

Electrochemistry

A-Level Chemistry 9701 · Topic 6 revision notes

Everything from the 2 lessons of this topic on a few pages: the key idea of each lesson, the diagram to remember, the points to learn and the traps examiners set. Read it, cover it, test yourself.

1. Oxidation numbers

2. Redox equations and half-equations

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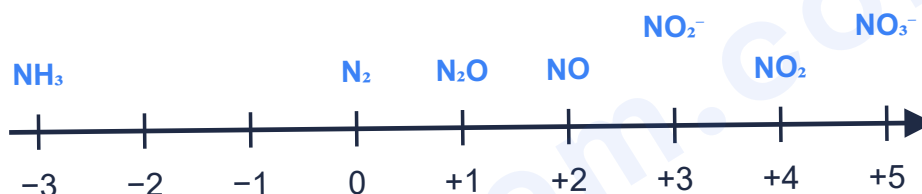
1 Oxidation numbers

Syllabus 6.1

An **oxidation number** is the charge an atom would have if all its bonds were ionic. Elements are 0; a simple ion has its charge; the numbers in a compound add up to 0, and in an ion to the ion's charge. Oxygen is usually -2 , hydrogen $+1$ and fluorine always -1 .

Oxidation is loss of electrons, an **increase** in oxidation number; **reduction** is gain of electrons, a **decrease**. An oxidising agent is itself reduced; a reducing agent is itself oxidised.

oxidation numbers of nitrogen



oxidation: oxidation number goes up $\rightarrow$

$\leftarrow$ reduction: oxidation number goes down

Fig. 1: nitrogen can take oxidation numbers from -3 to $+5$

Sum to zero, or to the charge.

- ✓ **Fixed values first:** F -1 , O -2 , H $+1$, Group 1 $+1$, Group 2 $+2$.
- ✓ **Up** = oxidised (lost e^-); **down** = reduced (gained e^-).
- ✓ **Oxidising agent** is reduced; **reducing agent** is oxidised.
- ✓ **Disproportionation:** one element goes both up and down.

Watch out!

The oxidising agent is the species that **is reduced**, because it takes electrons from something else. Students often name the species that was oxidised instead. Also remember that oxidation numbers are written with the sign first ($+3$), while ionic charges are written with the sign last ($3+$).

2 Redox equations and half-equations

Syllabus 6.1

A **half-equation** shows one species gaining or losing electrons. In acidic solution balance it in four steps: the element changing oxidation number, then **O with H_2O** , then **H with H^+** , then **charge with electrons**. The number of electrons equals the change in oxidation number. To combine two half-equations, multiply them so the electrons lost equal the electrons gained, add them, and cancel anything that appears on both sides. No electrons may remain in the final equation.

1 atoms



2 oxygen



3 hydrogen



4 charge



left charge +7 needs 5 electrons to match +2 on the right

Fig. 1: balancing the manganate(VII) half-equation in four steps

Atoms, O, H, then charge.

- ✓ **O** with H_2O , **H** with H^+ , **charge** with e^- .
- ✓ **Electrons** = change in oxidation number.
- ✓ **Match** electrons lost and gained, then add.
- ✓ **Check** atoms and charge; no e^- left over.

Watch out!

Electrons go on the **left** in a reduction and on the **right** in an oxidation, and they must all cancel in the full equation. A final equation that still shows e^- has not been combined properly. Always finish by checking that the charges balance as well as the atoms.