

# Edexcel Chemistry A-level

## Topic 3: Redox I

### Detailed Notes

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## Oxidation and Reduction

**Oxidation involves the loss of electrons. Reduction involves the gain of electrons.**

This redox rule is remembered using the acronym **OILRIG** (oxidation is loss, reduction is gain).

## Oxidation Number

The oxidation number gives the **oxidation state** of an element or ionic substance. Allocation of oxidation number to a species follows a number of rules:

- The 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**.
- **Halogens** have an oxidation number of **-1**.
- **Group I** metals have an oxidation number of **+1**.
- **Group II** metals have an oxidation number of **+2**.

However, there are some **exceptions** to these rules:

- **Oxygen** has an oxidation number of **-1** in **peroxides**.
- **Hydrogen** has an oxidation number of **-1** in **metal hydrides**.

These rules can be used to work out the oxidation number of species or elements in a reaction or compound.

*Example:*

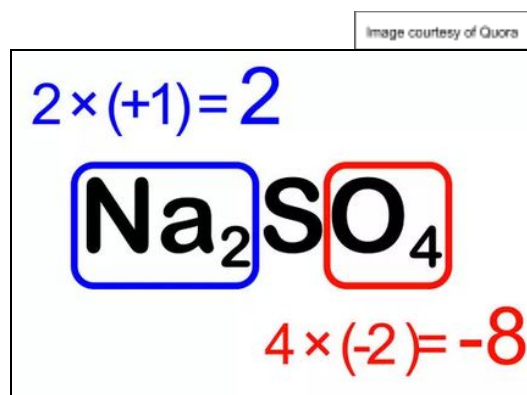
*This compound's total oxidation number is zero. Using the rules above, the oxidation number of sulfur can be found:*

Known oxidation numbers: Na=+1, O=-2.

$$2 - 8 + x = 0$$

$$-6 + x = 0$$

$$x = 6$$





Example:

What is the oxidation state of oxygen in hydrogen peroxide,  $\text{H}_2\text{O}_2$ ?

Hydrogen: +1, oxygen: -2 UNLESS in a peroxide  
 $\text{H}_2\text{O}_2$  is uncharged, therefore the sum of oxidation states must equal 0.

$$(2 \times +1) + (2 \times X) = 0$$

$$2 + 2X = 0$$

$$2X = -2$$

$$X = -1$$

Therefore, the oxidation state of oxygen in hydrogen peroxide is -1.

## Roman numerals

**Roman numerals** can be used to give the oxidation number of an element that has a **variable oxidation state**, depending on the compound it's in.

Example:

Copper(II) sulphate - this tells you the oxidation number of copper is +2

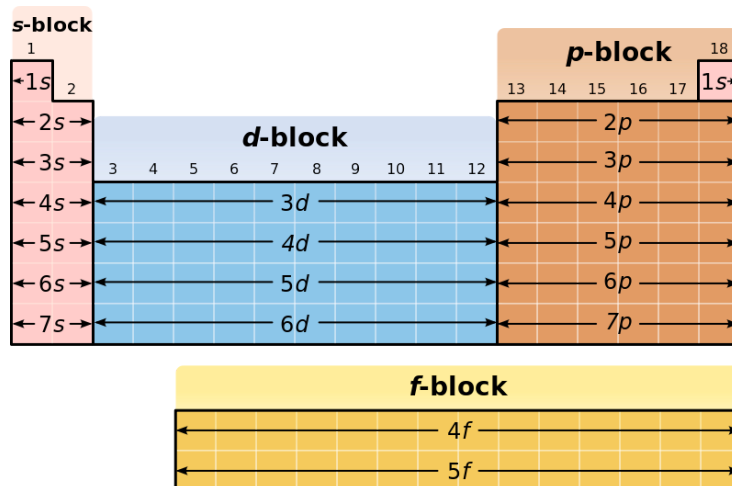
Iron(II) sulphate(VI) - this tells you the oxidation number of iron is +2 and the oxidation number of sulphur is +6

In the same way that oxidation numbers can be calculated from **formulas** of compounds, the formula of compounds may be deduced if the oxidation numbers of the elements (given by the **rules of oxidation states** and **roman numerals**) and the **overall charge** of the compound is known.

## Oxidation state and the periodic table

Electrons are held in **orbitals**. Elements are arranged in the periodic table by **proton number** and also by their orbitals. These orbitals correspond with **blocks** on the Periodic Table. Each element in the block has **outer electrons in that orbital**.





Elements within the same **block** react in similar ways since their outermost electron is in the same type of **orbital**. This leads to some **patterns** in oxidation number in the periodic table:

- **s block elements** (groups 1 and 2 metals) generally **lose electrons**, so are **oxidised** and form species with **positive oxidation numbers**.
- **p block non-metals** generally **gain electrons**, so are **reduced** and form species with **negative oxidation states**.

## 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

Reactions in which oxidation and reduction occur **simultaneously** take place when one species loses electrons, which are then donated and gained by the other species. These reactions are known as **redox** reactions (**reduction - oxidation**). Being able to work out the oxidation number of atoms in a reaction enables you to work out if a redox reaction is a **disproportionation** reaction too.

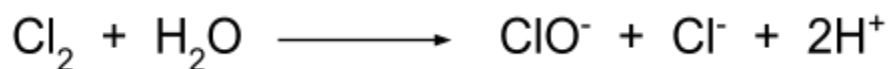




## Disproportionation Reactions

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<sup>-</sup>)** and **chloride ions**. The oxidation state goes from zero (in Cl<sub>2</sub>) to both **+1 (ClO<sup>-</sup>)** and **-1 (Cl<sup>-</sup>)**.



## Half Equations

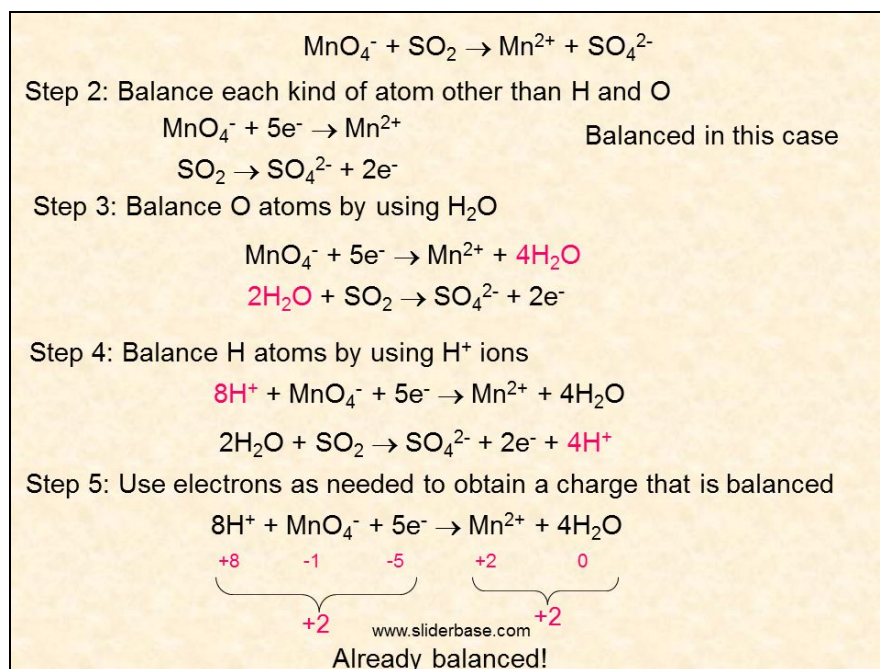
Half equations are used to show the **separate oxidation and reduction reactions** that occur in a redox reaction. They must be balanced in terms of the **species present and the charges** of the species on both sides of the equation.

In order to help write these equations, there is a useful method:

1. Balance all species excluding oxygen and hydrogen.
2. Balance oxygen using H<sub>2</sub>O.
3. Balance hydrogen using H<sup>+</sup> ions.
4. Balance charges using e<sup>-</sup> (electrons).

Following this method ensures the half equations are **correctly balanced**.

*Example:*

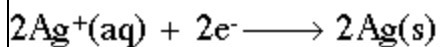
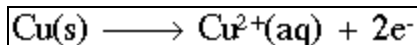




Half equations can be **combined** in order to determine the **overall redox reaction**. In order to do this, the number of **electrons must be the same** for both half equations. This can be done by scaling up the number of moles. Once the half equations are combined, the electrons should be **cancelled out** on each side of the equation.

*Example:*

Image courtesy of Shodor



*or*

