

WJEC Chemistry A-level

3.2: Redox Reactions

Detailed Notes

Welsh Specification

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Redox Titrations

Redox reactions can be used in **titrimetric analysis** to analyse reactions both theoretically and when carried out in practice. These redox titrations depend on the **transfer of electrons** between the two reacting species in solution. **Half equations** can be written for these transfers to work out the **ratios** between the reacting ions.

These calculations can be used before a reaction to work out the **theoretical amount** of reactant required in a reaction. This can then be compared to the actual amount required from a practical experiment.

Constructing half equations

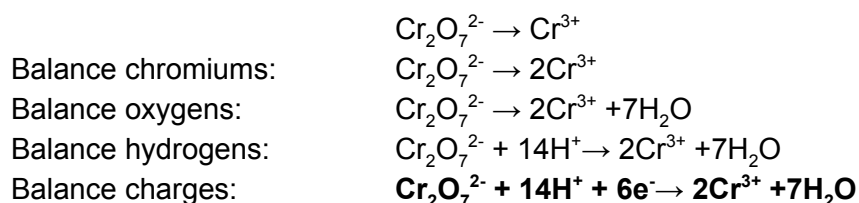
Half equations are used to show the **separate oxidation and reduction** reactions that occur in a redox reaction. They must be **balanced** in terms of the species present and the charges of the species on both sides of the equation.

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: Consider the reduction of $\text{Cr}_2\text{O}_7^{2-}$ to Cr^{3+} :



Potassium Dichromate(VI)

This compound is an **oxidising agent** used commonly in the **oxidation of alcohols**. In this process, $\text{Cr}_2\text{O}_7^{2-}$ is **reduced** (gains electrons) from to Cr^{3+} . A half equation for this reduction was deduced above:





Potassium Manganate(VII)

This compound is also an **oxidising agent**. MnO_4^- gains electrons and is **reduced** to Mn^{2+} ions. A half equation for this reduction reaction can be written as:



Thiosulfate

This compound is a **reducing agent**. $\text{S}_2\text{O}_3^{2-}$ donates electrons to become **oxidised** to $\text{S}_4\text{O}_6^{2-}$ ions. A half equation for this oxidation can be written as:

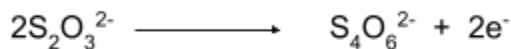
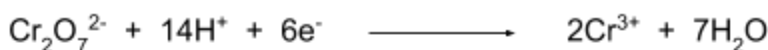


Combining Half Equations

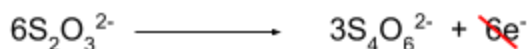
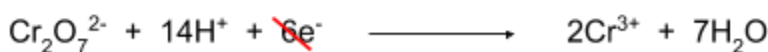
Half equations can be **combined** in order to determine the overall redox reaction. In order to do this, the number of electrons must be the **same** for both half equations. This can be done by scaling up the number of moles.

The **molar ratio** is crucial for redox titration calculations and different combinations of these half equations produce different molar ratios.

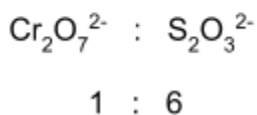
Example:



Balance the electrons and cancel:



Calculate the reacting ratio:





Cu²⁺ and I⁻ Redox Reaction

In this reaction, I⁻ isn't a strong enough reducing agent to completely reduce the Cu²⁺ ions, so they are only reduced to **Cu⁺ ions**. Therefore, the reduction reaction is a bit different.

Example:



The amount of iodine produced can be determined by titration with **sodium thiosulfate** solution of **known concentration**, since the following redox reaction takes place:



Clearly the **reacting molar ratio** of thiosulfate to iodine is 2:1 so if you **calculate the amount of sodium thiosulfate** required to react with all the iodine, then you can **calculate the amount of iodine** which was produced in the first reaction.

Redox Titrations

Redox **titrations** are carried out using a very similar method to acid-base titration where the concentration of an unknown substance can be **accurately determined** by measuring it against a **standardised titrant**.

A common example is the reaction between a standard solution of **potassium permanganate** (KMnO₄) and a solution containing an unknown concentration of **Fe²⁺ ions**.

When at the neutralisation point, the solution of KMnO₄ will turn from **bright purple** to almost **colourless** meaning there is a very clear endpoint to the titration.

Concordant results from redox titrations can then be used in redox calculations for the substances involved. Titre values are said to be **concordant** if they are within **0.20 cm³** of each other.

