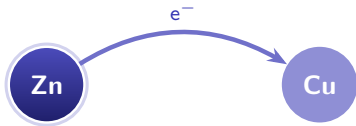


Electrochemistry

A-Level Chemistry ■ Topic 6

A-Level Chemistry



What we will cover

- 1 Oxidation number
- 2 Redox & OIL RIG
- 3 Balancing redox equations
- 4 Disproportionation
- 5 Oxidising & reducing agents



redox pushes electrons round



Oxidation number

The **oxidation number** shows electrons lost or gained vs the free element.

- element = 0; simple ion = its charge
- H = +1, O = -2, F = -1
- they sum to 0 (or the ion's charge)

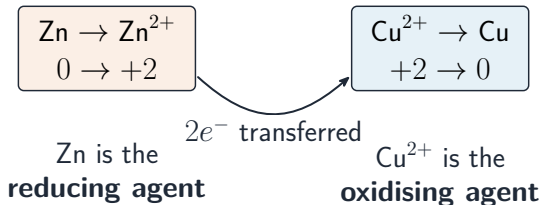


they sum to 0: $(+1) + \text{Mn} + (-8) = 0$, so $\text{Mn} = +7$

Redox and OIL RIG

oxidation: loses $2e^-$
(oxidation number $\uparrow$)

reduction: gains $2e^-$
(oxidation number $\downarrow$)

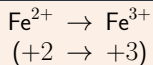


A **redox** reaction transfers electrons:

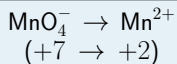
- **oxidation** is loss (number up)
- **reduction** is gain (number down)

Remember **OIL RIG**; they always happen together.

Balancing redox equations



oxidised: rises 1 ($\times 5$)



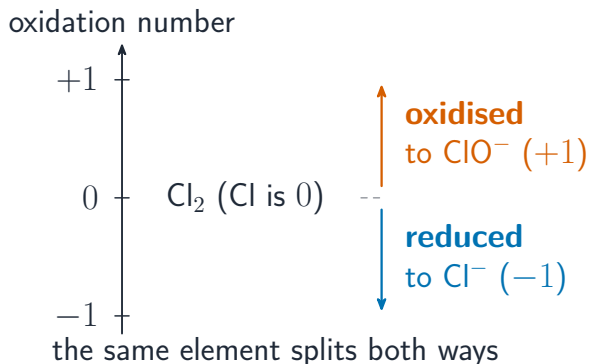
reduced: falls 5 ($\times 1$)

5 electrons lost = 5 gained $\Rightarrow$ ratio **5 Fe²⁺ : 1 MnO₄⁻**

Use oxidation-number changes:

The total **rise** in oxidation number must equal the total **fall** (every electron lost is gained) – this fixes the ratio.

Disproportionation



Disproportionation is when the **same** element is both oxidised and reduced at once. Chlorine (0) $\rightarrow$ ClO^- (+1) *and* Cl^- (-1) in cold water.

Oxidising and reducing agents



- an **oxidising agent** takes electrons – it is itself **reduced**
- a **reducing agent** gives electrons – it is itself **oxidised**

Rusting is the oxidation of iron by oxygen.

Exam tips

- Apply the oxidation-number rules in order: O is -2 **except** in peroxides (-1); H is $+1$ **except** in metal hydrides (-1).
- In a redox equation the **total increase equals the total decrease** in oxidation number – use this to fix the ratio, then balance O with H_2O and H with H^+ .
- An **oxidising agent is itself reduced**; name both what is oxidised/reduced *and* the agent for full marks.
- **Disproportionation** is the same element both oxidised and reduced – show both oxidation-number changes.



Worked example – oxidation number

Find the oxidation number of sulfur in SO_4^{2-} .

- Let the oxidation number of S be x . Each O is -2 .
- The numbers must add up to the **ion's charge**, -2 :
- $x + 4(-2) = -2$
- $x = -2 + 8 = +6$.

Oxidation numbers sum to the **charge** (0 for a neutral compound). O is -2 and H is $+1$ – except in peroxides and metal hydrides.

Topic 6 – key takeaways

1

Oxidation number

electrons lost/gained vs the element

2

OIL RIG

oxidation is loss; reduction is gain

3

Balancing

total rise = total fall

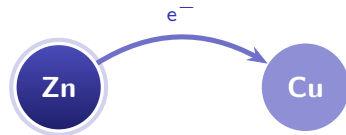
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Agents

oxidising agent is reduced;
reducing agent oxidised

Check yourself

- 1 Oxidation number of Mn in KMnO_4 ?
- 2 Does oxidation lose or gain electrons?
- 3 What is it called when one element is both oxidised and reduced?
- 4 An oxidising agent is itself what?



Answers 1. +7 2. loses electrons 3. disproportionation 4. reduced