  A. NaCl dissolution in water
  B. Combustion of methane
  C. Melting of ice
  D. Boiling of water
- **Q2:** Redox reactions are important in which of the following fields?
  A. Metallurgy
  B. Agriculture
  C. Pharmaceuticals
  D. All of the above
- **Q3:** Which environmental issue involves redox reactions?
  A. Acid rain
  B. Ozone hole formation
  C. Global warming
  D. Soil erosion

### Answer Key

- **Q1:** B - Combustion involves oxidation and reduction.
- **Q2:** D - Redox reactions are used extensively in all listed fields.
- **Q3:** B - Ozone hole formation is a redox phenomenon.

### Memory Aids

- Redox = Reduction + Oxidation happening together.
- Remember: OIL RIG - Oxidation Is Loss, Reduction Is Gain (of electrons).
- Environmental redox: Think hydrogen fuel and ozone layer.

## Classical Idea of Redox Reactions-Oxidation and Reduction Reactions

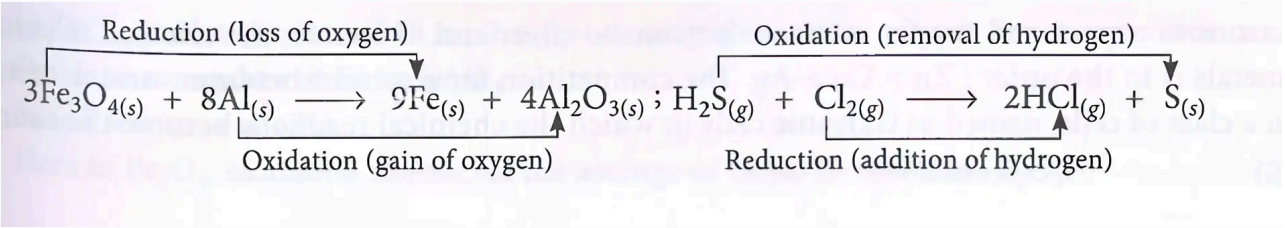
### Oxidation

- Oxidation: Addition of oxygen or electronegative element to a substance.
- Or removal of hydrogen or electropositive element from a substance.
- Examples:
  - $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$
  - $2H_2S_{(g)} + O_{2(g)} \rightarrow 2S_{(s)} + 2H_2O_{(l)}$
  - $Mg_{(s)} + F_{2(g)} \rightarrow MgF_{2(s)}$
  - $2K_4[Fe(CN)_6]_{(aq)} + H_2O_{2(aq)} \rightarrow 2K_3[Fe(CN)_6]_{(aq)} + 2KOH_{(aq)}$

### Reduction

- Reduction: Removal of oxygen or electronegative element from a substance.
- Or addition of hydrogen or electropositive element to a substance.
- Examples:
  - $2HgO_{(s)} \xrightarrow{\Delta} 2Hg_{(l)} + O_{2(g)}$ ; $2FeCl_{3(aq)} + H_{2(g)} \longrightarrow 2FeCl_{2(aq)} + 2HCl_{(aq)}$
  - $CH_2{=}CH_{2(g)} + H_{2(g)} \longrightarrow H_3C{-}CH_{3(g)}$
  - $2HgCl_{2(aq)} + SnCl_{2(aq)} \longrightarrow Hg_2Cl_{2(s)} + SnCl_{4(aq)}$

## Redox Reactions

- Redox reactions involve simultaneous oxidation and reduction.
- Examples:
  - 

    - Oxidation: Loss of electrons (gain of oxygen or loss of hydrogen).
    - Reduction: Gain of electrons (loss of oxygen or gain of hydrogen).

## Solved Examples

- **Ex 1:** Identify oxidation and reduction in $2Mg + O_2 \rightarrow 2MgO$.
  Step 1: Mg gains oxygen → oxidation.
  Step 2: $O_2$ loses oxygen → reduction.
- **Ex 2:** In $Fe_3O_4 + Al \rightarrow Al_2O_3 + Fe$, identify oxidized and reduced species.
  Step 1: Fe gains electrons (reduction).
  Step 2: Al loses electrons (oxidation).

## Practice Set

- **Q1:** Which is oxidation in classical terms?
  A. Addition of hydrogen
  B. Removal of oxygen
  C. Addition of oxygen
  D. Removal of electronegative element
- **Q2:** In the reaction $FeCl_3 + H_2 \rightarrow FeCl_2 + HCl$, what happens to Fe?
  A. Oxidized
  B. Reduced
  C. No change
  D. Decomposed
- **Q3:** Which reaction is not a redox reaction?
  A. $2Mg + O_2 \rightarrow 2MgO$
  B. $H_2 + Cl_2 \rightarrow 2HCl$

C. NaCl dissolution in water

D. $Fe_3O_4 + Al \rightarrow Al_2O_3 + Fe$

### Answer Key

- **Q1:** C - Oxidation is addition of oxygen.
- **Q2:** B - Fe is reduced by gaining electrons.
- **Q3:** C - Dissolution is not a redox reaction.

### Memory Aids

- Oxidation: OIL - Oxidation Is Loss (of electrons).
- Reduction: RIG - Reduction Is Gain (of electrons).
- Oxidation = add oxygen or remove hydrogen.
- Reduction = remove oxygen or add hydrogen.

## Redox Reactions Electron Transfer

### Concepts

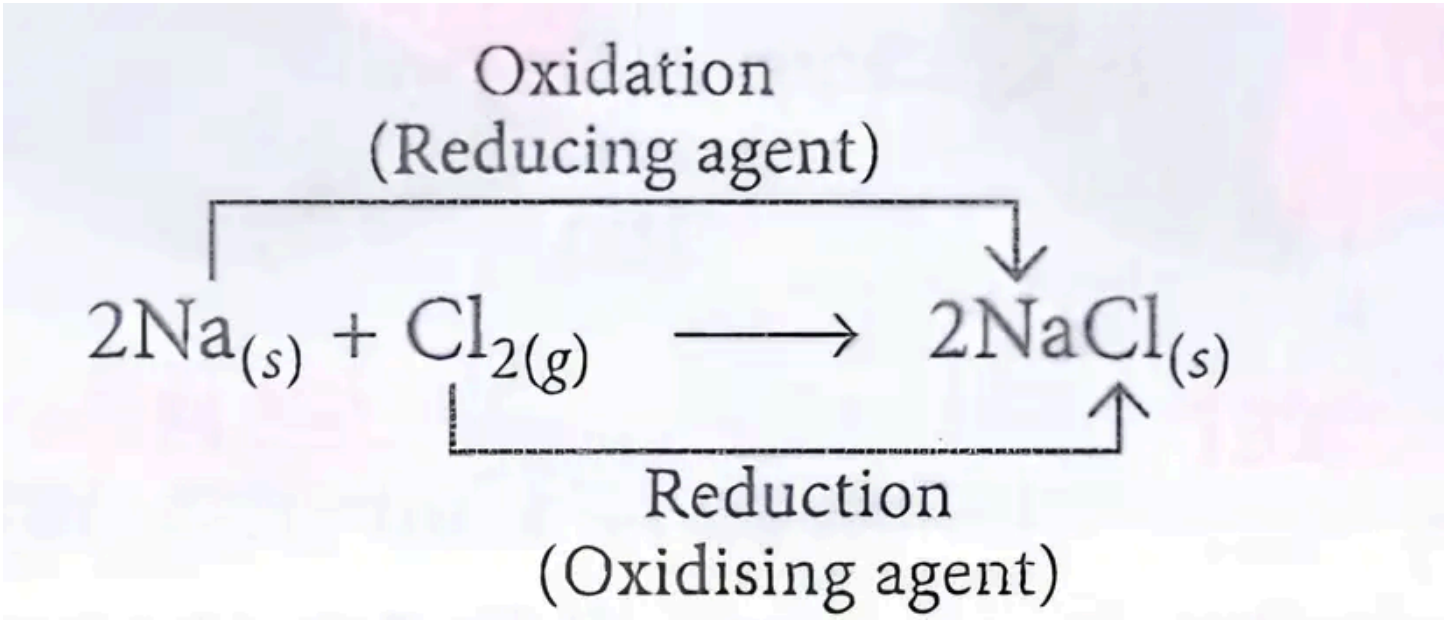
- Oxidation: Loss of electrons in half reactions.
- Reduction: Gain of electrons in half reactions.
- Oxidising agent (oxidant): Removes electrons from other substances and gets reduced itself.
- Example: Nitrate ion oxidises Cu to $Cu^{2+}$:

  $Cu_{(s)} + 4H^+_{(aq)} + 2NO_3^-{}_{(aq)} \square 1\ Cu^{2+}_{(aq)} + 2H_2O_{(l)} + 2NO_{2(g)}$

- Reducing agent (reductant): Donates electrons to reduce another substance and gets oxidised itself.
- Example: $Sn^{2+}$ reduces $Fe^{3+}$:

  $Sn^{2+}_{(aq)} + 2Fe^{3+}_{(aq)} \square 1\ Sn^{4+}_{(aq)} + 2Fe^{2+}_{(aq)}$

- Example reaction: Sodium (Na) oxidised, Chlorine ($Cl_2$) reduced forming NaCl.




## Solved Examples

- Ex 1: Identify oxidising and reducing agents in $Cu + 4H^+ + 2NO_3^-$ reaction.
  - Oxidising agent: $NO_3^-$ (gains electrons, reduced)
  - Reducing agent: Cu (loses electrons, oxidised)
- Ex 2: In $Sn^{2+} + 2Fe^{3+}$ reaction, $Sn^{2+}$ is reducing agent, $Fe^{3+}$ is oxidising agent.

## Practice Set

- Q1: Which is oxidised in $Cu + 4H^+ + 2NO_3^-$ reaction?
  - A. Cu
  - B. $NO_3^-$
  - C. $H^+$
  - D. $H_2O$
- Q2: Reducing agent donates:
  - A. Protons
  - B. Electrons
  - C. Neutrons
  - D. Photons
- Q3: In $Sn^{2+} + 2Fe^{3+}$ reaction, $Fe^{3+}$ is:
  - A. Oxidised
  - B. Reduced
  - C. Neither

- D. Both

### Answer Key

- Q1: A - Cu loses electrons (oxidised).
- Q2: B - Reducing agent donates electrons.
- Q3: B - $Fe^{3+}$ gains electrons (reduced).

### Memory Aids

- Oxidation = OIL (Oxidation Is Loss of electrons)
- Reduction = RIG (Reduction Is Gain of electrons)
- Oxidising agent gets reduced; reducing agent gets oxidised.

## Competitive Electron Transfer Reactions

### Concepts

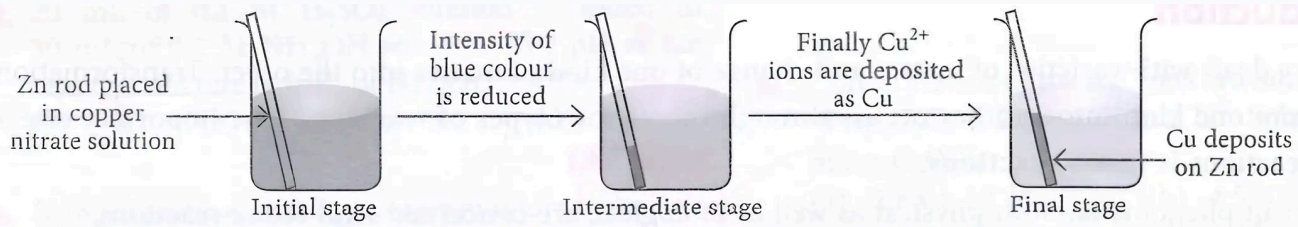
- Zn reacts with $Cu^{2+}$ solution: Zn oxidised to $Zn^{2+}$, $Cu^{2+}$ reduced to Cu metal.

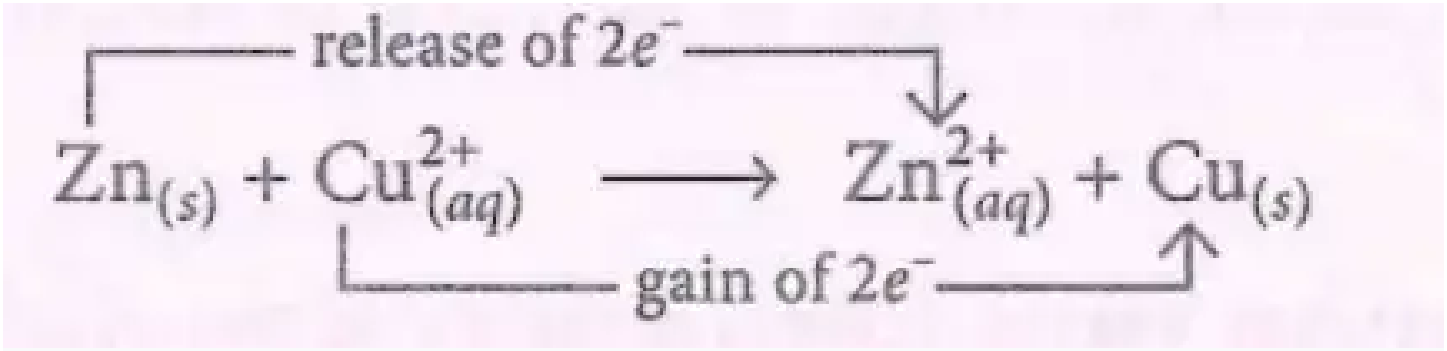
  $Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$

- Cu reacts with $Ag^{+}$ solution: Cu oxidised to $Cu^{2+}$, $Ag^{+}$ reduced to Ag metal.

  $Cu_{(s)} + 2Ag^{+}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$

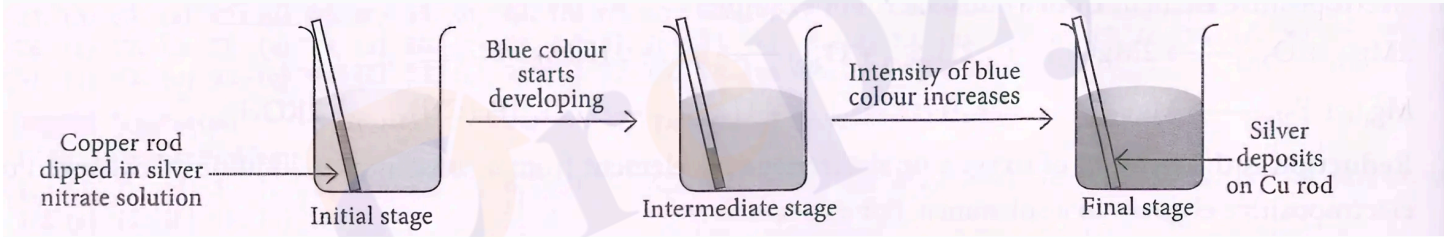
- Electron releasing tendency order: Zn > Cu > Ag.
- These reactions demonstrate displacement and electron transfer, basis for Galvanic cells.






## Solved Examples

- Ex 1: Write half reactions for Zn + $Cu^{2+}$ reaction.
  - Oxidation: Zn □1 $Zn^{2+}$ + $2e^-$
  - Reduction: $Cu^{2+}$ + $2e^-$ □1 Cu
- Ex 2: Identify oxidising and reducing agents in Cu + $2Ag^+$ reaction.
  - Oxidising agent: $Ag^+$
  - Reducing agent: Cu




## Practice Set

- Q1: In Zn + $Cu^{2+}$ reaction, which is oxidised?
  - A. Zn
  - B. $Cu^{2+}$
  - C. Cu
  - D. $Zn^{2+}$
- Q2: Electron releasing tendency order is:
  - A. Ag > Cu > Zn
  - B. Zn > Cu > Ag
  - C. Cu > Zn > Ag
  - D. Zn > Ag > Cu
- Q3: In Cu + $2Ag^+$ reaction, silver deposits because:
  - A. Ag is oxidised

- B. $Ag^+$ is reduced
- C. Cu is reduced
- D. $Cu^{2+}$ is reduced

### Answer Key

- Q1: A - Zn loses electrons (oxidised).
- Q2: B - Zn has highest electron releasing tendency.
- Q3: B - $Ag^+$ gains electrons and deposits as Ag metal.

### Memory Aids

- Electron releasing order: Zn > Cu > Ag (Zinc displaces copper, copper displaces silver).
- Displacement reactions show redox and electron transfer.
- Basis for Galvanic cells producing electrical energy.

## Oxidation Number

### Rules for Assigning Oxidation Number

- Oxidation number is the charge an atom appears to have in a compound.
- Elements in free state have oxidation number zero.
- For monoatomic ions, oxidation number equals the ion charge.
- Oxygen is usually -2; exceptions: peroxides (-1), superoxides (-1/2), $OF_2$ (+2), $O_2F_2$ (+1).
- Hydrogen is +1; in metal hydrides, it is -1.
- Fluorine is always -1 except in HOF (+1).
- Sum of oxidation numbers in a neutral compound is zero.
- Sum of oxidation numbers in a polyatomic ion equals the ion charge.
- **Fractional oxidation state**: average oxidation state when atoms have different states.
- **Stock notation**: oxidation state shown by Roman numerals in parentheses, e.g., iron(II) sulfate for $FeSO_4$.
- Oxidation: increase in oxidation number; Reduction: decrease in oxidation number.
- **Oxidising agent**: increases oxidation number (oxidant).
- **Reducing agent**: decreases oxidation number (reductant).

- **Redox reactions**: involve change in oxidation numbers.

## Solved Examples

**Ex 1:** Find oxidation number of Fe in $[Fe(H_2O)_5(NO)]SO_4$.

- Let Fe = x.
- 5 waters: 5 × 0 = 0, NO = +1, $SO_4$ = -2.
- Equation: x + 0 + 1 - 2 = 0 → x - 1 = 0 → x = +1.

**Ex 2:** Oxidation number of Fe in $Fe_3O_4$.

- Let Fe = x.
- 3x + 4(-2) = 0 → 3x - 8 = 0 → x = +8/3 (fractional state).
- Average of FeO (+2) and $Fe_2O_3$ (+3).

## Practice Set

- Q1: What is the oxidation number of oxygen in peroxides?
  - A. -2
  - B. -1
  - C. -1/2
  - D. +2
- Q2: In $FeSO_4$ (iron(II) sulfate), the oxidation state of Fe is:
  - A. +1
  - B. +2
  - C. +3
  - D. 0
- Q3: Which of the following is a reducing agent?
  - A. $MnO_4^-$
  - B. Zn
  - C. $O_2$
  - D. $F_2$

### Answer Key

- Q1: B. Oxygen in peroxides has oxidation number -1.
- Q2: B. Fe in $FeSO_4$ is +2 (iron(II) sulfate).
- Q3: B. Zn is a reducing agent as it lowers oxidation number of other elements.

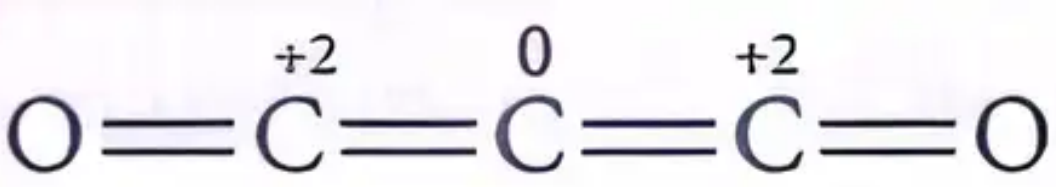
### Memory Aids

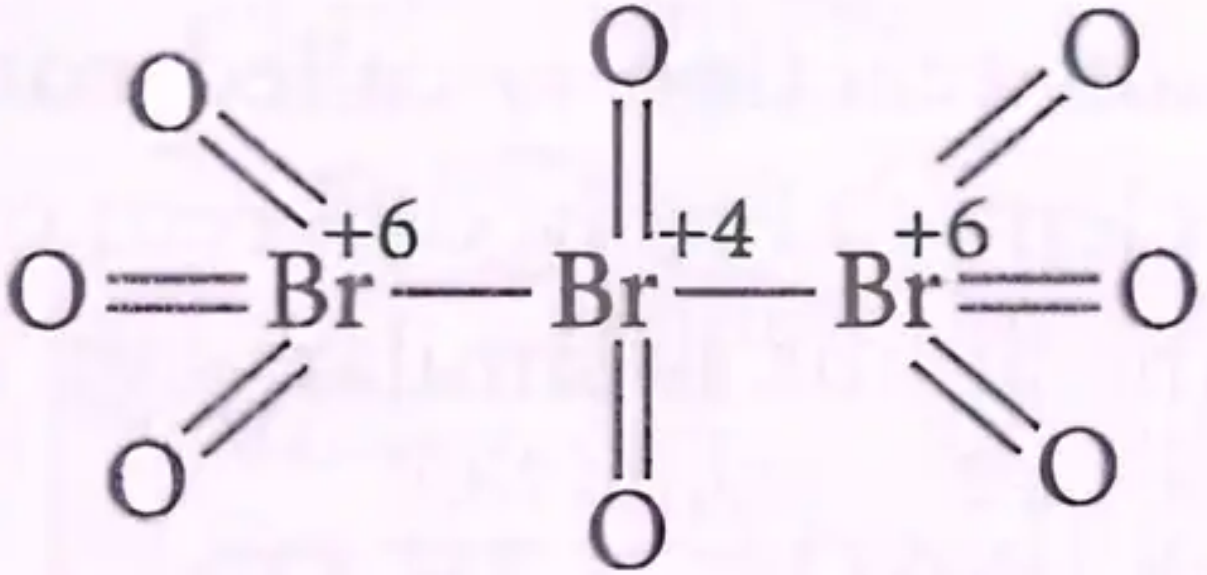
- Oxygen exceptions: peroxides (-1), superoxides (-1/2), $OF_2$ (+2).
- Hydrogen is +1 except in metal hydrides (-1).
- Fluorine always -1 except in HOF (+1).
- Stock notation uses Roman numerals for metal oxidation states.

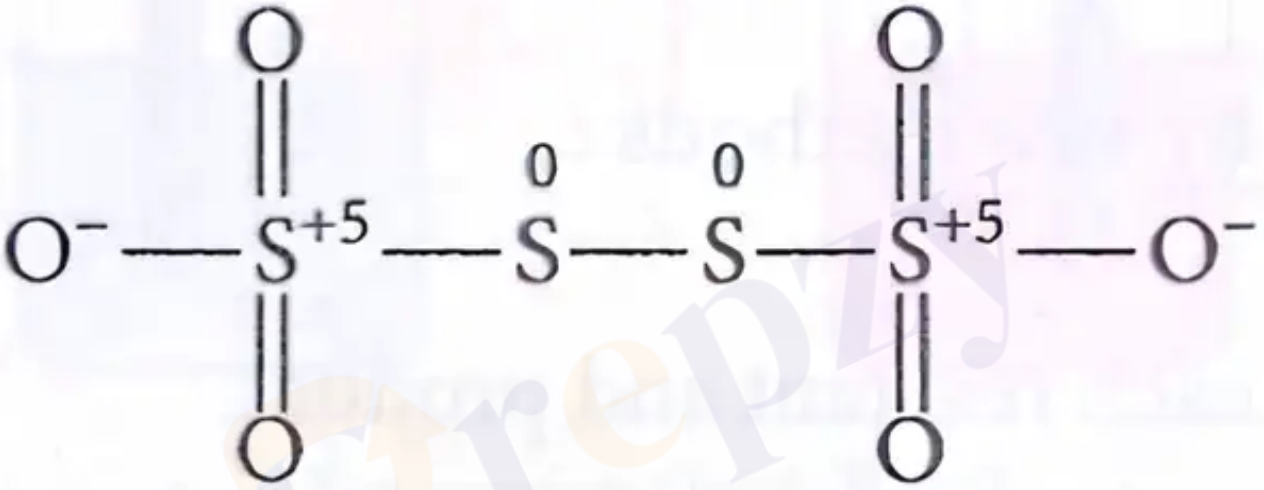
## Oxidation Number of Important Compounds

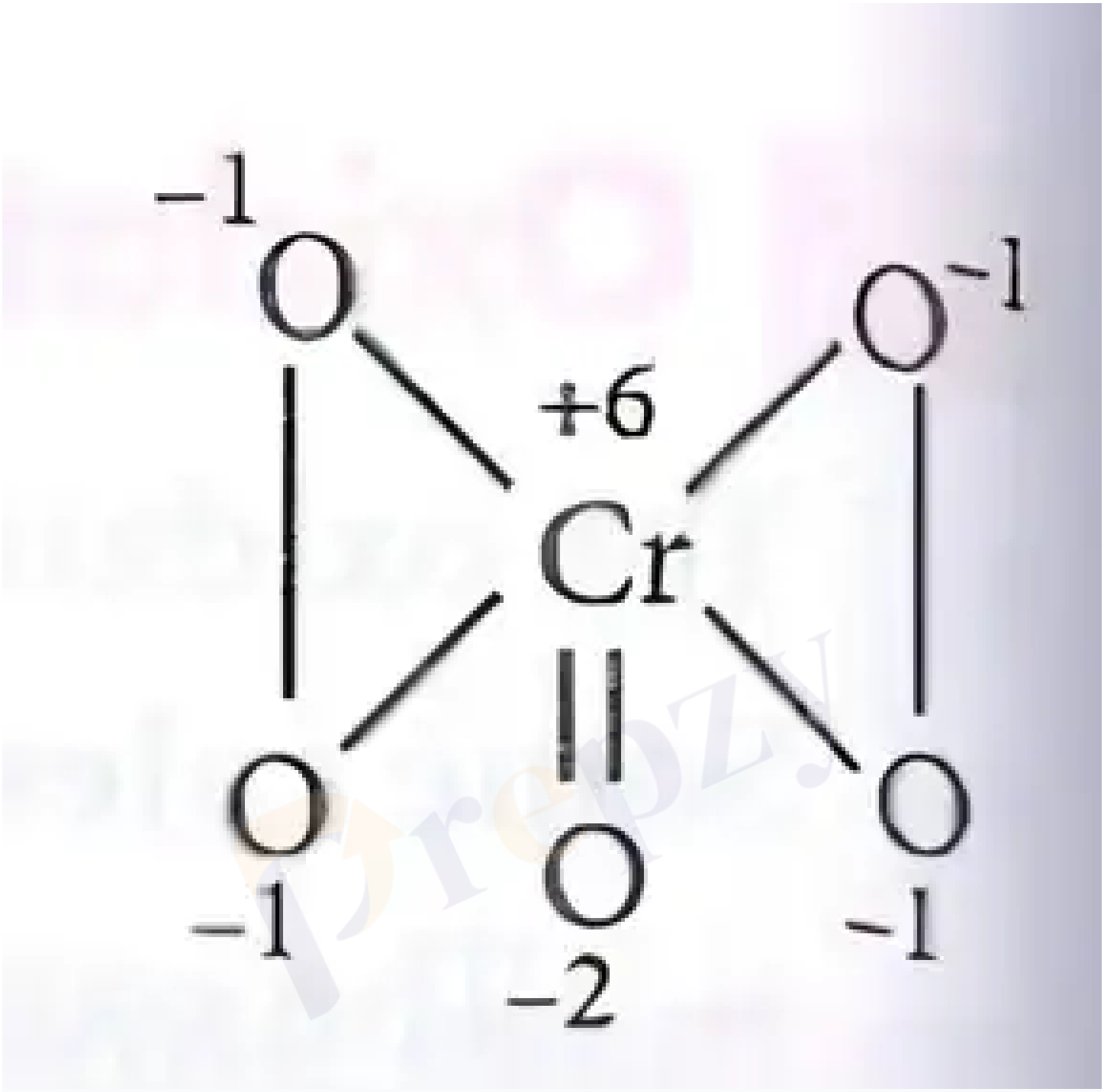
### Key Examples

- **C in $C_3O_2$ (carbon suboxide):** Two carbons +2, middle carbon 0.
- **Br in $Br_3O_4$ (tribromooctaoxide):** Two Br +6, one Br +4.
- **S in $S_4O_6^{2-}$ (tetrathionate ion):** Two S +5, two S 0.
- **Cl in Ca(OCl)Cl:** Cl in $OCl^-$ is +1; other Cl is -1.
- **Cr in $CrO_5$:** Cr is +6; peroxo oxygen -1, normal oxygen -2.




O
O
+6
O
Br
Br
+4
Br
+6
O
O
O
O
O
O


O
O
O⁻
S+5
S
S
S+5
O⁻
O
O




## Solved Examples

**Ex 1:** Calculate oxidation number of Cl in Ca(OCl)Cl.

- In $OCl^-$, Cl = y; -2 + y = -1 → y = +1.
- Sum: +2 + (-2) + 1 + x = 0 → x = -1 (other Cl).

## Practice Set

- Q1: Oxidation state of Cr in $CrO_5$ is:
  - A. +4
  - B. +5
  - C. +6
  - D. +7
- Q2: In $S_4O_6^{2-}$, how many S atoms have zero oxidation state?
  - A. 1
  - B. 2
  - C. 3
  - D. 4
- Q3: In $Br_3O_4$, the oxidation state of the middle Br atom is:
  - A. +4
  - B. +5
  - C. +6
  - D. 0

## Answer Key

- Q1: C. Cr in $CrO_5$ is +6 due to peroxo bonds.
- Q2: B. Two S atoms have zero oxidation state in $S_4O_6^{2-}$.
- Q3: A. Middle Br in $Br_3O_4$ is +4.

## Memory Aids

- $CrO_5$ has peroxo bonds; peroxo oxygen is -1.
- Tetrathionate ion has mixed S oxidation states (+5 and 0).
- Br in $Br_3O_4$ varies: outer +6, middle +4.

# Types of Redox Reactions

## Combination Reactions

- Two atoms/molecules combine to form one compound.
- At least one reactant is in elemental form.
- Example: $H_2 + Cl_2 \rightarrow 2HCl$ (H: 0 to +1, Cl: 0 to -1).
- Example: $H_2 + O_2 \rightarrow 2H_2O$ (H: 0 to +1, O: 0 to -2).

$$\overset{0}{H_2} + \overset{0}{Cl_2} \rightarrow 2\overset{+1}{H}\overset{-1}{Cl} \quad ; \quad 2\overset{0}{H_2} + \overset{0}{O_2} \rightarrow 2\overset{+1}{H_2}\overset{-2}{O}$$

## Decomposition Reactions

- Compound breaks down into two or more products.
- Reverse of combination reaction.
- For redox, one product must be elemental form.
- Example: $2H_2O \rightarrow 2H_2 + O_2$ (on heating).
- Example: $2LiNO_3 \rightarrow Li_2O + 2NO_2 + O_2$ (on heating).

$$2\overset{+1}{H_2}\overset{-2}{O} \xrightarrow{\Delta} 2\overset{0}{H_2} + \overset{0}{O_2} \quad ; \quad 4\overset{+1}{Li}\overset{+5}{N}\overset{-2}{O_3} \xrightarrow{\Delta} 2Li_2O + 4\overset{+4}{N}\overset{-2}{O_2} + \overset{0}{O_2}$$

## Displacement Reactions

- One atom/ion replaces another in a compound: $AB + C \rightarrow AC + B$.
- **Metal displacement:** metal replaces metal ion.
- Example: $TiCl_4 + Mg \rightarrow Ti + MgCl_2$ (Ti: +4 to 0, Mg: 0 to +2).
- Example: $CuSO_4 + Zn \rightarrow Cu + ZnSO_4$ (Cu: +2 to 0, Zn: 0 to +2).

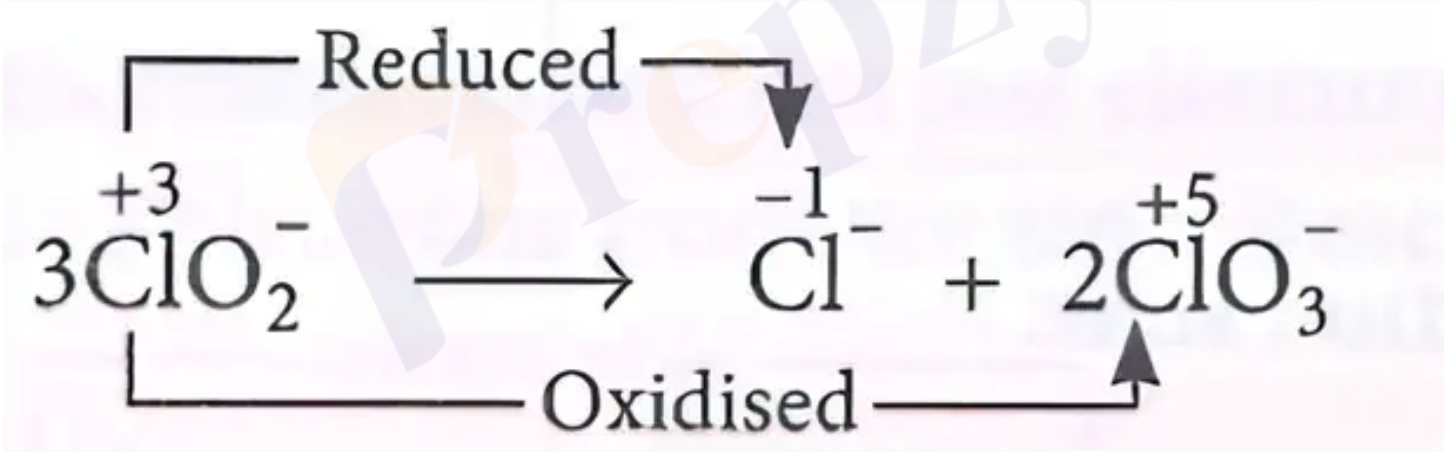
$$\overset{+4}{Ti}Cl_4 + 2\overset{0}{Mg} \rightarrow \overset{0}{Ti} + 2\overset{+2}{Mg}Cl_2 \quad ; \quad \overset{+2}{Cu}\overset{+6}{S}\overset{-2}{O_4} + \overset{0}{Zn} \rightarrow \overset{0}{Cu} + \overset{+2}{Zn}\overset{+6}{S}\overset{-2}{O_4}$$

- **Non-metal displacement:** non-metal replaces metal or non-metal ion.
- Example: $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$ (Zn oxidized, H reduced).
- Example: $Al + H_2O \rightarrow Al_2O_3 + H_2$ (Al oxidized, H reduced).
- Example: $Cl_2 + KBr \rightarrow KCl + Br_2$ (Cl reduced, Br oxidized).

$$\overset{0}{Zn} + \overset{+1}{H_2}SO_4 \rightarrow \overset{+2}{Zn}SO_4 + \overset{0}{H_2} \quad ; \quad \overset{0}{2Al} + 3\overset{+1}{H_2}O \rightarrow \overset{+3}{Al_2}O_3 + 3\overset{0}{H_2} \; ; \quad \overset{0}{Cl_2} + 2\overset{+1\,-1}{KBr} \rightarrow 2\overset{+1-1}{KCl} + \overset{0}{Br_2}$$

## Disproportionation Reactions

- Single species is simultaneously oxidized and reduced.
- Element must have at least three variable oxidation states.
- Example: $3ClO_2^- \rightarrow 2ClO_3^- + Cl^-$ (Cl: +3 to +5 and -1).
- Reverse is comproportionation: two reactants with different oxidation states form product with same oxidation state.
- Example: $Pb + PbO_2 + H_2SO_4 \rightarrow PbSO_4 + H_2O$ (lead storage battery).




$$\overset{0}{Pb_{(s)}} + \overset{+4}{PbO_{2(s)}} + 2H_2SO_{4(aq)} \longrightarrow \overset{+2}{2PbSO_{4(s)}} + 2H_2O_{(l)}$$

## Practice Set

- Q1: Which reaction is a combination reaction?
  - A. $2H_2 + O_2 \rightarrow 2H_2O$
  - B. $2H_2O \rightarrow 2H_2 + O_2$
  - C. $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
  - D. $3ClO_2^- \rightarrow 2ClO_3^- + Cl^-$

- Q2: In which reaction does disproportionation occur?
  - A. $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
  - B. $3ClO_2^- \rightarrow 2ClO_3^- + Cl^-$
  - C. $TiCl_4 + Mg \rightarrow Ti + MgCl_2$
  - D. $2LiNO_3 \rightarrow Li_2O + 2NO_2 + O_2$
- Q3: In the reaction $TiCl_4 + Mg \rightarrow Ti + MgCl_2$, the oxidation state of Ti changes from:
  - A. 0 to +4
  - B. +4 to 0
  - C. +2 to +3
  - D. 0 to -2

### Answer Key

- Q1: A. Combination reaction forms one product from elements.
- Q2: B. Disproportionation involves simultaneous oxidation and reduction of same species.
- Q3: B. Ti reduces from +4 to 0 by gaining electrons.

### Memory Aids

- Combination: elements combine to form compound.
- Decomposition: compound breaks into elements or simpler compounds.
- Displacement: one element replaces another.
- Disproportionation: same element oxidized and reduced.

## Balancing of Redox Reactions

### Oxidation Number Method

- Write correct formulas of reactants and products.
- Assign oxidation numbers to all atoms.
- Identify atoms with change in oxidation number.
- Calculate total increase and decrease in oxidation numbers; equalize by multiplying.
- Add $H^+$ (acidic) or $OH^-$ (basic) to balance charges.
- Balance hydrogen by adding $H_2O$; check oxygen atoms.

## Half Reaction Method

- Write ionic form of reaction.
- Separate into oxidation and reduction half reactions.
- Balance atoms except O and H in each half.
- In acidic medium: add $H_2O$ for O, $H^+$ for H; in basic medium: add $H_2O$ and $OH^-$ accordingly.
- Add electrons to balance charge; multiply half reactions to equalize electrons.
- Add half reactions; cancel electrons to get balanced equation.

## Practice Set

- Q1: In acidic medium, to balance oxygen atoms in half reaction, add:
  - A. $H_2O$
  - B. $OH^-$
  - C. $H^+$
  - D. Electrons
- Q2: In half reaction method, electrons are added to:
  - A. Oxidation half only
  - B. Reduction half only
  - C. Both halves to balance charge
  - D. Neither half
- Q3: Which step is NOT part of oxidation number method?
  - A. Assign oxidation numbers
  - B. Add $H^+$ or $OH^-$
  - C. Separate into half reactions
  - D. Balance hydrogen and oxygen

## Answer Key

- Q1: A. $H_2O$ is added to balance oxygen in acidic medium.
- Q2: C. Electrons are added to balance charge in both half reactions.
- Q3: C. Separating into half reactions is not part of oxidation number method.

### Memory Aids

- Oxidation number method focuses on changes in oxidation states.
- Half reaction method splits reaction into oxidation and reduction parts.
- Use $H^+$ in acidic and $OH^-$ in basic medium for balancing.

## Redox Reactions as the Basis for Titrations

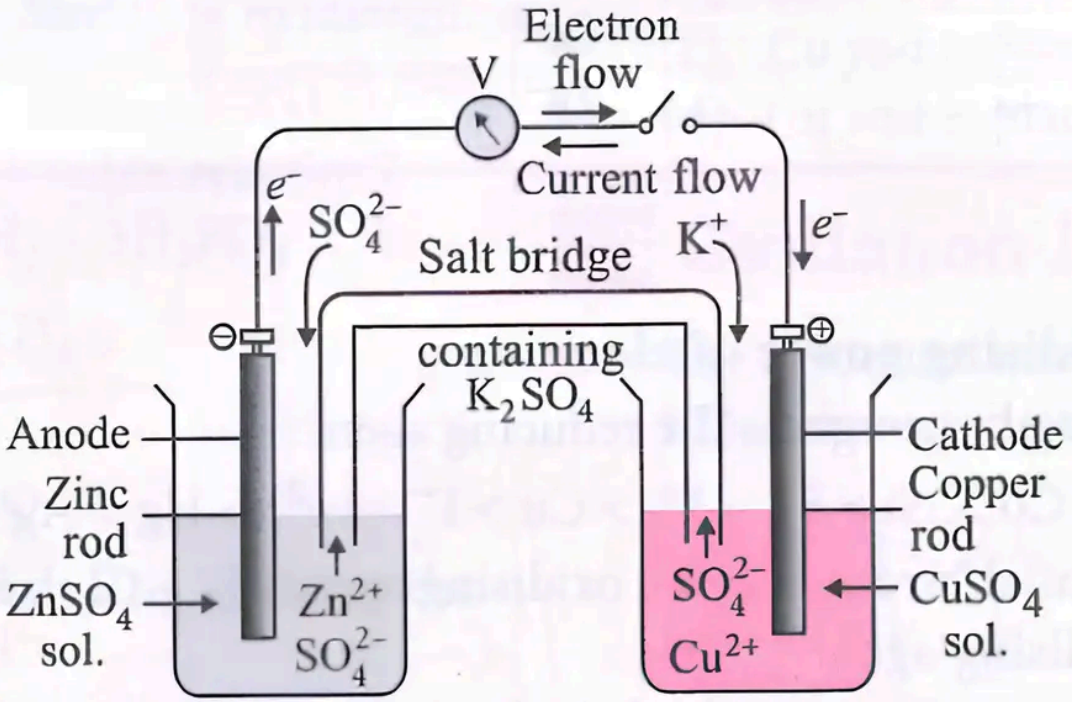
### Redox Titrations

- Used to determine strength of oxidants or reductants.
- Indicators show end point by color change.
- Some reagents (e.g., $MnO_4^-$) act as self-indicators due to intense color.
- Other indicators (e.g., diphenylamine) change color when oxidized at equivalence point.
- Starch is used as indicator for reactions involving $I^-$ or $I_2$, giving intense blue color.

## Redox Reactions and Electrode Processes

### Key Concepts

- **Redox couple**: Oxidised and reduced forms of a substance in a half-reaction, e.g., $Zn^{2+}/Zn$, $Cu^{2+}/Cu$ in Daniell cell.
- **Salt bridge**: Completes circuit, allows ion flow without mixing solutions, maintains electrical neutrality.
- **Electrode potential**: Potential difference at metal/ion interface; tendency to gain or lose electrons.
- **Standard electrode potential (EB0)**: Electrode potential at 1 M concentration, 298 K; hydrogen electrode EB0 = 0.00 V.
- **Electrochemical series**: List of reducing agents in decreasing strength; negative EB0 means stronger reducing agent than $H^+/H_2$.




Set-up for Daniell cell

Setup of Daniell cell: Zn rod in $ZnSO_4$, Cu rod in $CuSO_4$. Zn oxidises to $Zn^{2+}$, electrons flow to Cu cathode where $Cu^{2+}$ reduces to Cu metal. Salt bridge maintains neutrality. Electric current flows through external circuit.

Standard reduction potentials at 298 K (Electrochemical series)

| Reduction half-reaction | | | Standard reduction potential $E^\circ$ (in volts) |
|---|---|---|---|
| $Li^+ + e^-$ | $\rightarrow$ | Li | − 3.05 |
| $K^+ + e^-$ | $\rightarrow$ | K | − 2.93 |
| $Ba^{2+} + 2e^-$ | $\rightarrow$ | Ba | − 2.90 |
| $Ca^{2+} + 2e^-$ | $\rightarrow$ | Ca | − 2.87 |
| $Na^+ + e^-$ | $\rightarrow$ | Na | − 2.71 |
| $Mg^{2+} + 2e^-$ | $\rightarrow$ | Mg | − 2.36 |
| $Al^{3+} + 3e^-$ | $\rightarrow$ | Al | − 1.66 |
| $Mn^{2+} + 2e^-$ | $\rightarrow$ | Mn | − 1.18 |
| $2H_2O_{(l)} + 2e^-$ | $\rightarrow$ | $H_{2(g)} + 2OH^-_{(aq)}$ | − 0.83 |
| $Zn^{2+} + 2e^-$ | $\rightarrow$ | Zn | − 0.76 |
| $Cr^{3+} + 3e^-$ | $\rightarrow$ | Cr | − 0.74 |
| $Fe^{2+} + 2e^-$ | $\rightarrow$ | Fe | − 0.44 |
| $Cd^{2+} + 2e^-$ | $\rightarrow$ | Cd | − 0.40 |
| $Co^{2+} + 2e^-$ | $\rightarrow$ | Co | − 0.28 |
| $Ni^{2+} + 2e^-$ | $\rightarrow$ | Ni | − 0.25 |
| $Sn^{2+} + 2e^-$ | $\rightarrow$ | Sn | − 0.14 |
| $Pb^{2+} + 2e^-$ | $\rightarrow$ | Pb | − 0.13 |
| $2H^+ + 2e^-$ | $\rightarrow$ | $H_2$ | 0.00 |
| $Sn^{4+} + 2e^-$ | $\rightarrow$ | $Sn^{2+}$ | + 0.15 |
| $Cu^{2+} + 2e^-$ | $\rightarrow$ | Cu | + 0.34 |
| $Cu^+ + e^-$ | $\rightarrow$ | Cu | + 0.52 |
| $I_2 + 2e^-$ | $\rightarrow$ | $2I^-$ | + 0.54 |
| $Fe^{3+} + e^-$ | $\rightarrow$ | $Fe^{2+}$ | + 0.77 |
| $Ag^+ + e^-$ | $\rightarrow$ | Ag | + 0.80 |
| $Br_2 + 2e^-$ | $\rightarrow$ | $2Br^-$ | + 1.09 |
| $O_{2(g)} + 4H^+_{(aq)} + 4e^-$ | $\rightarrow$ | $2H_2O$ | + 1.23 |
| $Cl_2 + 2e^-$ | $\rightarrow$ | $2Cl^-$ | + 1.36 |
| $Au^{3+} + 3e^-$ | $\rightarrow$ | Au | + 1.40 |
| $Co^{3+} + e^-$ | $\rightarrow$ | $Co^{2+}$ | + 1.82 |
| $F_2 + 2e^-$ | $\rightarrow$ | $2F^-$ | + 2.87 |

Increasing strength of oxidising agent ↓

Increasing strength of reducing agent ↑

Standard reduction potentials table ranks substances by electron gain ability. Negative EB0: stronger oxidising agent; positive EB0: stronger reducing agent.

## Solved Examples

- **Ex 1:** Identify anode and cathode in Daniell cell.
  - Anode: Zn (oxidation)
  - Cathode: Cu (reduction)
- **Ex 2:** Calculate equivalent mass of oxidising agent if change in oxidation state is 2 and molar mass is 80.

- Equivalent mass = 80 / 2 = 40

## Practice Set

- Q1: What is the role of salt bridge in an electrochemical cell?
  A. Prevents electron flow
  B. Maintains electrical neutrality
  C. Acts as anode
  D. Acts as cathode
- Q2: Which has the highest standard electrode potential?
  A. $Zn^{2+}/Zn$
  B. $Cu^{2+}/Cu$
  C. $H^{+}/H_2$
  D. $Li^{+}/Li$
- Q3: In Daniell cell, which species is reduced?
  A. $Zn^{2+}$
  B. $Cu^{2+}$
  C. Zn
  D. Cu

## Answer Key

- Q1: B - Salt bridge maintains electrical neutrality by allowing ion flow.
- Q2: B - $Cu^{2+}/Cu$ has higher EB0 than Zn and Li; $H^{+}/H_2$ is zero by convention.
- Q3: B - $Cu^{2+}$ gains electrons and is reduced to Cu metal.

## Memory Aids

- Redox couple = oxidised + reduced forms.
- Salt bridge = ion flow, no mixing.
- Electrode potential: tendency to gain/lose electrons.
- Standard hydrogen electrode EB0 = 0.00 V.
- Electrochemical series: negative EB0 = strong reducing agent.

# Predicting Reducing and Oxidising Power

## Key Concepts

- Lower electrode potential → stronger reducing agent.
- Order of reducing agents: Li > Mg > Al > Zn > Fe > Co > Ni > Sn > Pb > Cu > $I^-$ > $Fe^{2+}$ > Hg > Ag > $Br^-$ > $Cl^-$ > $F^-$
- Higher electrode potential → stronger oxidising agent.
- Order of oxidising agents: $F_2$ > $Cl_2$ > $Br_2$ > $I_2$
- Fluorine is the strongest oxidising agent.
- Lithium is the strongest reducing agent due to high hydration energy.
- Salt bridge electrolytes must have cations and anions with similar ionic mobilities.
- Equivalent mass of oxidising/reducing agent = Molar mass / Change in oxidation state per mole.

## Solved Examples

- **Ex 1:** Identify the strongest reducing agent among Li, Cu, and Ag.
  - Li has lowest electrode potential → strongest reducing agent.
- **Ex 2:** Arrange $F_2$, $Cl_2$, $Br_2$ in order of oxidising power.
  - $F_2$ > $Cl_2$ > $Br_2$

## Practice Set

- Q1: Which is the strongest oxidising agent?
  A. $Cl_2$
  B. $F_2$
  C. $Br_2$
  D. $I_2$
- Q2: Which element is the strongest reducing agent?
  A. Cu
  B. Li
  C. Ag
  D. Hg

- Q3: What is the equivalent mass of an oxidising agent with molar mass 96 and oxidation state change 3?
  A. 32
  B. 48
  C. 96
  D. 288

## Answer Key

- Q1: B - Fluorine ($F_2$) is the strongest oxidising agent.
- Q2: B - Lithium (Li) is the strongest reducing agent due to low electrode potential and high hydration energy.
- Q3: A - Equivalent mass = 96 / 3 = 32.

## Memory Aids

- Reducing power ↑ as electrode potential ↓.
- Oxidising power ↑ as electrode potential ↑.
- Fluorine = strongest oxidiser.
- Lithium = strongest reducer.
- Equivalent mass = molar mass / change in oxidation state.

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