What are redox reactions (A level Chemistry)

A **RedOx** reaction is simply a reaction in which **Red**uction and **Ox**idation take place at the same time.

Oxidation in simple terms

We can define **oxidation** as the gain of oxygen by a substance. For example, when magnesium reacts with oxygen, the magnesium combines with oxygen to form magnesium oxide. This actually means that magnesium has been oxidised by oxygen to produce magnesium oxide.

The chemical equation for this reaction is:

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Reduction is simple terms

We can define **reduction** as the loss of oxygen by a substance. For example, when copper(II) oxide reacts with hydrogen, it loses oxygen to form copper.

The reaction is as follows:

$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$

Copper(II) oxide is reduced by hydrogen to form copper. However if we look carefully at the reaction we can see that whilst copper(II) oxide is being reduced to copper, hydrogen is also oxidised to water. Therefore reduction and oxidation have taken place together. Oxidation and reduction always take place together and this type of reaction is called a redox reaction.

Question

In each of the following equations, state which substance has been oxidised and which substance has been reduced.

i. $PbO + H_2 \rightarrow Pb + H_2O$
ii. $CO + Ag_2O \rightarrow 2Ag + CO_2$

iii. $2Mg + CO_2 \rightarrow 2MgO + C$

Redox in terms of electron transfer

In terms of electron transfer:

- Oxidation is the loss of electrons.
- Reduction is the gain of electrons.

For example:

Sodium reacts with chlorine to form sodium chloride as follows:

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

To see which substance has been oxidised or reduced we first convert this reaction into ionic an ionic equation.

$2Na(s) + Cl_2(g) \rightarrow 2Na^+(s)\ 2Cl^-(s)$

Then split the ionic equation into half-equations:

$2Na(s)$ becomes $2Na^+$ by the loss of electrons. (Oxidation)

Cl_2 becomes $2Cl^-$ by the gain of electrons. (Reduction)

Therefore, we have two redox half-equations becomes:

The half-equation shows that sodium is oxidised.:

$2Na(s) - 2e \rightarrow 2Na^+(s)$

This half-equation shows that sodium is oxidised.

$Cl_2(g) + 2e \rightarrow 2Cl^-(s)$

Example

Write two half-equations for the following reactions. For each half-equation state whether oxidation or reduction is occurring:

a. $2Fe + 3Cl_2 \rightarrow 2FeCl_3$
b. $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$

Solution a

$2Fe + 3Cl_2 \rightarrow 2FeCl_3$

First split the $2FeCl_3$ into ions. Leave 2Fe and $3Cl_2$ as they are, because they are free elements:

$2Fe + 3Cl_2 \rightarrow$ **$2Fe^{3+}$ + $6Cl^-$**

the charge on the iron ion is +3 because 1 iron atom is bonded to 3 chlorine atoms. Each chlorine atom requires 1 electron to form a stable configuration. Three chlorine atoms require 3 electrons from iron when bonding with iron, leaving the iron as Fe^{3+}.

Now separate the equation into 2 equations based on the products and reactants.

$2Fe \rightarrow 2Fe^{3+}$

and

$3Cl_2 \rightarrow 6Cl^-$

Reduce the first equation to:

$Fe \rightarrow Fe^{3+}$

and balance it using electrons. A charge of +3 means a loss of 3 electrons:

- **$Fe - 3e \rightarrow Fe^{3+}$**
- Oxidation has occurred.

Reduce the second equation to:

Reduce the second equation to:

$Cl_2 \rightarrow 2Cl^-$

and balance it using electrons.

- **$Cl_2 + 2e \rightarrow 2Cl^-$**
- Reduction has occurred.

Solution b

$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$

$Fe + Cu^{2+} + SO_4^{2-} \rightarrow Fe^{2+} + SO_4^{2-} + Cu$

Remove the spectator ions:

$Fe + Cu^{2+} + \mathbf{SO_4^{2-}} \rightarrow Fe^{2+} + \mathbf{SO_4^{2-}} + Cu$

$Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$

Split into half-equations and balance them using electrons.

- **$Fe - 2e \rightarrow Fe^{2+}$**
- Oxidation has occurred.

And

- **$Cu^{2+} + 2e \rightarrow Cu$**
- Reduction has occurred.

Given two half-equations it is also possible to construct a balanced ionic equation by balancing the numbers of electrons lost and gained and then adding the two half-equations together.

Example

Construct the balanced ionic equation for the reaction between nickel and iron(III) ions, Fe^{3+}, using the following half-equations:

- $Ni(s) \rightarrow Ni^{2+}(aq) + 2e$
- $Fe^{3+}(aq) + e \rightarrow Fe^{2+}(aq)$

Solution

From the half-equations we can see that:

- *each Ni atom loses two electrons when it is oxidised to $Ni^{2+}(aq)$*
- *and each Fe^{3+} ion gains one electron when it is reduced to Fe^{2+}.*

So to produce a balanced ionic equation we first have to make sure that the electrons lost is equal to the electrons gained. This means that the two Fe^{3+} ions need to gain the two electrons lost when each Ni^{2+} ion is formed.

Balance the electrons gained and the electrons lost.

$2Fe^{3+}(aq) + 2e \rightarrow 2Fe^{2+}(aq)$

$Ni(s) \rightarrow Ni^{2+}(aq) + 2e$

Add the 2 half-equations together:

$2Fe^{3+}(aq) + 2e + Ni(s) \rightarrow 2Fe^{2+}(aq) + Ni^{2+}(aq) + 2e$

Then cancel out the electrons from both sides. The balanced ionic equation is:

$Ni(s) + 2Fe^{3+}(aq) \rightarrow Ni^{2+}(aq) + 2Fe^{2+}(aq)$

Example

Construct the balanced ionic equation for the reaction of iodide ions (I^-) with manganate(VII) ions (MnO_4^-) in the presence of hydrogen ions (H^+), using the following two half-equations:

- $2I^-(aq) \rightarrow I_2(aq) + 2e$
- $MnO_4^-(aq) + 8H^+(aq) + 5e \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

Solution

Multiply the first equation by 5 and the second equation by 2 to balance their number of electrons.

- $10I^-(aq) \rightarrow 5I_2(aq) + 10e$
- $2MnO_4^-(aq) + 16H^+(aq) + 10e \rightarrow 2Mn^{2+}(aq) + 8H_2O(l)$

Combine the 2 equations and cancel out their electrons.

$10I^-(aq) + 2MnO_4^-(aq) + 16H^+(aq) \rightarrow 5I_2(aq) + 2Mn^{2+}(aq) + 8H_2O(l)$

Example

Zinc metal reacts with IO^{3-} ions in acidic solution. Construct a balanced ionic equation for this reaction, using the two half-equations below:

- $2IO^{3-} + 12H^+ + 10e \rightarrow I_2 + 6H_2O$
- $Zn \rightarrow Zn^{2+} + 2e$

Solution

Balance the electrons on both equations by multiplying the first equation by 1 and the second equation by 5.

- $2IO^{3-} + 12H^+ + 10e \rightarrow I_2 + 6H_2O$
- $5Zn \rightarrow 5Zn^{2+} + 10e$

Combine the equations and cancel out the electrons

$2IO^{3-} + 12H^+ + 5Zn \rightarrow I_2 + 6H_2O + 5Zn^{2+}$

Balancing redox equations in acidic conditions

When a redox reactions occur in acidic conditions the following stages are followed when balancing the equation.

- divide the redox equation into two half-equations, one for oxidation and the other for reduction.
- balance all elements except for oxygen and hydrogen.
- oxygen is balanced by adding H_2O to the appropriate side of the equation and hydrogen is balanced by adding H^+ to the appropriate side of the equation.
- balance the charge by adding electrons to the appropriate side of the equation.
- combine two half-equations and cancel out electrons on both sides of the redox equation.

Example

Balance and complete the following redox equation:

$Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+}$

Solution

- Divide the equation into two equations according to the reactants and products:

$Cr_2O_7^{2-} \rightarrow Cr^{3+}$ (reduction)

$Fe^{2+} \rightarrow Fe^{3+}$ (oxidation)

- Balance all the atoms except for oxygen and hydrogen:

$Cr_2O_7^{2-} \rightarrow \mathbf{2}Cr^{3+}$

$Fe^{2+} \rightarrow Fe^{3+}$

- Balance the oxygen by adding water molecules:

$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + \mathbf{7H_2O}$

$Fe^{2+} \rightarrow Fe^{3+}$

- Balance the hydrogen by adding H^+ ions:

$Cr_2O_7^{2-} + \mathbf{14H^+} \rightarrow 2Cr^{3+} + 7H_2O$

$Fe^{2+} \rightarrow Fe^{3+}$

- Balance the charges on each equation separately by adding the corresponding number of electrons:

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

$Fe^{2+} \rightarrow Fe^{3+} + \mathbf{e}$

- Balance the charges equating the electrons gained in the first equation to the electrons lost in the second equation.

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

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

$6Fe^{2+} \rightarrow 6Fe^{3+} + 6e$

- Combine the two equations.

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

- Cancel out the electrons on both side.

$\mathbf{Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \rightarrow 2Cr^{3+} + 7H_2O + 6Fe^{3+}}$