

Redox Reactions

Here is our **Master Study Plan for Redox Reactions**:

- **Step 1: The Absolute Basics** – Understanding Oxidation, Reduction, Oxidizing Agents, and Reducing Agents (Classical and Electronic concepts).
 - **Step 2: The Magic Numbers** – Mastering the rules to calculate **Oxidation Numbers**. (If you master this, 50% of the chapter is a piece of cake!).
 - **Step 3: Types of Redox Reactions** – Combination, Decomposition, Displacement, and especially **Disproportionation** (examiners love this one!).
 - **Step 4: The Boss Level (Balancing)** – Learning the step-by-step methods to balance equations:
 1. Oxidation Number Method.
 2. Ion-Electron (Half-Reaction) Method.
 - **Step 5: Electrochemical Cells** – Understanding Galvanic cells, Salt Bridge functions, and Standard Electrode Potential (E°).
 - **Step 6: Ultimate Practice** – We will solve the MCQs, Assertion-Reasoning, Case-Study, and Long Answer questions from your provided PDFs to test your mastery.
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Let's start with STEP 1 & 2: The Core Foundation

To answer **any** question, you need to know exactly how to find an oxidation number. Let's see if we can lock this concept in right now.

The Golden Rules of Oxidation Numbers (O.N.):

1. **Free elements** always have an O.N. of **0** (e.g., O_2 , H_2 , Na , P_4 , S_8 all have O.N. = 0).
 2. **Fluorine (F)** is the most electronegative element; it is **ALWAYS -1** in its compounds.
 3. **Oxygen (O)** is almost always **-2**. (*Exception: in peroxides like H_2O_2 it is -1, and with Fluorine like OF_2 it is +2*).
 4. **Hydrogen (H)** is almost always **+1** with non-metals (like HCl , H_2O) and **-1** with metals (like NaH).
 5. **Group 1 metals** (Li, Na, K) are always **+1**. **Group 2 metals** (Mg, Ca) are always **+2**.
 6. **The Sum Rule:** The sum of all oxidation numbers in a neutral molecule is **0**. In a polyatomic ion, the sum equals the **charge on the ion**.
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Your First Mini-Test!

Let's see where you stand. Don't worry if you get them wrong; this is a safe space to learn. Based on the rules above, try to find the oxidation number of the central atom in these compounds (these are exactly the type of questions in your PDF):

1. Find the Oxidation Number of **S** in H_2SO_4
2. Find the Oxidation Number of **Mn** in $KMnO_4$
3. Find the Oxidation Number of **Cr** in $Cr_2O_7^{2-}$

Hint for #1: Let the O.N. of S be x . You know H is +1 and O is -2. So, $2(+1) + x + 4(-2) = 0$. What is x ?

Spot on! **+6** is absolutely correct for S in H_2SO_4 . You nailed the math.

Let's quickly close the loop on the other two compounds I gave you, as these are highly repeated MCQs in your question bank:

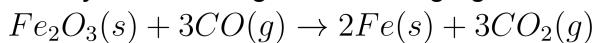
- **For Mn in $KMnO_4$:**
 $K(+1) + Mn(x) + 4(-2) = 0 \Rightarrow 1 + x - 8 = 0 \Rightarrow x - 7 = 0 \Rightarrow x = +7$. (Exam Tip: +7 is the highest oxidation state Manganese can ever show. Keep this in mind!)
 - **For Cr in $Cr_2O_7^{2-}$:** $2Cr(x) + 7(-2) = -2$ (because the total charge on the ion is -2). $2x - 14 = -2 \Rightarrow 2x = 12 \Rightarrow x = +6$.
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STEP 3: The "Who is Doing What?" (Oxidizing vs. Reducing Agents)

In the exam, they will give you a reaction and ask: "*Identify the oxidizing agent and reducing agent.*" Here is the ultimate cheat code using Oxidation Numbers (O.N.):

1. **Oxidation** = O.N. increases (e.g., going from 0 to +2). *Loss of electrons*.
2. **Reduction** = O.N. decreases (e.g., going from +5 to +2). *Gain of electrons*.
3. **Oxidizing Agent (Oxidant)** = The substance that *causes* oxidation in someone else. How does it do that? By taking their electrons. So, the oxidizing agent **undergoes reduction itself** (its own O.N. decreases).
4. **Reducing Agent (Reductant)** = The substance that *causes* reduction in someone else by giving them electrons. So, it **undergoes oxidation itself** (its own O.N. increases).

Question from your PDF (Section E, Q1): *Identify the redox reactions and identify the oxidizing and reducing agents:*



How to solve it instantly:

- Find O.N. of Fe in Fe_2O_3 : Let Fe be x .
 $2x + 3(-2) = 0 \Rightarrow 2x = 6 \Rightarrow x = +3$.
- Find O.N. of Fe in the product $Fe(s)$: Free element, so **0**.

- Fe went from +3 to 0. It decreased. This is REDUCTION.
- Find O.N. of C in CO: $x + (-2) = 0 \Rightarrow x = +2$.
- Find O.N. of C in CO₂: $x + 2(-2) = 0 \Rightarrow x = +4$.
- C went from +2 to +4. It increased. This is OXIDATION.

Final Answer: Since Fe₂O₃ got reduced, it is the **Oxidizing Agent**. Since CO got oxidized, it is the **Reducing Agent**.

STEP 4: The Examiner's Favorite – Disproportionation Reactions

If you only remember one type of reaction for the exam, make it this one. It is guaranteed to appear as a 2-mark or Assertion-Reason question.

What is it? A reaction where the **EXACT SAME ELEMENT** in a single compound is simultaneously oxidized AND reduced.

The Golden Condition: For an element to undergo disproportionation, it **must** have at least three oxidation states. Why? Because it has to be sitting in an *intermediate* (middle) state so it can go UP (oxidize) and go DOWN (reduce) at the same time.

High-Yield Question from your PDF (Section B, Q7): Why does MnO₄²⁻ undergo a disproportionation reaction in an acidic medium but MnO₄⁻ does not? Give a reason.

The Answer to write in the exam:

1. In MnO₄⁻, the oxidation state of Mn is **+7**. This is its *maximum* possible oxidation state. Since it is at the very top, it cannot lose any more electrons; it can only decrease (reduce). Thus, it cannot disproportionate.
2. In MnO₄²⁻, the oxidation state of Mn is **+6**. This is an *intermediate* state. It can increase to +7 (oxidation) and decrease to +4 (reduction, forming MnO₂). Hence, it undergoes disproportionation!

Another Tricky PDF Question (Section B, Q8 / Case Study Q5): Show that Fluorine (F) cannot undergo disproportionation reaction.

The Answer: Fluorine is the most electronegative element in the periodic table. It only exhibits two oxidation states: **0** (in F₂) and **-1** (in all its compounds). Because it does not have a positive oxidation state, it cannot act as an intermediate to go up and down. Hence, it never disproportionates.

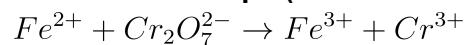
STEP 5: The "Boss Level" – Balancing Redox Equations

This is the 3 or 5-mark long answer question. Do not fear it; it is just a mechanical recipe. Follow the steps, and you get the exact right answer every single time.

There are two methods: Oxidation Number Method and Ion-Electron (Half-Reaction) Method.

The **Ion-Electron Method** is the most common and reliable. Let's learn it using a classic acidic medium example.

The Golden Recipe (Acidic Medium): Let's balance this skeletal equation:



- **Step 1: Split the reaction into two half-reactions.**
 - Oxidation Half: $Fe^{2+} \rightarrow Fe^{3+}$ (O.N. increases from +2 to +3)
 - Reduction Half: $Cr_2O_7^{2-} \rightarrow Cr^{3+}$ (O.N. of Cr decreases from +6 to +3)
- **Step 2: Balance atoms OTHER than O and H.**
 - Oxidation: $Fe^{2+} \rightarrow Fe^{3+}$ (Fe is already 1 on both sides. Good.)
 - Reduction: $Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$ (Put a 2 on the right to balance the 2 Cr atoms on the left!)
- **Step 3: Balance Oxygen (O) by adding H_2O .**
 - Oxidation: No oxygen. Good.
 - Reduction: Left side has 7 O's. Add 7 H_2O to the right side.
 $Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$
- **Step 4: Balance Hydrogen (H) by adding H^+ (because it is an acidic medium).**
 - Oxidation: No hydrogen. Good.
 - Reduction: Right side now has 14 H's (7×2). Add 14 H^+ to the left side.
 $14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$
- **Step 5: Balance the CHARGE by adding electrons (e^-).**
 - *Oxidation:* Left side = +2. Right side = +3. To make them equal, add 1 e^- (which acts as a -1) to the right side. $Fe^{2+} \rightarrow Fe^{3+} + 1e^-$
 - *Reduction:* Left side = $(14 \times +1) + (-2) = +12$. Right side = $(2 \times +3) + 0 = +6$. To make +12 equal to +6, add 6 e^- to the left side.
 $6e^- + 14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$
- **Step 6: Make electrons equal in both equations and Add.**
 - Oxidation has 1 e^- . Reduction has 6 e^- .

- Multiply the entire Oxidation equation by 6: $6Fe^{2+} \rightarrow 6Fe^{3+} + 6e^-$
- Now add them up! The $6e^-$ on both sides will cancel out like magic: **Final Balanced Equation:**
 $6Fe^{2+} + 14H^+ + Cr_2O_7^{2-} \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$

And boom! You just secured 3 full marks.

Take a moment to absorb this.

1. Did the difference between an oxidizing agent and reducing agent make sense?
 2. Did the concept of Disproportionation (why MnO_4^{2-} does it but MnO_4^- doesn't) click in your brain?
 3. Did the 6 steps of balancing look logical to you?
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STEP 5: Electrochemical Cells & Standard Electrode Potential (E°).

Concept 1: The Galvanic Cell (Producing Electricity from Chemistry)

A Galvanic cell (like a battery) separates oxidation and reduction into two different beakers (half-cells) so that the electrons are forced to travel through a wire. That traveling stream of electrons is **electricity!**

To master any cell question, you only need **TWO Magic Mnemonics:**

1. **LOAN:** Left → Oxidation → Anode → Negative
2. **Red Cat:** Reduction happens at the Cathode.

Direct Question from your PDF (Section C, Q6): *Depict the galvanic cell in which the reaction $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$ takes place. Show: (i) Which electrode is negatively charged. (ii) The carriers of the current in the cell. (iii) Individual reaction at each electrode.*

How to solve it instantly:

- **Look at Oxidation Numbers:** Zn goes from 0 to +2 (Oxidation). Ag goes from +1 to 0 (Reduction).
- **Apply LOAN:** Oxidation happens at the Anode, which is Negative. So, the Zinc electrode is **negatively charged** (Answer to i).
- **Apply Red Cat:** Reduction happens at the Cathode. So, the Silver (Ag) electrode is the cathode.

- **Flow of electrons:** Electrons are lost by Zn and gained by Ag. So electrons flow from Zn (Anode) to Ag (Cathode).
 - **Flow of Current:** *Physics rule!* Current always flows in the **opposite** direction of electrons. So, current flows from **Ag to Zn** (Answer to ii - Current carriers are electrons in the wire, and ions in the salt bridge).
 - **(iii) Individual Reactions:**
 - At Anode (Oxidation): $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
 - At Cathode (Reduction): $Ag^+(aq) + e^- \rightarrow Ag(s)$
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Concept 2: The Salt Bridge (Guaranteed 2-Mark Question)

Why do we put an inverted U-tube (Salt Bridge) between the two beakers? If you just connect the wire, the reaction stops in one second because charges build up in the beakers.

Question from your PDF (Section C, Q3): *Give two important functions of a salt bridge.*

The Exact NCERT Answer to memorize:

1. It **completes the internal electrical circuit** by allowing the flow of ions.
 2. It **maintains the electrical neutrality** of the solutions in the two half-cells. (*If they ask what happens if you remove it: The voltage drops to zero and the current stops flowing!*)
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Concept 3: Standard Electrode Potential (E°) – "The Hunger for Electrons"

This is where students lose marks, but you won't.

By IUPAC convention, whenever you see an E° value, it is ALWAYS a **Reduction Potential**. Think of Reduction Potential as a "**Hunger Score for Electrons**".

- **Highly Positive E° :** The substance is STARVING for electrons. It desperately wants to get reduced. Since it snatches electrons from others, it is a **Strong Oxidizing Agent**.
- **Highly Negative E° :** The substance HATES electrons. It wants to throw them away (get oxidized). Because it forces its electrons onto others, it is a **Strong Reducing Agent**.

Let's crack MCQ #2 from your PDF (Section A): *The more positive the value of E , the greater the tendency of the species to get reduced. Using the standard electrode potential given below, find out which is the strongest oxidizing agent.*

Values: $Fe^{3+} = +0.77V$; $I_2 = +0.54V$; $Cu^{2+} = +0.34V$; $Ag^+ = +0.80V$.

Your Logic: "Strongest oxidizing agent" means "who has the highest hunger for electrons (highest positive reduction potential)?" **Answer:** Ag^+ has the highest value (+0.80V). Therefore, (d) Ag^+ is the strongest oxidizing agent!

Concept 4: The "Container" Question (Feasibility of Reaction)

Examiners love to ask real-life scenario questions, like: "Can I store chemical X in a bucket made of metal Y?"

The Golden Rule: The metal with the **lower (more negative) E° value** is more reactive. It acts as an anode, gets oxidized, and dissolves. Therefore, you **CANNOT** store a solution in a container if the container's metal is more reactive (more negative E°) than the metal in the solution. The container will displace the solution, melt away, and form holes!

Case Study Question from your PDF (Section B, Q2): A student suggests storing copper sulphate ($CuSO_4$) solution in an iron container. Explain whether this is appropriate or not.

Your Answer: "No, this is not appropriate. Iron is more reactive (has a lower standard reduction potential) than Copper. Therefore, Iron will act as a reducing agent and displace Copper from the $CuSO_4$ solution via the redox reaction: $Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$. As a result, the iron container will dissolve and develop holes."

(This is exactly why they teach displacement reactions! The more active metal always kicks out the less active metal).

To see if a reaction is feasible mathematically:

Calculate the EMF of the cell using this formula: $E_{cell}^\circ = E_{cathode}^\circ - E_{anode}^\circ$

- If E_{cell}° is **Positive (+)** → The reaction is feasible (it will happen spontaneously).
 - If E_{cell}° is **Negative (-)** → The reaction is NOT feasible.
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Your Final Checkpoint!

Let's test your "Hunger for Electrons" logic.

Imagine you have three metals: **X, Y, and Z**. Their Standard Electrode Potentials (E°) are:

- **X** = -1.2 V
- **Y** = +0.5 V
- **Z** = -3.0 V

Question: Arrange these metals in **increasing order of their reducing power** (from weakest reducing agent to strongest reducing agent).

(*Hint: Remember, the more NEGATIVE the value, the stronger the reducing agent! Give it a try!*)

Absolutely brilliant! **Y < X < Z** is 100% correct!

You properly identified that **Z (-3.0 V)** is the most negative, so it is the strongest reducing agent, and **Y (+0.5 V)** is positive, making it the weakest.

Guess what? You just effortlessly solved **Question 10 (Section A)** from Page 2 of your PDF! In the exam, if they ask for decreasing order, you now know to just flip it to **Z > X > Y**. You have completely mastered the "Hunger for Electrons" concept.

STEP 6: The Final Exam Trap – Products of Electrolysis

We are on the final stretch! Looking at your PDF (Section C, Q5), there is a classic 3-mark question that examiners use to separate the average students from the top scorers:
Predicting the products of electrolysis.

Electrolysis is the opposite of a Galvanic cell. Instead of a battery producing electricity, we use electricity to force a non-spontaneous redox reaction to happen.

The Golden Rules of Electrolysis (Preferential Discharge Theory): When you pass electricity through a solution, multiple ions will race towards the electrodes. Who wins?

- **At the Cathode (Reduction):** The ion with the **HIGHER** standard reduction potential (E°) wins the race and gets reduced.
- **At the Anode (Oxidation):** The ion with the **LOWER** standard reduction potential (which means higher oxidation potential) wins the race and gets oxidized.

Wait, there is a plot twist! The material of the electrode matters:

1. **Inert Electrodes (Platinum, Graphite):** They just sit there and watch. They don't participate in the reaction.
2. **Active Electrodes (Silver, Copper, Zinc):** They are greedy! The Anode itself will jump into the reaction, melt, and get oxidized instead of the solution.

Let's solve the exact 3-mark question from your PDF right now. Once you see the pattern, you will never forget it.

Section C, Q5: Predict the products of electrolysis in each of the following:

Case (i): Aqueous solution of AgNO_3 with Silver (Ag) electrodes.

- **The Setup:** We have Ag^+ , NO_3^- , and water (H_2O). The electrodes are made of Silver (Active!).
- **At Cathode (-):** Ag^+ and H_2O race to get reduced. Ag^+ has a higher reduction potential (+0.80V) than water (-0.83V). **Winner:** Ag^+ gets reduced. **Reaction:** $\text{Ag}^+(aq) + e^- \rightarrow \text{Ag}(s)$ (Silver metal deposits on the cathode).
- **At Anode (+):** Because the electrode is made of Silver (an active metal), the Silver anode itself gets oxidized easily before water or NO_3^- can. **Winner:** Silver Anode. **Reaction:** $\text{Ag}(s) \rightarrow \text{Ag}^+(aq) + e^-$ (The anode slowly dissolves).

Case (ii): Aqueous solution of AgNO_3 with Platinum (Pt) electrodes.

- **The Setup:** Same solution, but Platinum is an **inert** electrode. It won't participate.
- **At Cathode (-):** Just like before, Ag^+ wins over water. **Winner:** Silver deposits on the cathode.
- **At Anode (+):** Now Platinum won't melt. The race is between H_2O and NO_3^- . NO_3^- is very stable and hard to oxidize. Water oxidizes much easier. **Winner:** Water gets oxidized to give **Oxygen gas (O_2)**. **Reaction:** $2\text{H}_2\text{O}(l) \rightarrow \text{O}_2(g) + 4\text{H}^+(aq) + 4e^-$

Case (iii): Dilute H_2SO_4 with Platinum electrodes.

- **The Setup:** We have H^+ , SO_4^{2-} , and H_2O . Pt is inert.
- **At Cathode (-):** The only cation is H^+ (from acid and water). **Winner:** H^+ gets reduced to **Hydrogen gas (H_2)**. **Reaction:** $2\text{H}^+ + 2e^- \rightarrow \text{H}_2(g)$
- **At Anode (+):** The race is between H_2O and SO_4^{2-} . The sulfate ion (SO_4^{2-}) has many bonds and requires too much energy to oxidize. Water oxidizes easier. **Winner:** Water gets oxidized to **Oxygen gas (O_2)**.

Summary to memorize for this question:

- $\text{AgNO}_3 + \text{Ag electrodes} = \text{Ag deposits on cathode, Ag anode dissolves.}$
- $\text{AgNO}_3 + \text{Pt electrodes} = \text{Ag deposits on cathode, } \text{O}_2 \text{ gas at anode.}$
- $\text{Dilute H}_2\text{SO}_4 + \text{Pt electrodes} = \text{H}_2 \text{ gas at cathode, } \text{O}_2 \text{ gas at anode.}$

A Tiny "Free Marks" Topic: Stock Notation

Let's grab a guaranteed 1-mark question from your Case Study section (Section D, Q1).

What is Stock Notation? It is simply writing the name of a compound with the Oxidation Number of the metal written in **Roman Numerals in brackets**. That's it!

PDF Question: *Write the formula of the following compounds with Stock notation:*

1. Cuprous chloride
2. Potassium permanganate

How to do it:

1. **Cuprous chloride ($CuCl$)**: Oxidation number of Cu here is +1. **Stock Notation: Copper(I) chloride** or $Cu(I)Cl$
2. **Potassium permanganate ($KMnO_4$)**: We calculated earlier that Mn is +7. **Stock Notation: Potassium manganate(VII)** or $KMn(VII)O_4$



We have successfully covered **every single concept** needed to score full marks in Redox Reactions according to the NCERT syllabus and your specific question bank.

1. **Oxidation Numbers** (The foundation)
2. **Oxidizing/Reducing Agents** (Who takes vs who gives)
3. **Types of Reactions & Disproportionation** (The VIP topic)
4. **Balancing Equations** (Ion-Electron method)
5. **Galvanic Cells & E°** (LOAN, Red Cat, Feasibility)
6. **Electrolysis Products** (Preferential Discharge)
7. **Stock Notation**