

AQA Chemistry A-level

3.1.7: Oxidation, Reduction and Redox Detailed Notes

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3.1.7.1 - Oxidation and Reduction

Oxidation is loss of electrons. Reduction is gain of electrons.

Oxidation and reduction occur **simultaneously** in a reaction because one species loses electrons which are then donated and gained by the other species. Therefore they are known as **redox** reactions (**re**duction - **ox**idation).

This redox rule is remembered using the acronym **OILRIG** (oxidation is loss, reduction is gain).

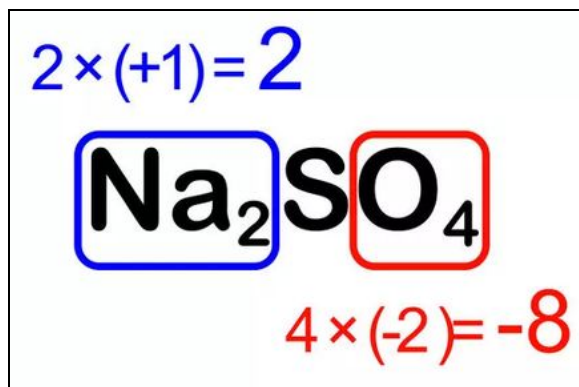
Oxidation Number

Oxidation number gives the **oxidation state** of an element or ionic substance. Allocation of oxidation state to a species follows a number of rules:

- Oxidation state of an **element is zero**.
- Oxidation states in a **neutral** compound add up to **zero**.
- Oxidation states in a charged compound add up to **total the charge**.
- **Hydrogen** has an oxidation state of **+1**.
- **Oxygen** has an oxidation state of **-2**.
- All **halogens** have an oxidation state of **-1**.
- **Group I** metals have an oxidation state of **+1**.

These rules can be used to work out the oxidation number of species or elements in a reaction.

Example:



This compound must total zero, therefore using the rules above, the oxidation number of Sulfur can be found.

$$2 - 8 + x = 0$$

$$-6 + x = 0$$

$$x = 6$$





Oxidising and Reducing Agents

An oxidising agent **accepts electrons** from the species that is being oxidised. Therefore it **gains electrons and is reduced**. This is seen as a **reduction** in oxidation number (gets more negative).

A reducing agent **donates electrons** to the species being reduced. Therefore it **loses electrons and is oxidised**. This is seen as an **increase** in oxidation number (gets more positive).

Half Equations

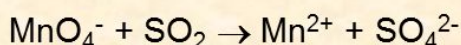
Half equations are used to show the **separate oxidation and reduction reactions** that occur in a redox reaction. They must be balanced in terms of the **species present and the charges** of the species on both sides of the equation.

In order to help write the equations, there is a useful method:

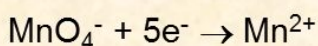
1. Balance all species excluding oxygen and hydrogen.
2. Balance oxygen using H_2O .
3. Balance hydrogen using H^+ ions.
4. Balance charges using e^- (electrons).

Following this method ensures the half equations are **correctly balanced**.

Example:



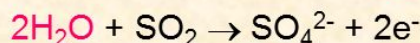
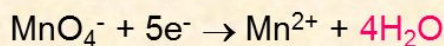
Step 2: Balance each kind of atom other than H and O



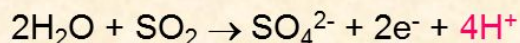
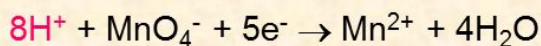
Balanced in this case



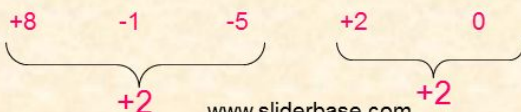
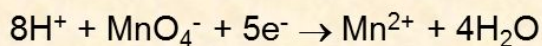
Step 3: Balance O atoms by using H_2O



Step 4: Balance H atoms by using H^+ ions



Step 5: Use electrons as needed to obtain a charge that is balanced



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Already balanced!

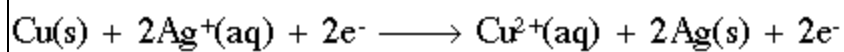
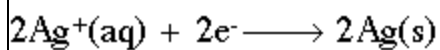
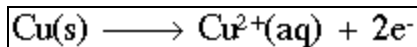




Half equations can be **combined** in order to determine the **overall redox reaction**. In order to do this, the number of **electrons must be the same** for both half equations. This can be done by scaling up the number of moles.

Example:

Image courtesy of Shodor



or

