

# CAIE Chemistry A-Level

## 6: Electrochemistry

### Notes

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## Redox

A redox reaction is a reaction in which **oxidation and reduction** takes place. Oxidation is the loss of electrons, or increase in oxidation number. Reduction is the gain of electrons, or decrease in oxidation number.

**Disproportionation** is when a substance is both reduced and oxidised simultaneously to give two different products. For example:



## Oxidation Numbers

Oxidation numbers are used to show what is being oxidised and reduced in a redox reaction. Below are some rules to follow when assigning oxidation states:

- **Uncombined elements** always have an oxidation state of **0** (this is still true when the element has a molecular structure like  $\text{O}_2$  or a giant structure like carbon).
- In **neutral compounds**, the **sum of the oxidation states** of all the atoms is **0**.
- In an ion, the sum of the oxidation states of all the atoms is **equal to the charge** of the ion.
- **More electronegative** elements in a substance have a **negative** oxidation state while **less electronegative** elements are given a **positive** oxidation state.
- The elements in the table below nearly always have the same oxidation state.

Element	Oxidation number
Group 1 metals	+1
Group 2 metals	+2
Oxygen	-2 (usually)
Hydrogen	+1 (usually)
Fluorine	-1
Hydrogen in metal hydride	-1
Oxygen with fluorine	+2
Oxygen in peroxides	-1



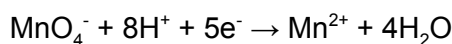
## Balancing Equations with Oxidation States

If an **oxidation state** increases by one unit, one electron is lost from that substance. If an oxidation state decreases by one unit, one electron has been gained. In a reaction, if the oxidation state of one substance decreases, this must be balanced by an increase in the oxidation state of something else.

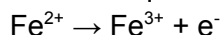
### Example:

*A solution of potassium manganate(VII),  $\text{KMnO}_4$ , acidified with dilute sulfuric acid, reacts with iron(II) ions to form iron(III) ions. The manganate(VII) ions are reduced to manganese(II) ions.*

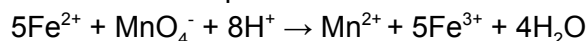
1. The oxidation state of manganese in the manganate(VII) ion is +7 and in the manganese(II) ion, it's +2. The oxidation state decreases by 5 units. The balanced equation is:



2. The oxidation state of iron in the iron(II) ions is +2 and in the iron(III) ions, it's +3. This is an increase of 1 unit. The balanced equation is:



3. For the oxidation numbers to balance, there must be  $5\text{Fe}^{2+}$  ions reacting with each  $\text{MnO}_4^-$  ion. Therefore the balanced equation can be written as:



## Oxidising and Reducing Agents

An oxidising agent gains electrons in order to oxidize another species, an **electron acceptor**.

A reducing agent loses electrons in order to reduce another species, an **electron donor**.

