

Oxidation numbers and balancing redox equations

Chemistry grades 11-12 (A Level · IB · Class 11-12) · Lesson

Quick answer: Redox reactions transfer electrons: oxidation is loss of electrons and an increase in oxidation number, reduction is gain of electrons and a decrease. Oxidation numbers follow simple rules, and half-equations balanced for atoms and charge are added so that electrons cancel.

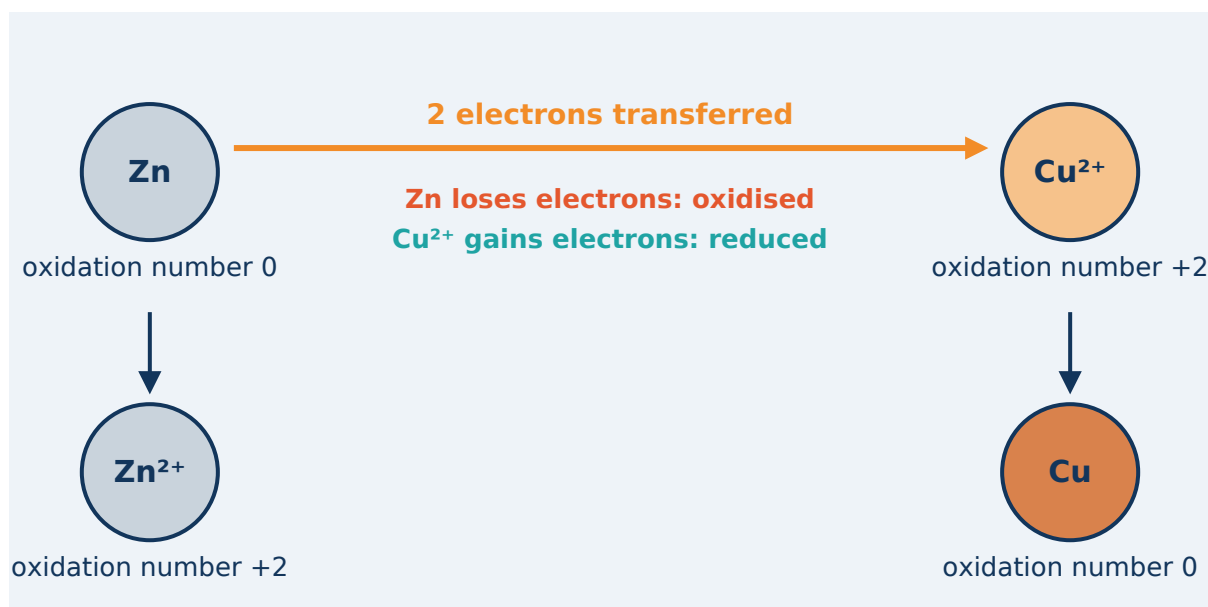
Level: Chemistry grades 11-12 (A Level · IB · Class 11-12) · **Curriculum:** IB R3.2

What are oxidation and reduction?

Redox reactions are reactions in which electrons are transferred. They power batteries, rust iron, bleach cloth and burn fuels. A simple memory aid is OIL RIG.

Definition

Oxidation is loss of electrons (oxidation number increases). Reduction is gain of electrons (oxidation number decreases). An oxidising agent takes electrons and is itself reduced; a reducing agent gives electrons and is itself oxidised.



Both processes always happen together. In the reaction $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$, zinc is the reducing agent and Cu²⁺ is the oxidising agent.

How do you work out an oxidation number?

An oxidation number is the charge an atom would have if all its bonds were treated as fully ionic. It lets you follow electrons even in covalent compounds. Apply the rules in order.

Rule	Examples
Uncombined element: 0	Cu, O ₂ , S ₈ all 0
Simple ion: equals its charge	Mg ²⁺ is +2; Cl ⁻ is -1
Sum for a compound is 0; for an ion it equals the charge	SO ₄ ²⁻ sums to -2
Group 1 is +1, group 2 is +2, aluminium is +3	Na in NaCl is +1
Hydrogen is +1 (-1 in metal hydrides)	H in NaH is -1
Fluorine is -1; oxygen is -2 (-1 in peroxides)	O in H ₂ O ₂ is -1

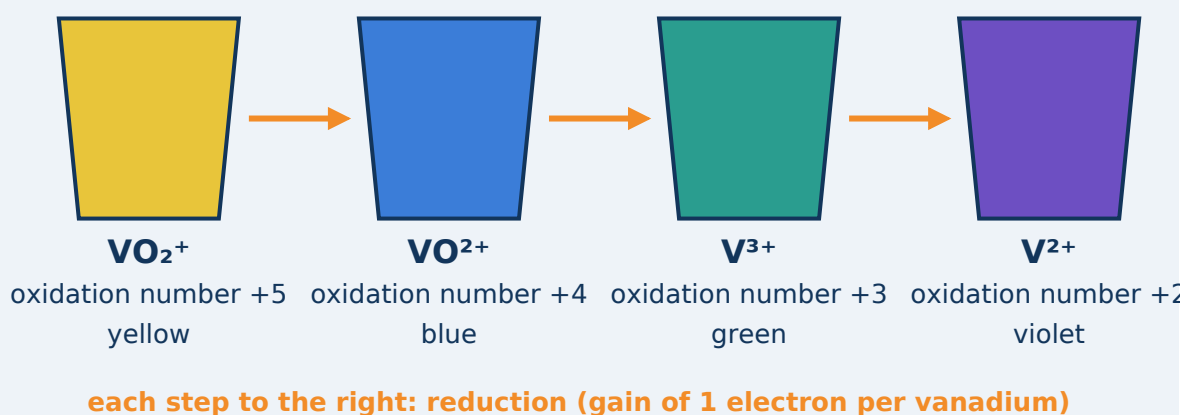
Worked example 1

Find the oxidation number of sulfur in S₂O₃²⁻ and of manganese in MnO₄⁻.

S₂O₃²⁻: 2x + 3(-2) = -2, so 2x = +4 and **S = +2**.

MnO₄⁻: x + 4(-2) = -1, so **Mn = +7**.

Vanadium species and their colours as zinc reduces them



Colour is a handy sign of oxidation number in transition metal ions.

Transition metals show their oxidation number through colour. Vanadium passes from yellow (+5) through blue (+4) and green (+3) to violet (+2) as it is reduced step by step.

How do you balance half-equations?

A half-equation shows only the oxidation or only the reduction. To write one, follow the steps in the figure, using H₂O and H⁺ in acid and OH⁻ in alkali. The final check is that atoms and charge are the same on both sides.

1 Balance atoms other than O and H
 $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$

2 Balance O by adding H_2O
 $\text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$

3 Balance H by adding H^+
 $\text{MnO}_4^- + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$

4 Balance charge by adding e^-
 $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$

5 In alkali: add OH^- to cancel every H^+
 $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$ on both sides

Method for balancing a half-equation, shown for permanganate in acid.

Method for overall equations

Multiply the half-equations so that electrons lost = electrons gained, add them, then cancel the electrons and any H^+ or H_2O that appears on both sides.

Worked example 2

Combine the half-equations of manganate(VII) and iron(II) in acid.

Reduction: $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$

Oxidation: $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$ (multiply by 5)

Overall: **$\text{MnO}_4^- + 8\text{H}^+ + 5\text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} + 5\text{Fe}^{3+}$**

Check: charge left $-1 + 8 + 10 = +17$ and right $+2 + 15 = +17$.

What is disproportionation?

Sometimes one substance is both oxidised and reduced. This is disproportionation. When chlorine is bubbled into cold sodium hydroxide, its oxidation number changes from 0 to -1 in NaCl and to $+1$ in NaClO : $\text{Cl}_2 + 2\text{NaOH} \rightarrow \text{NaCl} + \text{NaClO} + \text{H}_2\text{O}$.

How do you use redox in a titration?

The mole ratio comes from the overall equation, so always balance first. Potassium manganate(VII) is purple and its own indicator, because the first permanent pink tinge marks the end point.

Worked example 3

A 25.0 cm^3 sample of iron(II) solution needs 22.40 cm^3 of $0.0200 \text{ mol/dm}^3 \text{ MnO}_4^-$.

$n(\text{MnO}_4^-) = 0.02240 \text{ dm}^3 \times 0.0200 \text{ mol/dm}^3 = 4.48 \times 10^{-4} \text{ mol}$.

Ratio $\text{MnO}_4^- : \text{Fe}^{2+} = 1 : 5$, so $n(\text{Fe}^{2+}) = 2.24 \times 10^{-3} \text{ mol}$.

$c(\text{Fe}^{2+}) = 2.24 \times 10^{-3} \div 0.0250 \text{ dm}^3 = \mathbf{0.0896 \text{ mol/dm}^3}$.

Oxidising agent	Reduced form	Electrons gained
MnO_4^- (acid)	Mn^{2+}	5
$\text{Cr}_2\text{O}_7^{2-}$ (acid)	2Cr^{3+}	6
Cl_2	2Cl^-	2
H_2O_2 (acid)	$2\text{H}_2\text{O}$	2

Common mistake

Do not forget that oxidation numbers are written with the sign first (+2, −1) while ionic charges put the sign last (2+, 1−). Also, an element's oxidation number in a compound is not always its group charge: oxygen is −1 in H_2O_2 , not −2.

Lina's tip

After balancing, count charge on both sides before you finish. If the charges differ, an electron or an H^+ is missing. Charge is the quickest error detector.

Key points to remember

- OIL RIG: oxidation is loss, reduction is gain of electrons.
- Oxidation number of an element is 0; in a neutral compound the numbers add to 0, in an ion to its charge.
- Oxidising agents are reduced; reducing agents are oxidised.
- Balance half-equations in order: atoms other than O and H, O with H_2O , H with H^+ , charge with e^- .
- $\text{MnO}_4^- + 8\text{H}^+ + 5\text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} + 5\text{Fe}^{3+}$: mole ratio $\text{MnO}_4^- : \text{Fe}^{2+}$ is 1 : 5.

Oxidation numbers and balancing redox equations: frequently asked questions

How do you identify a redox reaction?

Assign oxidation numbers to every element on both sides. If any element changes its oxidation number, the reaction is redox. If none changes, as in precipitation and neutralisation, it is not.

What is the difference between an oxidising agent and a reducing agent?

An oxidising agent removes electrons from another substance and is reduced. A reducing agent supplies electrons and is oxidised. Look at which reactant's oxidation number goes down: that one is the oxidising agent.

How do you balance a redox half-equation in acid?

Balance the main atom, add H_2O to balance oxygen, add H^+ to balance hydrogen, then add electrons to balance charge. Finally check that atoms and charges match on both sides.

What is disproportionation?

It is a reaction in which the same element is both oxidised and reduced. Chlorine in cold alkali goes from 0 to -1 and $+1$, forming chloride and chlorate(I) ions.

Why is sulfuric acid used in manganate(VII) titrations?

The acid supplies the H^+ needed in the half-equation. Hydrochloric acid is avoided because chloride ions are oxidised by permanganate, and nitric acid is avoided because it is itself an oxidising agent.