WJEC Chemistry A-level

3.2: Redox Reactions

Detailed Notes

Welsh Specification

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 $Cr_2O_7^{2-}$ to Cr^{3+}:

	$Cr_2O_7^{2-} \rightarrow Cr^{3+}$
Balance chromiums:	$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$
Balance oxygens:	$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$
Balance hydrogens:	$Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$
Balance charges:	$\mathbf{Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O}$

Potassium Dichromate(VI)

This compound is an **oxidising agent** used commonly in the **oxidation of alcohols**. In this process, $Cr_2O_7^{2-}$ is **reduced** (gains elections) from to Cr^{3+}. A half equation for this reduction was deduced above:

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O$$

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:

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

Thiosulfate

This compound is a **reducing agent**. $S_2O_3^{2-}$ donates electrons to become **oxidised** to $S_4O_6^{2-}$ ions. A half equation for this oxidation can be written as:

$$\boxed{2S_2O_3^{2-} \longrightarrow S_4O_6^{2-} + 2e^-}$$

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:

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O$$

$$2S_2O_3^{2-} \longrightarrow S_4O_6^{2-} + 2e^-$$

Balance the electrons and cancel:

$$Cr_2O_7^{2-} + 14H^+ + \cancel{6e^-} \longrightarrow 2Cr^{3+} + 7H_2O$$

$$6S_2O_3^{2-} \longrightarrow 3S_4O_6^{2-} + \cancel{6e^-}$$

$$\Rightarrow Cr_2O_7^{2-} + 14H^+ + 6S_2O_3^{2-} \longrightarrow 2Cr^{3+} + 3S_4O_6^{2-} + 7H_2O$$

Calculate the reacting ratio:

$Cr_2O_7^{2-}$	:	$S_2O_3^{2-}$
1	:	6

Cu^{2+} and I^- Redox Reaction

In this reaction, **I^-** isn't a strong enough reducing agent to completely reduce the Cu^{2+} ions, so they are only reduced to **Cu^+ ions**. Therefore, the reduction reaction is a bit different.

Example:

$$2Cu^{2+} + 4I^- \longrightarrow 2CuI + I_2$$

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

$$2S_2O_3^{2-} + I_2 \longrightarrow 2I^- + S_4O_6^{2-}$$

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** ($KMnO_4$) and a solution containing an unknown concentration of **Fe^{2+} ions**.
When at the neutralisation point, the solution of $KMnO_4$ 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 cm^3** of each other.