CAIE Chemistry A-Level

6: Electrochemistry

Notes

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:

$$2H_2O_2 \longrightarrow 2H_2O + O_2$$

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 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), $KMnO_4$, 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:
$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$
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:
$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$
3. For the oxidation numbers to balance, there must be $5Fe^{2+}$ ions reacting with each MnO_4^- ion. Therefore the balanced equation can be written as:
$$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$$

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.**