REDOX Reactions

A REDOX reaction involves two half reactions - oxidation and reduction.

These half reactions can be written as ion-electron equations. Ion-electron equations are found on page 11 of the Data Booklet.

Oxidation involves the LOSS of electrons (OIL):

\[ \text{Fe} \rightarrow \text{Fe}^{2+} + 2e \quad \text{Mg} \rightarrow \text{Mg}^{2+} + 2e \]

The ion-electron equation for oxidation must be written in reverse.

Reduction involves the GAIN of electrons (RIG):

\[ \text{Cu}^{2+} + 2e \rightarrow \text{Cu} \quad \text{Ag}^+ + e \rightarrow \text{Ag} \]

Remember, oxidation and reduction are two halves of the same chemical reaction. The combined reaction is called a Redox reaction.

To form the overall redox reaction, the ion-electron equations for the oxidation and reduction must be combined, ensuring that the electrons cancel.

\[
\begin{align*}
\text{Oxidation:} & \quad \text{Mg}(s) \rightarrow \text{Mg}^{2+} + 2e \\
\text{Reduction:} & \quad \text{Ag}^+ + e \rightarrow \text{Ag}(s) \quad (x2) \\
\hline
\text{Mg}(s) + 2\text{Ag}^+ & \rightarrow \text{Mg}^{2+} + 2\text{Ag}(s) \\
\end{align*}
\]

\[
\begin{align*}
\text{Oxidation:} & \quad \text{Al}(s) \rightarrow \text{Al}^{3+} + 3e \quad (x2) \\
\text{Reduction:} & \quad 2\text{H}^+ + 2e \rightarrow \text{H}_2(g) \quad (x3) \\
\hline
\text{Mg}(s) + 2\text{Ag}^+ & \rightarrow \text{Mg}^{2+} + 2\text{Ag}(s) \\
\end{align*}
\]
Oxidising and Reducing Agents

In a redox reaction the species that is oxidised is described as a **reducing agent** - a species that allows reduction to occur.

Similarly, a species which is reduced is described as an **oxidising agent** - a species that allows oxidation to occur.

eg. $$\text{Mg}_\text{(s)} + 2\text{Ag}^+ \rightarrow \text{Mg}^{2+} + 2\text{Ag}_\text{(s)}$$

**Reducing agent**  **Oxidising agent**

Using the Data Booklet

When writing redox equations you must first identify the oxidation and reduction half reactions and then combine them to give the redox reaction. Use the Electrochemical Series on page 12 of the data booklet.

**Example:**

Iron(II)sulphate solution is reacted with potassium permanganate solution until the first appearance of a permanent pink colour.

**Step 1:** What four ions are present in the two solutions?

**Step 2:** Write the ion-electron equation for the oxidation reaction.

**Step 3:** Write the ion-electron equation for the reduction reaction.

**Step 4:** Add the two half equations together to give the redox equation (remember to balance the electrons first)
To find the species being oxidised in the solution start at the top right and work your way down until you find the first atom or ion being oxidised from the solution. Highlight it when you find it.

To find the species being reduced in the solution start at the bottom left and work your way up until you find the first atom or ion from the solution. Highlight it when you find it.
More Complicated Half-Equations

There are some ion-electron equations that are not given in the data book. You must learn the rules to work them out for yourself.

**Worked example**

Write an ion-electron equation for the following reaction:

\[
\text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}^{3+}
\]

**Step 1:** If necessary, balance the central atom / ion.

\[
\text{Cr}_2\text{O}_7^{2-} \rightarrow 2\text{Cr}^{3+}
\]

**Step 2:** Add water (H\(_2\)O) if it is needed to balance the oxygen atoms.

\[
\text{Cr}_2\text{O}_7^{2-} \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}
\]

**Step 3:** Balance the hydrogen in the water by adding hydrogen ions.

\[
\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}
\]

**Step 4:** Calculate the total electrical charge on each side of the equation.

\[
\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}
\]

\[
(12+) \quad (6+)
\]

**Step 5:** Add electrons to balance the electrical charges.

\[
\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}
\]

\[
(12+) \quad + \quad (6-) \quad \rightarrow \quad (6+)
\]