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Redox Reactions Class 11 Notes | CBSE Chemistry Chapter 7

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Redox Reactions Class 11 Notes | CBSE Chemistry

Chapter 7

Redox Reactions is Chapter 7 of CBSE Class 11 Chemistry — and it quietly powers almost everything around you, from the battery in your phone to the rusting of an iron gate. At its heart is one simple idea: electrons jumping from one species to another. Master that, and oxidation, reduction, balancing equations, and the whole of electrochemistry suddenly click into place.

By the end of these notes you will be able to assign oxidation numbers in seconds, identify the oxidising and reducing agents in any reaction, classify a reaction as combination, decomposition, displacement, or disproportionation, and balance even a messy redox equation by both the oxidation-number and half-reaction methods. This is a high-weightage chapter carrying roughly 6–8 marks in boards and a guaranteed question in NEET and JEE.

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Key Concepts

1. Oxidation and Reduction — the Classical View

The oldest way to spot a redox reaction is to watch oxygen and hydrogen move around. When a magnesium ribbon burns to form MgO , magnesium gains oxygen — it is **oxidised**. When CuO is heated with hydrogen to give copper, the oxide loses oxygen — it is **reduced**.

- **Oxidation:** addition of oxygen OR removal of hydrogen.
- **Reduction:** removal of oxygen OR addition of hydrogen.

This view works for simple cases but fails for reactions with no oxygen or hydrogen at all — so chemists upgraded to the electronic concept.

2. The Electron-Transfer (Modern) Concept

Every redox reaction is really electrons leaving one species and arriving at another. When zinc reacts with copper sulphate, zinc atoms hand over two electrons each — that is oxidation. Copper ions accept those electrons — that is reduction.

Oxidation = loss of electrons. Reduction = gain of electrons. The memory hook is **OIL RIG** — Oxidation Is Loss, Reduction Is Gain.

- Oxidation half: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ (electrons lost)
- Reduction half: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ (electrons gained)

Oxidation and reduction always happen together — electrons cannot simply vanish — which is why these are called **redox** (reduction–oxidation) reactions.

3. Oxidising and Reducing Agents

The species that *accepts* electrons is itself reduced, and by stripping electrons off the other species it acts as the **oxidising agent**. The species that *donates* electrons is itself oxidised, so it acts as the **reducing agent**.

- **Oxidising agent (oxidant):** gains electrons, gets reduced (e.g., KMnO_4 , $\text{K}_2\text{Cr}_2\text{O}_7$, O_2 , F_2).
- **Reducing agent (reductant):** loses electrons, gets oxidised (e.g., H_2 , C , Zn , SO_2).

Memory tip: the oxidising agent is the one that gets *reduced* — it does the opposite to itself of what its name suggests.

4. Oxidation Number (Oxidation State)

Oxidation number is the imaginary charge an atom would carry if every bond in the molecule were assumed to be fully ionic. It lets us track electron transfer even in covalent compounds where no real ions exist.

Rules for Assigning Oxidation Number

- Free elements (O_2 , H_2 , Na , Cl_2) have oxidation number **0**.
- For a monatomic ion, it equals the ionic charge ($\text{Na}^+ = +1$, $\text{S}^{2-} = -2$).
- Oxygen is usually **−2** (but -1 in peroxides like H_2O_2 , and $+2$ in OF_2).
- Hydrogen is usually **+1** (but -1 in metal hydrides like NaH).
- Fluorine is always **−1**; alkali metals always $+1$; alkaline earth metals always $+2$.
- Sum of oxidation numbers = 0 for a neutral molecule, and = charge for a polyatomic ion.

Worked rule example: In KMnO_4 , $\text{K} = +1$ and $\text{O} = -2$ ($\times 4 = -8$). So $+1 + \text{Mn} + (-8) = 0 \rightarrow \text{Mn} = \mathbf{+7}$.

5. Redox in Terms of Oxidation Number

The electron-transfer idea is restated cleanly using oxidation numbers, which is the form examiners love.

- **Oxidation:** an *increase* in oxidation number.
- **Reduction:** a *decrease* in oxidation number.

In $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$, zinc goes $0 \rightarrow +2$ (oxidised) and copper goes $+2 \rightarrow 0$ (reduced). Spotting the change in oxidation number is the fastest way to identify redox in an exam.

6. Types of Redox Reactions

CBSE classifies redox reactions into four neat families. Recognising the family often gives you the products instantly.

(a) Combination Reactions

Two or more substances combine into one, with at least one element changing oxidation number.



(b) Decomposition Reactions

A single compound breaks into two or more products; it is the reverse of combination.



(c) Displacement Reactions

One element displaces another from its compound. These can be metal or non-metal displacements.



(d) Disproportionation Reactions

The *same element* in an intermediate oxidation state is simultaneously oxidised and reduced.



[TABLE: Type → key feature → example. Combination: many→one; Decomposition: one→many; Displacement: element swaps in; Disproportionation: same element up and down.]

7. Balancing Redox Reactions — Oxidation Number Method

This method balances the total increase and decrease in oxidation number, since electrons lost must equal electrons gained.

Steps:

- Write the skeletal equation and assign oxidation numbers.
- Identify atoms whose oxidation number changes; find the increase and decrease.
- Multiply by suitable integers so total increase = total decrease.
- Balance the remaining atoms (O by adding H_2O , H by adding H^+), then balance charge.

Example skeleton: in $\text{MnO}_4^- + \text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + \text{Fe}^{3+}$, Mn falls $+7 \rightarrow +2$ (gain 5e^-) and Fe rises $+2 \rightarrow +3$ (lose 1e^-), so 5 Fe^{2+} are needed per MnO_4^- .

8. Balancing Redox Reactions — Half-Reaction (Ion-Electron) Method

Here the reaction is split into an oxidation half and a reduction half, each balanced separately, then recombined. This is the cleaner method for solutions.

Steps (acidic medium):

- Split into two half-reactions.
- Balance atoms other than O and H first.
- Balance O by adding H_2O , balance H by adding H^+ .
- Balance charge by adding electrons (e^-).
- Equalise electrons in both halves, then add and cancel.

Basic medium: do the same, then add OH^- to both sides to neutralise H^+ into H_2O .

Reduction half example: $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$. **Oxidation half:** $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$. Multiply the iron half by 5 and add.

9. Redox Reactions and Electrode Processes

If we physically separate the two half-reactions, the electrons are forced to travel through an external wire instead of jumping directly — and a flowing electron is electricity. This is the leap from chemistry to a working cell.

- **Anode:** oxidation occurs here (electrons released); it is the negative terminal in a galvanic cell.
- **Cathode:** reduction occurs here (electrons consumed); it is the positive terminal in a galvanic cell.

Memory hook: *An Ox* (Anode = Oxidation) and *Red Cat* (Reduction = Cathode).

10. Electrochemical (Galvanic) Cells — an Introduction

A **galvanic (voltaic) cell** converts the chemical energy of a spontaneous redox reaction directly into electrical energy. The classic example is the Daniell cell, built from the Zn–Cu reaction.

[DIAGRAM: Daniell cell — Zn rod in ZnSO_4 (anode, $-$) connected through a voltmeter to a Cu rod in CuSO_4 (cathode, $+$), with a salt bridge joining the two beakers.]

- **At the anode:** $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ (oxidation).
- **At the cathode:** $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ (reduction).
- **Salt bridge:** completes the circuit and keeps both solutions electrically neutral.

The **electrode potential** of each electrode measures its tendency to lose or gain electrons; the difference between the two gives the cell's EMF. Standard hydrogen electrode (SHE) is taken as the zero reference (0 V).

11. Applications of Redox Reactions

Redox is not just an exam topic — it runs the modern world.

- **Batteries and cells:** dry cells, lead storage batteries, fuel cells.
- **Corrosion:** rusting of iron is the oxidation of iron — a costly redox process.
- **Metallurgy:** extraction of metals by reduction of their ores (e.g., iron from haematite).
- **Bleaching and disinfection:** by oxidising agents like Cl_2 and KMnO_4 .
- **Photosynthesis and respiration:** the redox engines of life.

Weightage in Board & Entrance Exams

Exam	Typical Weightage	Most-Tested Areas
CBSE Board (Class 11)	6–8 marks	Oxidation number, identifying oxidant/reductant, balancing redox equations
JEE Main / Advanced	1–2 questions	Oxidation states, disproportionation, balancing in acidic/basic medium
NEET	1–2 questions	Oxidation number, types of redox reactions, oxidising/reducing agents

[TABLE: Question-type split — VSA (1 mark): definitions, oxidation-number calculation; SA (2–3 marks): identify oxidant/reductant, classify reaction type; LA (5 marks): balancing redox equations by both methods, Daniell cell.]

Important Definitions

Term	Definition
Oxidation	Loss of electrons / increase in oxidation number / addition of oxygen
Reduction	Gain of electrons / decrease in oxidation number / removal of oxygen
Redox reaction	A reaction in which oxidation and reduction occur simultaneously
Oxidising agent	Species that gains electrons and is itself reduced
Reducing agent	Species that loses electrons and is itself oxidised
Oxidation number	Imaginary charge on an atom if all bonds were assumed ionic
Disproportionation	Same element in one oxidation state is both oxidised and reduced
Anode	Electrode where oxidation occurs (negative terminal in a galvanic cell)
Cathode	Electrode where reduction occurs (positive terminal in a galvanic cell)
Galvanic cell	Device that converts chemical energy of a spontaneous redox reaction into electrical energy

Solved Examples

Example 1

Find the oxidation number of sulphur in H_2SO_4 .

Answer: $\text{H} = +1$ ($\times 2 = +2$), $\text{O} = -2$ ($\times 4 = -8$). So $2 + \text{S} - 8 = 0 \rightarrow \text{S} = \mathbf{+6}$.

Example 2

In the reaction $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$, identify the oxidising and reducing agents.

Answer: Zn goes $0 \rightarrow +2$ (oxidised) so it is the **reducing agent**; Cu^{2+} goes $+2 \rightarrow 0$ (reduced), so CuSO_4 is the **oxidising agent**.

Example 3

Find the oxidation number of chromium in $\text{K}_2\text{Cr}_2\text{O}_7$.

Answer: $\text{K} = +1$ ($\times 2 = +2$), $\text{O} = -2$ ($\times 7 = -14$). So $2 + 2\text{Cr} - 14 = 0 \rightarrow 2\text{Cr} = 12 \rightarrow \text{Cr} = \mathbf{+6}$.

Example 4

Show that $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$ is a disproportionation reaction.

Answer: Oxygen in H_2O_2 is -1 . In H_2O it becomes -2 (reduced) and in O_2 it becomes 0 (oxidised). The same element is both oxidised and reduced, so it is **disproportionation**.

Example 5

Balance the reduction half-reaction for MnO_4^- to Mn^{2+} in acidic medium.

Answer: $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$. (O balanced by 4 H_2O , H by 8 H^+ , charge by 5 electrons.)

Example 6

Find the oxidation number of nitrogen in NH_4^+ and in NO_3^- .

Answer: In NH_4^+ : $\text{N} + 4(+1) = +1 \rightarrow \text{N} = -3$. In NO_3^- : $\text{N} + 3(-2) = -1 \rightarrow \text{N} = +5$.

Important Questions for Board Exams

1-Mark Questions (VSA)

1. Define oxidation in terms of electrons.
2. What is the oxidation number of oxygen in OF_2 ?
3. Name the species that gets reduced in a redox reaction.
4. What is a disproportionation reaction? Give one example.
5. At which electrode does oxidation occur in a galvanic cell?

2–3-Mark Questions (SA)

1. Identify the oxidising and reducing agents in: $\text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + \text{Cl}_2 + 2\text{H}_2\text{O}$.
2. Explain why the same element can be both oxidised and reduced in a disproportionation reaction, using an example.
3. Calculate the oxidation number of phosphorus in H_3PO_4 and of manganese in KMnO_4 .
4. Distinguish between a combination and a displacement redox reaction with one example each.

5-Mark Questions (LA)

1. Balance the equation $\text{MnO}_4^- + \text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + \text{Fe}^{3+}$ in acidic medium by the ion-electron (half-reaction) method.
2. Describe the construction and working of a Daniell cell. Write the electrode reactions and state the function of the salt bridge.
3. Balance $\text{Cr}_2\text{O}_7^{2-} + \text{SO}_3^{2-} \rightarrow \text{Cr}^{3+} + \text{SO}_4^{2-}$ in acidic medium and identify the oxidising agent.

Quick Revision Points

- OIL RIG: Oxidation Is Loss, Reduction Is Gain (of electrons)
- Oxidation = increase in oxidation number; reduction = decrease
- Oxidising agent is itself reduced; reducing agent is itself oxidised
- Oxidation number of free elements = 0; O is usually -2 ; H usually $+1$; F always -1
- Four types: combination, decomposition, displacement, disproportionation
- Disproportionation: same element both oxidised and reduced (e.g., H_2O_2)
- Balancing: oxidation-number method (balance net change) or half-reaction method (split, balance, recombine)
- In acidic medium balance O with H_2O and H with H^+ ; in basic medium add OH^-
- An Ox, Red Cat: Anode = Oxidation, Cathode = Reduction
- Galvanic cell converts chemical energy of a spontaneous redox reaction into electricity; salt bridge keeps it neutral
- SHE is the reference electrode with potential 0 V

Next Chapter: Equilibrium Class 11 Notes



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