**Definitions**

* ***Oxidation is loss of electrons.***
* ***Reduction is gain of electrons.***

It is essential that you remember these definitions. There is a very easy way to do this. As long as you remember that you are talking about electron transfer:



**A simple example**

The equation shows a simple redox reaction which can obviously be described in terms of oxygen transfer.



Copper(II) oxide and magnesium oxide are both ionic. The metals obviously aren't. If you rewrite this as an ionic equation, it turns out that the oxide ions are spectator ions and you are left with:



**A last comment on oxidising and reducing agents**

If you look at the equation above, the magnesium is reducing the copper(II) ions by giving them electrons to neutralise the charge. Magnesium is a reducing agent.

Looking at it the other way round, the copper(II) ions are removing electrons from the magnesium to create the magnesium ions. The copper(II) ions are acting as an oxidising agent.

***Warning!***

This is potentially very confusing if you try to learn both what oxidation and reduction mean in terms of electron transfer, and also learn definitions of oxidising and reducing agents in the same terms.

Personally, I would recommend that you work it out if you need it. The argument (going on inside your head) would go like this if you wanted to know, for example, what an oxidising agent did in terms of electrons:

* An oxidising agent oxidises something else.
* Oxidation is loss of electrons (OIL RIG).
* That means that an oxidising agent takes electrons from that other substance.
* So an oxidising agent must gain electrons.

Or you could think it out like this:

* An oxidising agent oxidises something else.
* That means that the oxidising agent must be being reduced.
* Reduction is gain of electrons (OIL RIG).
* So an oxidising agent must gain electrons.

Understanding is a lot safer than thoughtless learning!

Oxidation state shows the total number of electrons which have been removed from an element (a positive oxidation state) or added to an element (a negative oxidation state) to get to its present state.

**Oxidation involves an increase in oxidation state**

**Reduction involves a decrease in oxidation state**

Recognising this simple pattern is the single most important thing about the concept of oxidation states. If you know how the oxidation state of an element changes during a reaction, you can instantly tell whether it is being oxidised or reduced without having to work in terms of electron-half-equations and electron transfers.

**Working out oxidation states**

You *don't* work out oxidation states by counting the numbers of electrons transferred. It would take far too long. Instead you learn some simple rules, and do some very simple sums!

* The oxidation state of an uncombined element is zero. That's obviously so, because it hasn't been either oxidised or reduced yet! This applies whatever the structure of the element - whether it is, for example, Xe or Cl2 or S8, or whether it has a giant structure like carbon or silicon.
* The sum of the oxidation states of all the atoms or ions in a neutral compound is zero.
* The sum of the oxidation states of all the atoms in an ion is equal to the charge on the ion.
* The more electronegative element in a substance is given a negative oxidation state. The less electronegative one is given a positive oxidation state. Remember that fluorine is the most electronegative element with oxygen second.
* Some elements almost always have the same oxidation states in their compounds:

|  |  |  |
| --- | --- | --- |
| **element** | **usual oxidation state** | **exceptions** |
| Group 1 metals | always +1 |   |
| Group 2 metals | always +2 |   |
| Oxygen | usually -2 | except in peroxides and F2O (see below) |
| Hydrogen | usually +1 | except in metal hydrides where it is -1 (see below) |
| Fluorine | always -1 |   |
| Chlorine | usually -1 | except in compounds with O or F (see below) |

In each of the following examples, we have to decide whether the reaction involves redox, and if so what has been oxidised and what reduced.

*Example 1:*

This is the reaction between magnesium and hydrochloric acid or hydrogen chloride gas:



Have the oxidation states of anything changed? Yes they have - you have two elements which are in compounds on one side of the equation and as uncombined elements on the other. Check all the oxidation states to be sure:.



The magnesium's oxidation state has increased - it has been oxidised. The hydrogen's oxidation state has fallen - it has been reduced. The chlorine is in the same oxidation state on both sides of the equation - it hasn't been oxidised or reduced.

*Example 2:*

The reaction between sodium hydroxide and hydrochloric acid is:



Checking all the oxidation states:



Nothing has changed. This isn't a redox reaction.

*Example 3:*

This is a sneaky one! The reaction between chlorine and cold dilute sodium hydroxide solution is:



Obviously the chlorine has changed oxidation state because it has ended up in compounds starting from the original element. Checking all the oxidation states shows:



The chlorine is the *only* thing to have changed oxidation state. Has it been oxidised or reduced? Yes! Both! One atom has been reduced because its oxidation state has fallen. The other has been oxidised.

This is a good example of a ***disproportionation*** reaction. A disproportionation reaction is one in which a single substance is both oxidised and reduced.