AQA Chemistry A-Level

3.1.7: Oxidation, Reduction and Redox

Detailed Notes

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

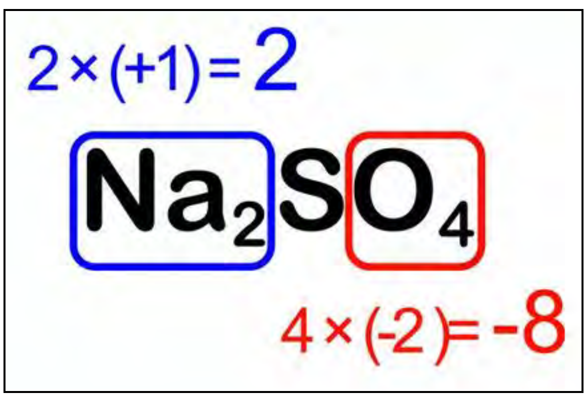
Oxidation Number

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

- Oxidation number of an **element is zero**.
- Oxidation numbers in a **neutral** compound add up to **zero**.
- Oxidation numbers in a charged compound add up to **total the charge**.
- **Hydrogen** has an oxidation number of **+1**.
- **Oxygen** has an oxidation number of **-2**.
- All **halogens** have an oxidation number of **-1**.
- **Group I** metals have an oxidation number 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 **gain electrons and is reduced**. This is seen as an **increase** in oxidation number (gets more positive).

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

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 changes using e^- (electrons).

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

$$MnO_4^- + SO_2 \rightarrow Mn^{2+} + SO_4^{2-}$$

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

$$MnO_4^- + 5e^- \rightarrow Mn^{2+}$$

$$SO_2 \rightarrow SO_4^{2-} + 2e^-$$

Balanced in this case

Step 3: Balance O atoms by using H_2O

$$MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

$$2H_2O + SO_2 \rightarrow SO_4^{2-} + 2e^-$$

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

$$8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

$$2H_2O + SO_2 \rightarrow SO_4^{2-} + 2e^- + 4H^+$$

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

$$8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

$$\underbrace{+8 \quad -1 \quad -5}_{+2} \qquad \underbrace{+2 \quad 0}_{+2}$$

www.sliderbase.com

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

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$$

$$2Ag^+(aq) + 2e^- \longrightarrow 2Ag(s)$$

$$Cu(s) + 2Ag^+(aq) + 2e^- \longrightarrow Cu^{2+}(aq) + 2Ag(s) + 2e^-$$

or

$$Cu(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$$