Edexcel Chemistry A-Level

Topic 3: Redox I

Detailed Notes
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:

\[
\begin{align*}
2 \times (+1) &= 2 \\
\text{Na}_2\text{SO}_4 \\
4 \times (-2) &= -8
\end{align*}
\]

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

\[
\begin{align*}
2 - 8 + x &= 0 \\
-6 + x &= 0 \\
x &= 6
\end{align*}
\]

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 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).

**Redox Equations**

**Disproportionation Reactions**

In a disproportionation reaction, a species is both oxidised and reduced, seen as both an increase and a decrease in oxidation number for that species.

An example is seen when chlorine reacts with cold water to produce Chlorate(I) ions (ClO\(^-\)) and chloride ions. The oxidation state goes from zero to both +1 and -1.  
*Example:*

\[
\text{Cl}_2 + \text{H}_2\text{O} \rightarrow \text{ClO}^- + \text{Cl}^- + 2\text{H}^+ 
\]

**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\(_2\)O.
4. Balance changes using e\(^-\) (electrons).

Following this method ensures the half equations are correctly balanced.

*Example:*
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:**

\[
\begin{align*}
\text{MnO}_4^- + \text{SO}_2 &\rightarrow \text{Mn}^{2+} + \text{SO}_4^{2-} \\
\text{Step 2: Balance each kind of atom other than H and O} &\\
\text{MnO}_4^- + 5\text{e}^- &\rightarrow \text{Mn}^{2+} & \text{Balanced in this case} \\
\text{SO}_2 &\rightarrow \text{SO}_4^{2-} + 2\text{e}^- \\
\text{Step 3: Balance O atoms by using H}_2\text{O} &\\
\text{MnO}_4^- + 5\text{e}^- &\rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} \\
2\text{H}_2\text{O} + \text{SO}_2 &\rightarrow \text{SO}_4^{2-} + 2\text{e}^- \\
\text{Step 4: Balance H atoms by using H}^+ \text{ ions} &\\
8\text{H}^+ + \text{MnO}_4^- + 5\text{e}^- &\rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} \\
2\text{H}_2\text{O} + \text{SO}_2 &\rightarrow \text{SO}_4^{2-} + 2\text{e}^- + 4\text{H}^+ \\
\text{Step 5: Use electrons as needed to obtain a charge that is balanced} &\\
8\text{H}^+ + \text{MnO}_4^- + 5\text{e}^- &\rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} \\
\begin{array}{ccc}
+8 & -1 & -5 \\
+2 & 0 & +2
\end{array}
\text{Already balanced!}
\end{align*}
\]

Cu(s) \rightarrow 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)