Edexcel Chemistry A-level

Topic 3: Redox I

Detailed Notes

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:

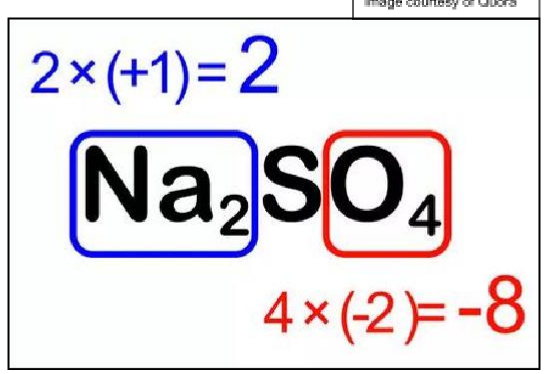
Known oxidation numbers: Na=+1, O=-2.

2 - 8 + x = 0

-6 + x = 0

X = 6

Image courtesy of Quora



Example:

What is the oxidation state of oxygen in hydrogen peroxide, H_2O_2?

Hydrogen: +1, oxygen: -2 UNLESS in a peroxide

H_2O_2 is uncharged, therefore the sum of oxidations 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.

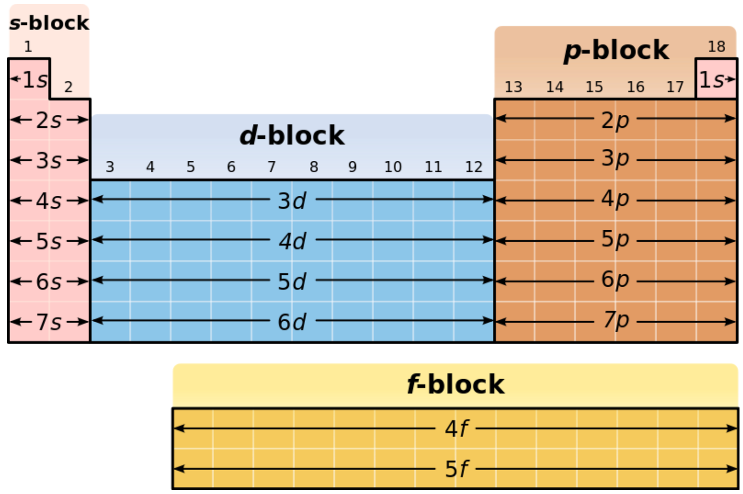
Example:

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 (**red**uction - **ox**idation). 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^-)** and **chloride ions**. The oxidation state goes from zero (in Cl_2) to both **+1 (ClO^-) and -1 (Cl^-)**.

$$\boxed{Cl_2 + H_2O \longrightarrow ClO^- + Cl^- + 2H^+}$$

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_2O.
3. Balance hydrogen using H^+ ions.
4. Balance charges using e^- (electrons).

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

Example:

$$MnO_4^- + SO_2 \rightarrow Mn^{2+} + SO_4^{2-}$$

Step 2: Balance each kind of atom other than H and O

$$MnO_4^- + 5e^- \rightarrow Mn^{2+}$$

Balanced in this case

$$SO_2 \rightarrow SO_4^{2-} + 2e^-$$

Step 3: Balance O atoms by using H_2O

$$MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

$$2H_2O + SO_2 \rightarrow SO_4^{2-} + 2e^-$$

Step 4: Balance H atoms by using H^+ ions

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

$$2H_2O + SO_2 \rightarrow SO_4^{2-} + 2e^- + 4H^+$$

Step 5: Use electrons as needed to obtain a charge that is balanced

$$\underbrace{\underset{+8}{8H^+} + \underset{-1}{MnO_4^-} + \underset{-5}{5e^-}}_{+2} \rightarrow \underbrace{\underset{+2}{Mn^{2+}} + \underset{0}{4H_2O}}_{+2}$$

Already balanced!

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

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$$

$$2Ag^+(aq) + 2e^- \longrightarrow 2Ag(s)$$

$$Cu(s) + 2Ag^+(aq) + 2e^- \longrightarrow Cu^{2+}(aq) + 2Ag(s) + 2e^-$$

or

$$Cu(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$$