Chapter 7: Redox Reactions

Comprehensive Study Notes

Class 11 Chemistry - NCERT Based

EXAM SPRINT - Complete Coverage for NEET and JEE Examinations

Introduction

Redox reactions represent one of the most important categories of chemical transformations. These reactions are fundamental to numerous processes including:

Everyday Applications:

- Combustion of fuels for energy
- Battery operations (dry and wet cells)
- Metal corrosion processes
- Industrial manufacturing (caustic soda production)

Advanced Applications:

- Electrochemical metal extraction
- Pharmaceutical synthesis
- Environmental processes (Hydrogen Economy, Ozone depletion)
- Biological systems (respiration, photosynthesis)

Key Principle: In redox reactions, oxidation and reduction occur simultaneously - "Where there is oxidation, there is always reduction."

7.1 CLASSICAL IDEA OF REDOX REACTIONS

Historical Evolution of Oxidation Concept

Stage 1: Oxygen-Based Definition

Original Definition: Oxidation = Addition of oxygen to an element/compound

Examples:

- $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$ *(Mg oxidized)*
- $S(s) + O_2(g) \rightarrow SO_2(g)$ *(S oxidized)*
- $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ *(CH_4 oxidized)*

Stage 2: Hydrogen-Based Extension

Expanded Definition: Oxidation = Removal of hydrogen from a compound

Example:

- $2H_2S(g) + O_2(g) \rightarrow 2S(s) + 2H_2O(l)$ *(H_2S oxidized by H removal)*

Stage 3: Electronegativity-Based Generalization

Modern Classical Definition:

- **Oxidation:** Addition of electronegative element OR removal of electropositive element
- **Reduction:** Addition of electropositive element OR removal of electronegative element

Examples:

- $Mg(s) + F_2(g) \rightarrow MgF_2(s)$ *(Mg oxidized - F added)*
- $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$ *(Mg oxidized - Cl added)*
- $2K_4[Fe(CN)_6] + H_2O_2 \rightarrow 2K_3[Fe(CN)_6] + 2KOH$ *(K removal = oxidation)*

Classical Reduction Examples

- $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$ *(O removal)*
- $2FeCl_3(aq) + H_2(g) \rightarrow 2FeCl_2(aq) + 2HCl(aq)$ *(Cl removal)*
- $CH_2{=}CH_2(g) + H_2(g) \rightarrow H_3C\text{-}CH_3(g)$ *(H addition)*

Key Insight: Oxidation and reduction always occur simultaneously, leading to the term "redox."

7.2 REDOX REACTIONS IN TERMS OF ELECTRON TRANSFER

Electronic Theory Foundation

Modern Definition:

- **Oxidation:** Loss of electrons by a species
- **Reduction:** Gain of electrons by a species
- **Oxidizing agent (Oxidant):** Electron acceptor
- **Reducing agent (Reductant):** Electron donor

Half-Reaction Approach

Example: Formation of NaCl

- **Overall:** $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
- **Oxidation half:** $2Na(s) \rightarrow 2Na^+(g) + 2e^-$
- **Reduction half:** $Cl_2(g) + 2e^- \rightarrow 2Cl^-(g)$

7.2.1 Competitive Electron Transfer Reactions

Metal Activity Series

Experimental Observations:

1. **$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$** (Spontaneous)
2. **$Cu + Zn^{2+} \rightarrow$ No reaction** (Non-spontaneous)

Activity Order: Zn > Cu > Ag (decreasing electron-releasing tendency)

Galvanic Cell Formation:

- Chemical energy → Electrical energy
- Basis for electrochemical series
- Foundation for battery technology

7.3 OXIDATION NUMBER

Definition and Purpose

Oxidation Number: A hypothetical charge assigned to an atom in a compound, assuming complete electron transfer to the more electronegative atom.

Purpose:

- Track electron movement in covalent compounds
- Systematic approach to identify redox changes
- Balancing complex redox equations

Rules for Assigning Oxidation Numbers

Fundamental Rules:

1. **Free elements:** Oxidation number = 0
   - Examples: H_2, O_2, Cl_2, P_4, S_8, Na, Mg

2. **Monatomic ions:** Oxidation number = Ionic charge
   - Na^+ = +1, Mg^{2+} = +2, Cl^- = -1, O^{2-} = -2
3. **Group trends:**
   - **Group 1:** Always +1 in compounds
   - **Group 2:** Always +2 in compounds
   - **Aluminum:** Always +3 in compounds
4. **Oxygen rules:**
   - **Most compounds:** -2
   - **Peroxides (H_2O_2, Na_2O_2):** -1
   - **Superoxides (KO_2, RbO_2):** -½
   - **With fluorine (OF_2, O_2F_2):** +2, +1 respectively
5. **Hydrogen rules:**
   - **Most compounds:** +1
   - **Metal hydrides (LiH, NaH, CaH_2):** -1
6. **Halogen rules:**
   - **Fluorine:** Always -1
   - **Other halogens:** Usually -1, positive in oxoacids/oxoanions
7. **Electroneutrality:**
   - **Neutral compounds:** Sum of oxidation numbers = 0
   - **Polyatomic ions:** Sum = Ionic charge

Oxidation Number Applications

Identifying Redox Changes:

- **Oxidation:** Increase in oxidation number

- **Reduction:** Decrease in oxidation number
- **Oxidizing agent:** Species undergoing reduction
- **Reducing agent:** Species undergoing oxidation

Stock Notation:

Roman numerals in parentheses indicate oxidation states:

- $AuCl$ → $Au(I)Cl$
- $AuCl_3$ → $Au(III)Cl_3$
- $SnCl_2$ → $Sn(II)Cl_2$
- $SnCl_4$ → $Sn(IV)Cl_4$

7.3.1 TYPES OF REDOX REACTIONS

1. Combination Reactions

Pattern: A + B → C (where A or B or both are elements)

Examples:

- $C(s) + O_2(g) \rightarrow CO_2(g)$ *(C: 0 → +4)*
- $3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$ *(Mg: 0 → +2, N: 0 → -3)*
- $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ *(C: -4 → +4)*

2. Decomposition Reactions

Pattern: Compound → Elements + Other compounds

Examples:

- $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$ *(H: +1 → 0, O: -2 → 0)*
- $2NaH(s) \rightarrow 2Na(s) + H_2(g)$ *(Na: +1 → 0, H: -1 → 0)*
- $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$ *(Cl: +5 → -1, O: -2 → 0)*

Note: Not all decomposition reactions are redox reactions.

- $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ *(Not redox - no oxidation state changes)*

3. Displacement Reactions

Metal Displacement:

Pattern: More active metal displaces less active metal from compound

Examples:

- $CuSO_4(aq) + Zn(s) \rightarrow Cu(s) + ZnSO_4(aq)$
- $V_2O_5(s) + 5Ca(s) \rightarrow 2V(s) + 5CaO(s)$
- $TiCl_4(l) + 2Mg(s) \rightarrow Ti(s) + 2MgCl_2(s)$

Hydrogen Displacement:

From water:

- $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
- $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

From steam:

- $Mg(s) + 2H_2O(g) \rightarrow Mg(OH)_2(s) + H_2(g)$
- $2Fe(s) + 3H_2O(g) \rightarrow Fe_2O_3(s) + 3H_2(g)$

From acids:

- $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
- $Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$

Non-metal Displacement:

Halogen activity series: $F_2 > Cl_2 > Br_2 > I_2$

Examples:

- $Cl_2(g) + 2KBr(aq) \rightarrow 2KCl(aq) + Br_2(l)$
- $Cl_2(g) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(s)$
- $Br_2(l) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(s)$

Industrial Application: Layer Test for halide identification

4. Disproportionation Reactions

Definition: Same element simultaneously oxidized and reduced

Requirements:

- Element in intermediate oxidation state
- Can exist in at least 3 different oxidation states
- Products have higher and lower oxidation states

Examples:

- $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$ *(O: -1 → -2, 0)*
- $3Cl_2(g) + 6OH^-(aq) \rightarrow 5Cl^-(aq) + ClO_3^-(aq) + 3H_2O(l)$ *(Cl: 0 → -1, +5)*
- $P_4(s) + 3OH^-(aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_2^-(aq)$ *(P: 0 → -3, +1)*

Exception: Fluorine cannot disproportionate (most electronegative, no positive oxidation state)

7.3.2 BALANCING REDOX REACTIONS

Method 1: Oxidation Number Method

Steps:

1. **Write correct formulas** for all reactants and products
2. **Assign oxidation numbers** to identify changing elements
3. **Calculate electron change** and equalize gain/loss
4. **Balance ionic charges** (add H^+ in acidic, OH^- in basic medium)
5. **Balance atoms** by adding H_2O molecules
6. **Verify balance** (atoms and charges)

Example: $Cr_2O_7^{2-} + SO_3^{2-} \rightarrow Cr^{3+} + SO_4^{2-}$ (acidic)

Step-by-step:

1. $Cr_2O_7^{2-}(aq) + SO_3^{2-}(aq) \rightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$
2. Cr: +6 → +3 (decrease of 3, × 2 = $6e^-$ gained) S: +4 → +6 (increase of 2, × 3 = $6e^-$ lost)
3. $Cr_2O_7^{2-} + 3SO_3^{2-} \rightarrow 2Cr^{3+} + 3SO_4^{2-}$
4. Add $8H^+$: $Cr_2O_7^{2-} + 3SO_3^{2-} + 8H^+ \rightarrow 2Cr^{3+} + 3SO_4^{2-}$
5. Add $4H_2O$: $Cr_2O_7^{2-} + 3SO_3^{2-} + 8H^+ \rightarrow 2Cr^{3+} + 3SO_4^{2-} + 4H_2O$

Method 2: Half-Reaction Method

Steps:

1. **Write unbalanced ionic equation**
2. **Separate into half-reactions**

3. **Balance atoms** (other than O and H)
4. **Add H_2O** to balance O atoms
5. **Add H^+** to balance H atoms (acidic medium)
6. **Add electrons** to balance charges
7. **Equalize electrons** in both half-reactions
8. **Add half-reactions** and simplify
9. **For basic medium:** Add OH^- to neutralize H^+

Example: $MnO_4^- + I^- \rightarrow MnO_2 + I_2$ (basic)

Oxidation half: $2I^- \rightarrow I_2 + 2e^-$ **Reduction half:** $MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$

For basic medium: Add $4OH^-$ to both sides of reduction half: $MnO_4^- + 4H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$

Balanced equation: $6I^- + 2MnO_4^- + 4H_2O \rightarrow 3I_2 + 2MnO_2 + 8OH^-$

7.3.3 REDOX TITRATIONS

Self-Indicator Systems

Example: Permanganate Titrations

- MnO_4^- (purple) → Mn^{2+} (colorless)
- End point: First permanent pink color
- Sensitivity: 10^{-6} mol/L detection limit

External Indicator Systems

Example: Dichromate Titrations

- Uses diphenylamine indicator
- Color change: Colorless → Intense blue
- Indicates complete oxidation of analyte

Iodometric Methods

Principle: Indirect determination using I_2/I^- couple

Steps:

1. **Liberation:** $Cu^{2+} + 4I^- \rightarrow Cu_2I_2 + I_2$
2. **Detection:** I_2 + starch → Blue complex
3. **Titration:** $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$
4. **End point:** Blue color disappears

7.4 REDOX REACTIONS AND ELECTRODE PROCESSES

Electrochemical Cells

Redox Couples

Definition: Oxidized and reduced forms of same species **Notation:** Oxidized form/Reduced form **Examples:** Cu^{2+}/Cu, Zn^{2+}/Zn, Fe^{3+}/Fe^{2+}

Daniell Cell

Construction:

- **Anode:** $Zn|Zn^{2+}$ (oxidation occurs)
- **Cathode:** $Cu^{2+}|Cu$ (reduction occurs)
- **Salt bridge:** Maintains electrical neutrality

- **External circuit:** Electron flow (Zn → Cu)

Cell Reactions:

- **Anode:** $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
- **Cathode:** $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
- **Overall:** $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

Standard Electrode Potentials

Definition and Convention

- **Standard conditions:** 298K, 1 atm, 1M concentration
- **Reference:** Standard Hydrogen Electrode (SHE) = 0.00 V
- **Sign convention:**
  - Negative E° = Stronger reducing agent than H^+/H_2
  - Positive E° = Weaker reducing agent than H^+/H_2

Electrochemical Series (Selected Values)

Half-Reaction	E° (V)
$Li^+ + e^- \rightarrow Li$	-3.05
$K^+ + e^- \rightarrow K$	-2.93
$Na^+ + e^- \rightarrow Na$	-2.71
$Mg^{2+} + 2e^- \rightarrow Mg$	-2.36
$Al^{3+} + 3e^- \rightarrow Al$	-1.66
$Zn^{2+} + 2e^- \rightarrow Zn$	-0.76
$Fe^{2+} + 2e^- \rightarrow Fe$	-0.44
$2H^+ + 2e^- \rightarrow H_2$	0.00
$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34
$I_2 + 2e^- \rightarrow 2I^-$	+0.54
$Ag^+ + e^- \rightarrow Ag$	+0.80
$Br_2 + 2e^- \rightarrow 2Br^-$	+1.09
$Cl_2 + 2e^- \rightarrow 2Cl^-$	+1.36
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	+1.51
$F_2 + 2e^- \rightarrow 2F^-$	+2.87

Applications:

- **Predicting spontaneous reactions**
- **Designing galvanic cells**
- **Understanding corrosion**
- **Metal extraction processes**

NEET/JEE Important Points

High-Yield Topics:

1. Oxidation Number Rules:

- Exception cases (peroxides, superoxides, metal hydrides)
- Stock notation for transition metals
- Fractional oxidation numbers (average states)

2. Balancing Equations:

- Both algebraic and half-reaction methods
- Acidic vs basic medium differences
- Common redox couples

3. Types of Reactions:

- Identification from equations
- Disproportionation conditions
- Industrial applications

4. Electrochemical Series:

- Predicting reaction feasibility
- Metal displacement reactions
- Relative reducing/oxidizing power

Common Exam Patterns:

1. Numerical Problems:

- Oxidation number calculations
- Equivalent weight determinations
- Titration calculations

2. Mechanism Questions:

- Electron transfer pathways
- Half-reaction writing
- Cell representations

3. Application Problems:

- Industrial processes
- Biological systems
- Environmental chemistry

Memory Aids and Mnemonics

Oxidation Number Rules:

"Free Elements Get Zero, Ions Get Their Charge"

- Free elements = 0
- Monatomic ions = charge

Redox Definitions:

"OIL RIG LEO GER"

- **O**xidation **I**s **L**oss (of electrons)

- **R**eduction **I**s **G**ain (of electrons)
- **L**oss of **E**lectrons **O**xidation
- **G**ain of **E**lectrons **R**eduction

Electrochemical Series:

"Please Stop Calling Me A Careless Zebra Instead Try Learning How Copper Shows Bravery In Acidic Conditions For Gold"

- **P**b, **S**n, **C**d, **M**g, **A**l, **C**r, **Z**n, **I**ron, **T**in, **L**ead, **H_2**, **C**u, **S**ilver, **B**romine, **I**odine, **A**u, **Cl_2**, **F_2**, **G**old

Balancing Steps:

"Write Assign Calculate Balance Verify"

- **W**rite skeleton equation
- **A**ssign oxidation numbers
- **C**alculate electron changes
- **B**alance charges and atoms
- **V**erify final equation

Practice Questions for NEET/JEE

Multiple Choice Questions:

1. In the reaction: $2FeCl_3 + H_2S \rightarrow 2FeCl_2 + 2HCl + S$, which species is oxidized? a) Fe^{3+} b) Cl^- c) H_2S d) S
2. The oxidation number of Cr in $K_2Cr_2O_7$ is: a) +3 b) +6 c) +7 d) +2
3. Which of the following is a disproportionation reaction? a) $2H_2O_2 \rightarrow 2H_2O + O_2$ b) $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$ c) $2KI + Cl_2 \rightarrow 2KCl + I_2$ d) $CaCO_3 \rightarrow CaO + CO_2$

Assertion-Reason Questions:

Assertion (A): In the reaction $Cl_2 + 2KI \rightarrow 2KCl + I_2$, Cl_2 is reduced. **Reason (R):** The oxidation number of Cl decreases from 0 to -1.

Short Answer Questions:

1. Balance the equation: $MnO_4^- + C_2O_4^{2-} \rightarrow Mn^{2+} + CO_2$ (acidic medium)
2. Calculate the oxidation number of S in $Na_2S_4O_6$.
3. Identify the oxidizing and reducing agents in: $Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$

Long Answer Questions:

1. Explain the electronic concept of redox reactions with examples.
2. Derive the relationship between oxidation number change and electrons transferred.
3. Describe the construction and working of Daniell cell.

Summary Table: Key Comparisons

Classical vs Electronic Concepts:

Aspect	Classical	Electronic
Oxidation	Addition of O_2/electronegative element	Loss of electrons
Reduction	Addition of H_2/electropositive element	Gain of electrons
Scope	Limited to specific elements	Universal applicability
Mechanism	Atom/ion transfer	Electron transfer

Balancing Methods Comparison:

Method	Best For	Limitations
Oxidation Number	Simple reactions, beginners	Complex polyatomic ions
Half-Reaction	Complex reactions, electrochemistry	Requires ion knowledge

Types of Redox Reactions:

Type	Characteristics	Examples
Combination	A + B → C	Metal + non-metal
Decomposition	AB → A + B	Electrolysis reactions
Displacement	A + BC → AC + B	Metal activity series
Disproportionation	Same element oxidized and reduced	$ClO^- \rightarrow Cl^- + ClO_3^-$

EXAM SPRINT - Master Redox Reactions through systematic study of electron transfer concepts, oxidation number rules, balancing techniques, and electrochemical applications. Regular practice of numerical problems and mechanism questions is essential for NEET/JEE success.

Source: NCERT Chemistry Class 11, Chapter 7 - Comprehensive coverage for competitive exam preparation