REDOX REACTIONS

Class 11 Chemistry – Chapter 7

CBSE Board 2025-26 Syllabus

Complete Notes | Tips & Tricks | Case Studies | Expected Questions

Prepared by: Math Love Institute, Indore / Raipur

★ ★ ★ COMPREHENSIVE STUDY MATERIAL FOR CLASS 11 CHEMISTRY ★ ★ ★

CHAPTER OVERVIEW

What You Will Learn:

- **Classical Idea of Redox:** Oxidation as addition of oxygen / removal of hydrogen; Reduction as opposite
- **Electronic Concept:** Oxidation = loss of electrons; Reduction = gain of electrons
- **Oxidation Number:** Rules for calculation and significance
- **Types of Redox Reactions:** Combination, Decomposition, Displacement, Disproportionation
- **Balancing Redox Equations:** Oxidation Number Method and Half-Reaction (Ion-Electron) Method
- **Redox Titrations:** Self-indicators, external indicators
- **Electrode Processes:** Galvanic cells, Standard Electrode Potential, Electrochemical Series

💡 Why This Chapter is Important:

Redox Reactions is a **high-weightage chapter** for both board exams and competitive tests. It directly connects to:

- ✅ CBSE Board Exams (5–8 marks typically)
- ✅ Electrochemistry (Class 12 – direct continuation!)
- ✅ JEE Main & Advanced (1–2 questions every year)
- ✅ NEET (Inorganic Chemistry backbone)
- ✅ Industrial processes: batteries, metallurgy, bleaching, corrosion

Key Focus Areas: Oxidation number rules, balancing redox equations, identifying oxidant/reductant, types of redox reactions

📜 Chapter Quote & Context:

"Where there is oxidation, there is always reduction — Chemistry is essentially a study of redox systems."

Redox reactions are all around us — burning of fuel, rusting of iron, operation of batteries, photosynthesis, respiration, extraction of metals, and even the development of the ozone hole all involve redox processes.

🔬 SECTION 7.1 – CLASSICAL IDEA OF REDOX REACTIONS

Oxidation (Classical Definition)

Oxidation = Addition of oxygen / electronegative element

OR

Removal of hydrogen / electropositive element

📌 Examples of Oxidation:

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$ [Addition of oxygen to Mg]

$2H_2S(g) + O_2(g) \rightarrow 2S(s) + 2H_2O(l)$ [Removal of hydrogen from H_2S]

$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$ [Addition of electronegative element Cl]

Reduction (Classical Definition)

Reduction = Removal of oxygen / electronegative element

OR

Addition of hydrogen / electropositive element

📌 Examples of Reduction:

$2HgO(s) \rightarrow 2Hg(l) + O_2(g)$ [Removal of oxygen]

$2FeCl_3(aq) + H_2(g) \rightarrow 2FeCl_2(aq) + 2HCl(aq)$ [Removal of electronegative Cl]

$CH_2{=}CH_2(g) + H_2(g) \rightarrow CH_3{-}CH_3(g)$ [Addition of hydrogen]

🎯 Key Insight — "REDOX" = Both Always Together!

Oxidation and reduction **always occur simultaneously**. You can never have one without the other. That's why the combined term **"REDOX"** was coined.

Memory Trick: "OIL RIG"

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

⚡ SECTION 7.2 – REDOX IN TERMS OF ELECTRON TRANSFER

Oxidation: Loss of electron(s) by any species → Species is *oxidised*

Reduction: Gain of electron(s) by any species → Species is *reduced*

Oxidising Agent (Oxidant): Acceptor of electron(s)

Reducing Agent (Reductant): Donor of electron(s)

📌 Classic Example – Formation of NaCl:

Oxidation half: $2Na(s) \rightarrow 2Na^{+}(g) + 2e^{-}$

Reduction half: $Cl_2(g) + 2e^{-} \rightarrow 2Cl^{-}(g)$

Overall: $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

→ Na is the **reducing agent** (loses electrons); Cl_2 is the **oxidising agent** (gains electrons)

7.2.1 Competitive Electron Transfer Reactions

Zinc displacing Copper:

$$\mathrm{Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)}$$

Zn is oxidised (loses $2e^-$); Cu^{2+} is reduced (gains $2e^-$). Blue colour of $CuSO_4$ solution fades as Cu^{2+} is consumed.

Copper displacing Silver:

$$\mathrm{Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)}$$

Solution turns blue (Cu^{2+} formed); silver deposits on copper strip.

🎯 Electron Releasing Tendency Order:

From the above reactions, electron releasing (reducing) tendency: **Zn > Cu > Ag**

This leads to the concept of the **Electrochemical Series / Activity Series**.

🔢 SECTION 7.3 – OXIDATION NUMBER

Oxidation Number = The oxidation state of an element in a compound, assuming complete transfer of electrons to the more electronegative atom.

Rules for Calculating Oxidation Number

📋 The 6 Golden Rules (MUST MEMORISE):

1. **Free elements:** Oxidation number = **0** (e.g., Na, H_2, O_2, S_8 → all zero)
2. **Monoatomic ions:** Oxidation number = charge on ion (e.g., Na^+ = +1, Fe^{3+} = +3, Cl^- = −1)
3. **Oxygen:** Usually **−2**. Exceptions:
    - Peroxides (H_2O_2, Na_2O_2): O = **−1**
    - Superoxides (KO_2, RbO_2): O = **−½**
    - Fluoride compounds (OF_2): O = **+2**; (O_2F_2): O = **+1**
4. **Hydrogen:** Usually **+1**. Exception: Metal hydrides (LiH, NaH, CaH_2) → H = **−1**
5. **Fluorine:** Always **−1**. Other halogens: −1 in halides, positive in oxoacids/oxoanions
6. **Sum Rule:** Algebraic sum of all oxidation numbers = 0 (neutral molecule) or = charge (polyatomic ion)

📝 Worked Examples – Finding Oxidation Numbers:

Example 1: K_2MnO_4

K = +1 (×2), O = −2 (×4). Let Mn = x

$2(+1) + x + 4(-2) = 0 \rightarrow x = +6$. *Mn is +6*

Example 2: $Cr_2O_7^{2-}$

O = −2 (×7). Let Cr = x

$2x + 7(-2) = -2 \rightarrow 2x = +12 \rightarrow x = +6$. *Cr is +6*

Example 3: MnO_4^-

O = −2 (×4). Let Mn = x

$x + 4(-2) = -1 \rightarrow x = +7$. *Mn is +7*

Example 4: NH_4^+

H = +1 (×4). Let N = x

$x + 4(+1) = +1 \rightarrow x = -3$. *N is −3*

Example 5: H_2O_2 (Peroxide)

H = +1 (×2). Let O = x

$2(+1) + 2x = 0 \rightarrow x = -1$. *O is −1 (NOT −2 — peroxide!)*

⚠️ Common Mistakes to Avoid:

- ❌ Assigning O = −2 in H_2O_2 (it is a PEROXIDE, so O = −1)
- ❌ Assigning H = +1 in NaH (metal hydride, so H = −1)
- ❌ Forgetting to account for the charge of the ion when applying Sum Rule
- ❌ Confusing oxidation number with valency
- ❌ In $Na_2S_2O_3$, the two S atoms have different oxidation states; average is calculated

Stock Notation

🎯 Stock Notation (by Alfred Stock – German Chemist):

The oxidation number is written as a Roman numeral in parentheses after the metal symbol.

Compound Name	Stock Notation	Oxidation State of Metal
Aurous chloride ($AuCl$)	$Au(I)Cl$	+1
Auric chloride ($AuCl_3$)	$Au(III)Cl_3$	+3
Stannous chloride ($SnCl_2$)	$Sn(II)Cl_2$	+2
Stannic chloride ($SnCl_4$)	$Sn(IV)Cl_4$	+4
Ferrous oxide (FeO)	$Fe(II)O$	+2
Ferric oxide (Fe_2O_3)	$Fe(III)_2O_3$ or $Fe_2(III)O_3$	+3
Manganese dioxide (MnO_2)	$Mn(IV)O_2$	+4

Oxidation Number Based Definitions

Oxidation = Increase in oxidation number

Reduction = Decrease in oxidation number

Oxidising Agent = Species that increases oxidation number of other species (itself gets reduced)

Reducing Agent = Species that decreases oxidation number of other species (itself gets oxidised)

Redox Reaction = Any reaction involving change in oxidation number

SECTION 7.3.1 – TYPES OF REDOX REACTIONS

Type 1: Combination (Synthesis) Reactions

General Form: A + B → C

At least one reactant must be in elemental form for it to be a redox reaction.

$C(s) + O_2(g) \rightarrow CO_2(g)$ [C: 0 → +4; O: 0 → −2]

$3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$ [Mg: 0 → +2; N: 0 → −3]

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ [C: −4 → +4; O: 0 → −2]

Note: All combustion reactions are combination redox reactions!

Type 2: Decomposition Reactions

General Form: C → A + B (at least one product in elemental state)

Note: Not all decomposition reactions are redox reactions (e.g., $CaCO_3 \rightarrow CaO + CO_2$ is NOT redox).

$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$ [H: +1 → 0; O: −2 → 0]

$2NaH(s) \rightarrow 2Na(s) + H_2(g)$ [H: −1 → 0; Na: +1 → 0]

$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$ [Cl: +5 → −1; O: −2 → 0]

Type 3: Displacement Reactions

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

(a) Metal Displacement:

$Zn(s) + CuSO_4(aq) \rightarrow Cu(s) + ZnSO_4(aq)$ [Zn: 0 → +2; Cu: +2 → 0]

$TiCl_4(l) + 2Mg(s) \rightarrow Ti(s) + 2MgCl_2(s)$ [Industrial use: Ti extraction]

$Cr_2O_3(s) + 2Al(s) \rightarrow Al_2O_3(s) + 2Cr(s)$ [Goldschmidt thermite process]

(b) Non-Metal Displacement – Hydrogen Displacement:

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$ [Alkali metals displace H from cold water]

$Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(s) + H_2(g)$ [Mg reacts with steam]

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ [Lab preparation of H_2]

(c) Non-Metal Displacement – Halogen Displacement:

$Cl_2(g) + 2KBr(aq) \rightarrow 2KCl(aq) + Br_2(l)$ [Cl displaces Br]

$Cl_2(g) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(s)$ [Layer Test – basis]

$Br_2(l) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(s)$ [Br displaces I]

Oxidising power of halogens: $F_2 > Cl_2 > Br_2 > I_2$

Reducing power of hydrohalic acids: HI > HBr > HCl > HF

Type 4: Disproportionation Reactions

Disproportionation = An element in one oxidation state simultaneously gets both oxidised AND reduced in the same reaction.

The element must be capable of existing in at least **3 oxidation states**.

📌 Key Examples:

$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$ [O: −1 → −2 (reduced) and −1 → 0 (oxidised)]

$Cl_2(g) + 2OH^-(aq) \rightarrow ClO^-(aq) + Cl^-(aq) + H_2O(l)$ [Cl: 0 → +1 and 0 → −1]

Household bleach formation: ClO^- (hypochlorite) formed in above reaction oxidises colour-bearing stains.

$P_4(s) + 3OH^-(aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_2^-(aq)$ [P: 0 → −3 and 0 → +1]

🎯 How to Identify Disproportionation:

- The SAME element appears on BOTH sides of the equation with DIFFERENT oxidation numbers
- One oxidation number is higher and one is lower than the original
- The original element is at an INTERMEDIATE oxidation state
- **ClO_4^- does NOT disproportionate** — Cl is already at its highest oxidation state (+7)
- **F_2 does NOT disproportionate** — F cannot exhibit positive oxidation state (most electronegative)

⚠️ Comproportionation – Opposite of Disproportionation!

When two different oxidation states of the same element combine to give a single intermediate oxidation state, it is called **Comproportionation** (or Synproportionation).

Example: $Cu^0 + Cu^{2+} \rightarrow 2Cu^+$ (reverse of disproportionation)

⚖️ SECTION 7.3.2 – BALANCING REDOX REACTIONS

Two Methods Available:

- **Method 1:** Oxidation Number Method
- **Method 2:** Half-Reaction (Ion-Electron) Method

Method 1: Oxidation Number Method

📋 Steps (5 Steps):

1. **Step 1:** Write correct formulas for all reactants and products.
2. **Step 2:** Assign oxidation numbers to ALL atoms; identify atoms whose oxidation number changes.
3. **Step 3:** Calculate increase and decrease in oxidation number. Multiply by suitable coefficients to make them equal.
4. **Step 4:** Balance the ionic charges by adding H^+ (in acidic medium) or OH^- (in basic medium).
5. **Step 5:** Balance H atoms by adding H_2O molecules. Then verify O atom balance.

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

Step 2: Cr: +6 → +3 (decrease of 3 per Cr, total 6 for 2 Cr)
S: +4 → +6 (increase of 2 per S)

Step 3: To equalise: multiply Cr term by 1 (×2 Cr already) and S term by 3:
→ $Cr_2O_7^{2-} + 3SO_3^{2-} \rightarrow 2Cr^{3+} + 3SO_4^{2-}$

Step 4: Add $8H^+$ on left to balance charges (acidic medium)

Step 5: Add $4H_2O$ on right to balance H atoms

Final: $Cr_2O_7^{2-}(aq) + 3SO_3^{2-}(aq) + 8H^+(aq) \rightarrow 2Cr^{3+}(aq) + 3SO_4^{2-}(aq) + 4H_2O(l)$

Method 2: Half-Reaction (Ion-Electron) Method

📋 Steps (7 Steps):

1. **Step 1:** Write the unbalanced ionic equation.
2. **Step 2:** Separate into two half-reactions (oxidation and reduction).
3. **Step 3:** Balance all atoms other than O and H in each half-reaction.
4. **Step 4:** Balance O atoms using H_2O; balance H atoms using H^+ (acidic) or OH^- (basic).
5. **Step 5:** Balance charges by adding electrons. Equalize electrons in both half-reactions by multiplying.
6. **Step 6:** Add the two half-reactions; cancel electrons and common terms.
7. **Step 7:** Verify atom count and charge balance on both sides.

For basic medium: First balance as in acidic medium, then add OH^- to both sides for each H^+, combine H^+ and OH^- to form H_2O.

📌 Worked Example: $Fe^{2+} + Cr_2O_7^{2-} \rightarrow Fe^{3+} + Cr^{3+}$ (acidic medium)

Oxidation half: $Fe^{2+} \rightarrow Fe^{3+} + e^-$ (×6)

$\rightarrow 6Fe^{2+} \rightarrow 6Fe^{3+} + 6e^-$

Reduction half: $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

Add both:

Final: $6Fe^{2+}(aq) + Cr_2O_7^{2-}(aq) + 14H^+(aq) \rightarrow 6Fe^{3+}(aq) + 2Cr^{3+}(aq) + 7H_2O(l)$

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

Oxidation half: $2I^- \rightarrow I_2 + 2e^-$ (×3) $\rightarrow 6I^- \rightarrow 3I_2 + 6e^-$

Reduction half: $MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$ (×2) $\rightarrow 2MnO_4^- + 4H_2O + 6e^- \rightarrow 2MnO_2 + 8OH^-$

Final: $6I^-(aq) + 2MnO_4^-(aq) + 4H_2O(l) \rightarrow 3I_2(s) + 2MnO_2(s) + 8OH^-(aq)$

SECTION 7.3.3 – REDOX TITRATIONS

Types of Indicators in Redox Titrations:

1. **Self-Indicators:** The reagent itself is coloured.
    - Example: **$KMnO_4$** (permanganate, purple/violet colour) acts as self-indicator
    - End point: First lasting pink tinge appears (all reductant consumed)
    - Detection limit: As low as 10^{-6} mol L^{-1}
2. **External Indicators:** Added separate indicator
    - Example: **$K_2Cr_2O_7$** uses *diphenylamine* indicator → intense blue colour at end point
3. **Iodometric Titrations (Starch Indicator Method):**
    - Iodine gives intense blue colour with starch
    - $2Cu^{2+}(aq) + 4I^{-}(aq) \rightarrow Cu_2I_2(s) + I_2(aq)$ — I_2 liberated
    - $I_2 + 2S_2O_3^{2-} \rightarrow 2I^{-} + S_4O_6^{2-}$ — I_2 consumed, blue colour disappears at end point

🔋 SECTION 7.4 – REDOX REACTIONS AND ELECTRODE PROCESSES

The Daniell Cell

Construction and Working:

- One beaker: $CuSO_4$ solution + Cu rod (cathode / reduction)
- Other beaker: $ZnSO_4$ solution + Zn rod (anode / oxidation)
- Salt bridge: U-tube with KCl or NH_4NO_3 in agar-agar — connects the two solutions electrically without mixing
- Metallic wire with ammeter connects the two rods

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

Cathode (reduction): $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Overall: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

🎯 Key Terminology – Electrode Processes:

Term	Definition
Anode	Electrode where oxidation occurs; negatively charged in galvanic cell; electrons flow OUT from anode
Cathode	Electrode where reduction occurs; positively charged in galvanic cell; electrons flow INTO cathode
Electrode Potential	Potential associated with each electrode; depends on the species and its concentration
Standard Electrode Potential (E°)	Electrode potential when all species are at unit concentration (1 M), gas at 1 atm, at 298 K
Redox Couple	Oxidised and reduced forms of a species written together (e.g., Zn^{2+}/Zn; Cu^{2+}/Cu)
Salt Bridge	Provides electrical contact between two solutions; maintains electrical neutrality; prevents mixing

Standard Electrode Potentials (E° Values)

Reference Electrode: **Standard Hydrogen Electrode (SHE) — E° = 0.00 V**

Negative E° → **Stronger reducing agent than H^+/H_2 couple**

Positive E° → **Weaker reducing agent than H^+/H_2 couple**

Standard Electrode Potentials at 298 K (Selected Values):

Reaction (Reduction)	E° (V)	Category
$F_2 + 2e^- \rightarrow 2F^-$	+2.87	Strongest oxidant
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	+1.51	Strong oxidant
$Cl_2 + 2e^- \rightarrow 2Cl^-$	+1.36	Strong oxidant
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	+1.33	Strong oxidant
$Br_2 + 2e^- \rightarrow 2Br^-$	+1.09	Moderate oxidant
$Ag^+ + e^- \rightarrow Ag$	+0.80	Moderate oxidant
$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34	Weak oxidant
$2H^+ + 2e^- \rightarrow H_2$	**0.00**	**Reference**
$Ni^{2+} + 2e^- \rightarrow Ni$	−0.25	Weak reductant
$Fe^{2+} + 2e^- \rightarrow Fe$	−0.44	Moderate reductant
$Zn^{2+} + 2e^- \rightarrow Zn$	−0.76	Good reductant
$Al^{3+} + 3e^- \rightarrow Al$	−1.66	Strong reductant
$Na^+ + e^- \rightarrow Na$	−2.71	Very strong reductant
$Li^+ + e^- \rightarrow Li$	−3.05	Strongest reductant

Top to bottom in the table: Increasing strength of reducing agent ↑ ; Increasing strength of oxidising agent ↓

🎯 Using E° to Predict Feasibility:

A redox reaction is feasible if **E°cell = E°cathode − E°anode > 0**

The species with higher E° value acts as the oxidising agent (gets reduced).

Example: Is $Fe^{3+}(aq) + I^{-}(aq)$ feasible?

E° (Fe^{3+}/Fe^{2+}) = +0.77 V ; E° (I_2/I^{-}) = +0.54 V

E°cell = +0.77 − 0.54 = +0.23 V > 0 → **Feasible** ✅

Fe^{3+} is reduced, I^{-} is oxidised.

🧠 SPECIAL TOPICS

Fractional Oxidation Numbers – The Paradox!

When Oxidation Numbers Appear Fractional:

Sometimes, the calculated oxidation number comes out as a fraction. This happens when the same element exists in **more than one actual oxidation state** in the same molecule, and we are calculating the **average**.

Species	Average O.N.	Reality (by Structure)
C_3O_2 (carbon suboxide)	Carbon = +4/3	Two C at +2; one C at 0
Br_3O_8 (tribromooctaoxide)	Bromine = +16/3	Two Br at +6; one Br at +4
$S_4O_6^{2-}$ (tetrathionate ion)	Sulphur = +2.5	Two S at +5; two S at 0
Fe_3O_4 (mixed oxide)	Iron = +8/3	One Fe^{2+}; two Fe^{3+}

Conclusion: Fractional oxidation numbers represent the *average* of different whole-number oxidation states. The reality is always revealed by the structure.

Limitations of Oxidation Number Concept

Important Limitation:

The oxidation number concept assumes **complete transfer of electrons**, which is only an approximation. In reality (especially for covalent compounds), there is only a **partial shift in electron density**.

Modern definition:

- **Oxidation** = Decrease in electron density around the atom
- **Reduction** = Increase in electron density around the atom

📝 CASE-BASED QUESTIONS (CBSE PATTERN)

Case Study 1: The Daniell Cell and Electrochemistry

A student sets up a Daniell cell by placing a zinc strip in zinc sulphate solution and a copper strip in copper sulphate solution. The two solutions are connected by a salt bridge containing KCl solution in agar-agar. The metal strips are connected by a metallic wire with a switch and ammeter.

1. What happens at the zinc electrode when the switch is on? Write the half-reaction.
2. Why does the blue colour of $CuSO_4$ solution decrease over time?
3. What is the role of the salt bridge? Why is KCl preferred?
4. Define standard electrode potential. What is the reference electrode?
5. The E° values for $Zn^{2+}/Zn = -0.76$ V and $Cu^{2+}/Cu = +0.34$ V. Calculate the E°cell.

Answers:

1. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ (oxidation takes place at the anode — zinc rod)

2. Cu^{2+} ions are being reduced to Cu (deposited on copper strip), consuming the Cu^{2+} ions in solution.

3. Salt bridge maintains electrical neutrality in both half-cells by allowing ion migration (K^+ moves to cathode compartment, Cl^- moves to anode compartment). KCl is preferred as K^+ and Cl^- have similar mobility.

4. Standard electrode potential is the electrode potential measured at 298K with all species at unit concentration (1 M), gases at 1 atm. Reference: Standard Hydrogen Electrode (SHE), E° = 0.00 V.

5. $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = +0.34 - (-0.76) = +1.10$ V

Case Study 2: Redox Reactions in Everyday Life

The bleaching action of household bleach is based on the disproportionation of chlorine in alkaline medium. The hyochlorite ion formed oxidises the colour-bearing stains. Permanganate is used as a powerful oxidant in organic chemistry labs. Rusting of iron is an electrochemical redox process.

1. Write the equation for the disproportionation of Cl_2 in NaOH solution. Identify oxidation and reduction.
2. What is the oxidation state of Mn in $KMnO_4$? Why is it a strong oxidant?
3. Write the equation for rusting of iron (formation of Fe_2O_3). Is it a redox reaction?
4. Why does F_2 NOT show disproportionation in alkaline medium?
5. Give one example each of: combination redox, decomposition redox, displacement redox.

Answers:

1. $Cl_2(g) + 2OH^-(aq) \rightarrow ClO^-(aq) + Cl^-(aq) + H_2O(l)$. Cl(0)→Cl(+1): oxidation; Cl(0)→Cl(−1): reduction.

2. Mn is +7 in $KMnO_4$. It is a strong oxidant because Mn at +7 has a high tendency to gain electrons and be reduced to lower oxidation states.

3. $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$. Yes, it is redox: Fe 0→+3 (oxidised), O 0→−2 (reduced).

4. F is the most electronegative element and cannot exhibit positive oxidation states. So Cl cannot undergo disproportionation to give ClO^- equivalent because F cannot exist in +1 state.

5. Combination: $C + O_2 \rightarrow CO_2$; Decomposition: $2KClO_3 \rightarrow 2KCl + 3O_2$; Displacement: $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

Case Study 3: Balancing Redox Equations

A chemistry teacher asks students to balance the following reactions using the ion-electron method:

(A) $MnO_4^- + SO_2 \rightarrow Mn^{2+} + HSO_4^-$ (acidic medium)

(B) $P_4 + OH^- \rightarrow PH_3 + H_2PO_2^-$ (basic medium)

1. For reaction (A), identify the oxidant and reductant.
2. For reaction (A), write the two half-reactions with electron balance.
3. Write the fully balanced equation for reaction (A).
4. For reaction (B), what type of redox reaction is this?
5. Explain why the two half-reactions must have equal electrons before being added.

Answers:

1. Mn in MnO_4^- goes from +7 to +2 (reduction) → MnO_4^- is the oxidant. S in SO_2 goes from +4 to +6 in HSO_4^- (oxidation) → SO_2 is the reductant.

2. Reduction: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$; Oxidation: $SO_2 + 2H_2O \rightarrow HSO_4^- + 3H^+ + 2e^-$

3. $2MnO_4^-(aq) + 5SO_2(g) + 2H_2O(l) \rightarrow 2Mn^{2+}(aq) + 5HSO_4^-(aq) + H^+(aq)$

4. Disproportionation — P(0) in P_4 simultaneously becomes P(−3) in PH_3 and P(+1) in $H_2PO_2^-$.

5. Electrons are neither created nor destroyed. The number of electrons lost in oxidation must equal the number of electrons gained in reduction, so the electrons cancel completely when half-reactions are added.

❓ MOST EXPECTED EXAM QUESTIONS

1-Mark Questions:

- What is the oxidation number of Cr in $Cr_2O_7^{2-}$? *(+6)*
- What is the oxidation number of S in $Na_2S_2O_3$? *(+2)*
- What is the oxidation number of O in H_2O_2? *(−1)*
- Define disproportionation reaction.
- What is the standard electrode potential of SHE?
- In electrolytic cells, which electrode is the anode?
- Name the indicator used in $KMnO_4$ titrations.
- Which halogen is the best oxidising agent? *(F_2)*
- What is the oxidation number of Fe in Fe_3O_4? *(+8/3 average; or +2 for 1 Fe and +3 for 2 Fe)*
- Define a redox couple. Give example.

2-Mark Questions:

- Identify the oxidising agent and reducing agent in: $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
- Why does HF not act as a reducing agent?
- Write the two half-reactions for: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$
- Classify the reaction: $2KClO_3 \rightarrow 2KCl + 3O_2$ and assign oxidation numbers.
- State the role of a salt bridge in a Daniell cell.
- Explain why F_2 does not show disproportionation.
- Distinguish between oxidation number and oxidation state.

3-Mark Questions:

- Balance by oxidation number method: $Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+}$ (acidic)
- Balance by half-reaction method: $MnO_4^- + I^- \rightarrow MnO_2 + I_2$ (basic)
- Justify that $2Cu_2O + Cu_2S \rightarrow 6Cu + SO_2$ is a redox reaction. Identify oxidant and reductant.
- Using E° values, predict which of the following reactions will occur: (a) $Fe^{3+} + I^-$, (b) $Cu + Ag^+$, (c) $Ag + Fe^{3+}$
- Explain competitive electron transfer reactions with the example of Zn and Cu^{2+}.
- What is disproportionation? Give two examples with equations.

5-Mark Questions (Long Answer):

- Describe the Daniell cell with labelled diagram. Write electrode reactions, define redox couple, and calculate E°cell.
- State and apply the rules for calculating oxidation number. Find O.N. of: (a) S in H_2SO_5, (b) Cr in $Cr_2O_7^{2-}$, (c) N in NO_3^-
- Classify the following as combination, decomposition, displacement, or disproportionation redox reactions with reasoning: (a) $N_2 + O_2 \rightarrow 2NO$, (b) $2Pb(NO_3)_2 \rightarrow 2PbO + 4NO_2 + O_2$, (c) $NaH + H_2O \rightarrow NaOH + H_2$, (d) $2NO_2 + 2OH^- \rightarrow NO_2^- + NO_3^- + H_2O$
- Balance by ion-electron method: $H_2O_2 + Fe^{2+} \rightarrow Fe^{3+} + H_2O$ (acidic medium)

📖 EXAM PREPARATION STRATEGY

🗓️ Topic-wise Priority for Exams:

1. **Oxidation Number Calculations (HIGHEST PRIORITY)**
    - Practice 20+ compounds including peroxides, metal hydrides, polyatomic ions
    - Know exceptions for O and H thoroughly
2. **Balancing Redox Equations**
    - Master BOTH methods (oxidation number method AND half-reaction method)
    - Practice at least 5 acidic and 5 basic medium examples
3. **Types of Redox Reactions**
    - Be able to identify and classify any given reaction
    - Know 2 examples of each type
4. **Standard Electrode Potential Table**
    - Memorise top 5 and bottom 5 values
    - Know how to use the table for predicting feasibility
5. **Daniell Cell and Electrode Processes**
    - Know diagram, definitions, salt bridge function
    - E°cell calculation

⚠️ Last Minute Checklist (1 Day Before Exam):

- ✅ Revise all 6 rules for oxidation number (especially O and H exceptions)
- ✅ Practice 5 oxidation number calculations including complex compounds
- ✅ Revise both balancing methods with one example each
- ✅ Memorise the 4 types of redox reactions with one equation each
- ✅ Know key E° values: F_2 (+2.87), MnO_4^- (+1.51), Cl_2 (+1.36), Cu (+0.34), H (+0.00), Zn (−0.76), Na (−2.71), Li (−3.05)
- ✅ Practice E°cell = E°cathode − E°anode
- ✅ Know OIL RIG mnemonic
- ✅ Revise disproportionation with H_2O_2 and Cl_2 examples
- ✅ Stock notation for at least 5 compounds
- ✅ Revise the Daniell cell diagram

🎯 Exam Writing Tips:

1. **Always show oxidation numbers** above atoms when justifying a redox reaction
2. **Show all steps** in balancing (partial marks available)
3. **Write half-reactions clearly** with electron count in ion-electron method
4. **State the medium** (acidic/basic) when writing balanced ionic equations
5. **Identify oxidant and reductant** explicitly when asked
6. **Use arrows** to show which element increased/decreased in oxidation number
7. **Verify** your balanced equation — check both atom count and charge on both sides

📝 Formula Sheet – Chapter 7: Redox Reactions

Concept	Key Formula / Rule
OIL RIG	Oxidation Is Loss; Reduction Is Gain (of electrons)
O.N. in neutral molecule	Sum of all oxidation numbers = 0
O.N. in polyatomic ion	Sum of all oxidation numbers = charge on ion
Oxygen (normal)	−2; Peroxide → −1; Superoxide → −½; OF_2 → +2
Hydrogen (normal)	+1; Metal hydrides → −1
Fluorine	Always −1
E°cell	$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$ (if > 0, reaction is feasible)
Oxidising agent	Higher E° value (gets reduced); increases O.N. of other species
Reducing agent	Lower E° value (gets oxidised); decreases O.N. of other species
Disproportionation	Same element: one oxidised + one reduced simultaneously

📚 QUICK REVISION NOTES

🔑 Key Definitions (Must Know Word-to-Word)

1. Oxidation (Classical): Addition of oxygen/electronegative element or removal of hydrogen/electropositive element from a substance.

2. Reduction (Classical): Removal of oxygen/electronegative element or addition of hydrogen/electropositive element to a substance.

3. Oxidation (Electronic): Loss of electron(s) by any species.

4. Reduction (Electronic): Gain of electron(s) by any species.

5. Oxidising Agent: A species that accepts electrons and oxidises other species; itself undergoes reduction.

6. Reducing Agent: A species that donates electrons and reduces other species; itself undergoes oxidation.

7. Oxidation Number: The oxidation state of an element in a compound, ascertained assuming complete electron transfer to the more electronegative element.

8. Disproportionation: A redox reaction in which the same element in an intermediate oxidation state is simultaneously oxidised and reduced.

9. Redox Couple: The oxidised and reduced forms of a substance involved in a half-reaction, represented as Oxidised form / Reduced form (e.g., Cu^{2+}/Cu).

10. Standard Electrode Potential (E°): Electrode potential at 298K with all ionic species at 1 M concentration and gases at 1 atm pressure; measured relative to SHE ($E^\circ = 0.00V$).

One-Line Facts (Speed Revision)

- Oxidation ↑ in O.N. ; Reduction ↓ in O.N.
- The oxidant is reduced; the reductant is oxidised
- F_2 is the strongest oxidising agent ($E^\circ = +2.87$ V)
- Li is the strongest reducing agent ($E^\circ = -3.05$ V)
- Reducing power of hydrohalic acids: HI > HBr > HCl > HF
- Oxidising power of halogens: $F_2 > Cl_2 > Br_2 > I_2$
- $CaCO_3 \rightarrow CaO + CO_2$ is NOT a redox reaction (no O.N. change)
- $Cl_2 + 2OH^- \rightarrow ClO^- + Cl^- + H_2O$ → This forms household bleach
- $KMnO_4$ is a self-indicator in redox titrations
- $E^\circ_{cell} > 0$ → spontaneous / feasible reaction
- Anode: Oxidation (electron source); Cathode: Reduction (electron sink)
- Salt bridge: maintains electrical neutrality; prevents mixing

🎯 Top 10 Numerical/Application Types (Practice These!)

1. Calculate oxidation number of specific element in complex ion/compound
2. Identify oxidation, reduction, oxidant, reductant in a given equation
3. Classify reaction as combination/decomposition/displacement/disproportionation
4. Balance redox equation in acidic medium (oxidation number method)
5. Balance redox equation in basic medium (half-reaction method)
6. Justify that a given reaction is/is not a redox reaction
7. Predict feasibility using E° values
8. Calculate E°cell from given E° values
9. Represent a galvanic cell using standard notation
10. Identify oxidant/reductant from E° table for given species

📚 Study Material Information

This comprehensive study material on **Redox Reactions (Chapter 7)** has been meticulously prepared following the latest CBSE curriculum and examination pattern for Class 11 Chemistry (2025-26 session). The content includes detailed explanations of all topics — classical and electronic concepts of redox, complete oxidation number rules with all exceptions, types of redox reactions with examples, both methods of balancing redox equations (Oxidation Number Method and Half-Reaction Method), redox titrations, electrode processes, the Daniell cell, standard electrode potentials, and the electrochemical series. Also included are NCERT-aligned content, tips and tricks, case-based questions as per new CBSE pattern, most expected questions for board exams, and a complete formula sheet.

Key Features of This Material:

- ✅ Complete Chapter 7 coverage with crystal-clear concepts
- ✅ Classical, Electronic, and Oxidation Number concepts of Redox
- ✅ All 6 rules for oxidation number with exceptions and worked examples
- ✅ Stock notation and oxidation state table
- ✅ 4 types of redox reactions: Combination, Decomposition, Displacement, Disproportionation
- ✅ Competitive electron transfer reactions and activity series
- ✅ Both balancing methods — step-by-step with multiple worked examples
- ✅ Redox Titrations — self-indicator, external indicator, iodometric method
- ✅ Daniell Cell — construction, working, terminology, electrode notation
- ✅ Standard Electrode Potential table (34 reduction reactions)
- ✅ Predicting feasibility of reactions using E° values
- ✅ Fractional oxidation numbers explained with examples
- ✅ 3 complete case studies as per latest CBSE pattern
- ✅ Most expected questions (1-mark, 2-mark, 3-mark, 5-mark)
- ✅ Common mistakes to avoid
- ✅ Exam preparation strategy and last-minute checklist
- ✅ Quick revision notes and formula sheet

🎯 Why Redox is a BRIDGE Chapter:

Chapter 7 is not just another chapter — it's the **bridge between Class 11 and Class 12 Chemistry!** Redox concepts are directly applied in:

- ✓ Electrochemistry (Class 12 — entire chapter!)
- ✓ Chemical Kinetics (rate of redox reactions)
- ✓ Hydrogen (fuel cells, ozone formation)
- ✓ p-Block Elements (oxoacids, oxidation states)
- ✓ d-Block & f-Block Elements (variable oxidation states)
- ✓ Metallurgy (extraction of metals by reduction)
- ✓ JEE & NEET (1–3 questions every year from this chapter!)

Master Redox = Foundation of Inorganic Chemistry!

Content is based on NCERT syllabus and CBSE guidelines for 2025-26.
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Disclaimer: This material is prepared as a comprehensive study aid for Class 11 students. While every effort has been made to ensure accuracy and alignment with CBSE curriculum, students are advised to refer to their NCERT textbooks and official CBSE guidelines for examination preparation. This material covers Chapter 7: Redox Reactions from Class 11 Chemistry NCERT textbook (Reprint 2025-26). All numerical values, formulas, and definitions are as per the latest NCERT edition.

📞 **Contact: +91 7869553517**
📍 **Location: Raipur | Indore**
🌐 **Website: www.mathlove.in**
📧 **Email: info@mathlove.in**