# Oxidation States

half-equations
ionic equations

## Oxidation/reduction

As you should know, oxidation is loss of electrons and reduction is gain of electrons: OILRIG!

✓ oxidation and reduction always occur simultaneously. You cannot have one without the other.

An oxidizing agent oxidizes something i.e. removes electrons. Those electrons have to go *somewhere*, so the oxidising agent accepts them. The oxidising agent is therefore reduced (gains electrons) and vice-versa for a reducing agent.

an oxidizing agent is an **electron acceptor**

a reducing agent is an **electron donor**

✓ Oxidising agents have **high oxidation states** e.g. in $MnO_4^-$, we have $Mn^{7+}$. This means it has plenty too to accept electrons.

✓ Reducing agents have **low oxidation states** e.g. metals such as Na, Zn. They have plenty of electrons to lose.

The overall process is called **redox** (reduction and oxidation).

## Oxidation states

An oxidation state just means **how many electrons an element has gained or lost**. For example, if you look at NaCl again, then you would say the oxidation state of Na is 1+ and the Cl is 1-.

In the ionic bonding tutorial we looked at how to work out the formulae of ionic compounds.

With this topic the questions often ask you to work out the oxidation state (or oxidation number) of an atom within a compound, which is very similar to the ionic formulae stuff.

The point of this is to allow you to see what has been oxidised and what has been reduced.

This differs from working out the formulae of ionic compounds as **oxidation states can change**. It just depends what compound is being formed. So you have to look at each example individually.

### The Rules

In the ionic bonding tutorial we looked at a few basic rules for working out what charge an ion has. As with all Chemistry topics there are of course some exceptions to the rules:

**Oxygen:** is 99% of the time $O^{2-}$ but if it is a peroxide such as $H_2O_2$, then it has a 1- charge. Peroxide means it has an O-O single bond.

**Hydrogen:** is 99% of the time 1+ but if it is part of a metal hydride such as NaH, then it has a 1- charge.

**Group 1-3:** always +1 to +3

**Group 5-7**: often are 3-, 2- and 1- but they can change their oxidation state.

### Working out oxidation states

You need to apply the above rules and use a bit of common sense.

Most of the examples will have more than one element in a compound. This means you can only have **one unknown** and you have to assume that the other elements are following the above basic rules.

### **Example:** *what is the oxidation state of N in $KNO_3$?*

As N is the unknown, you have to assume that the K is 1+ and oxygen is 2- in this case. But there are 3 oxygens, so we need to do 2- x 3 → 6-.

Add up the charges: 1+ and 6- → 5-.

**The overall charge of a compound is always 0**, therefore we need to go from 5- to 0, which means N must be 5+.

### **Example:** *what is the oxidation state of O in $OF_2$?*

O is the unknown this time, so we have to assume that F is in its' usual 1- oxidation state.

Therefore we can say 1- x 2 for the two F atoms → 2-. We need to get back to 0 again, so O must be 2+.

### **Example:** *what is the oxidation number of Cl in $Cl_2O_7$?*

We have to assume that the O is 2-. Therefore we have 2- x 7 → 14-.

We need to get back to 0 again, so we need 14+, but as there are two chlorines, we can do 14/2 → 7+ per Cl atom.

## Oxidation or Reduction?

As was mentioned in the first paragraph, oxidation and reduction occur **simultaneously**. You have to be able to look at equations and identify what is being oxidised and what is being reduced.

For example:

$$\mathbf{Fe_2O_3 + 2Al \rightarrow 2Fe + Al_2O_3}$$

The easiest way to do this is simply work out the oxidation state for all elements:

**Left hand side:** $Fe^{3+}$ and $O^{2-}$ and $Al^0$

**Right hand side:** $Fe^0$ and $Al^{3+}$ and $O^{2-}$

**$Fe^{3+} \rightarrow Fe^0$**, which means Fe has **gained** 3 electrons, so it has been reduced.

**$Al^0 \rightarrow Al^{3+}$**, which means Al has **lost** 3 electrons, so it has been oxidised.

✓ when working out oxidation states, ignore the 'big' numbers e.g. the 2 in front of the Al and 2 in front of the Fe. They are just there to balance the equation.

As you improve at this topic, you might not want to work out every single oxidation state of every element as it is time consuming. Look for:

- **Elements.** If there is an element then that element must be involved with oxidation or reduction. It can't be an element on both sides of an equation.

- **C, N, S** and **halogens.** These elements commonly change their oxidation states and worth looking out for.

## Which is the oxidising/reducing agent?

What is going on in the following reaction? **$Cl_2(g) + 2Na(s) \rightarrow 2NaCl(s)$**

Cl has gone from **0 to -1** and Na has gone from **0 to +1**.

Chlorine has been **reduced** and sodium has been **oxidised**. This means that chlorine has gained an electron, which has come from the sodium.

As sodium is **giving away** an electron, sodium is therefore a reducing agent. You can reverse this argument to show that $Cl_2$ is the oxidising agent.

✓ Understanding this usually causes problems for students. If something is 'being' reduced, it itself is an oxidising agent i.e. $Cl_2$ from above. If something is being oxidised, then it itself is a reducing agent i.e. Na from above.

**Multiple Choice**

Some multiple choice questions ask things like which reaction is or isn't redox. A couple of simple things to look for:

- **acid + base → salt + water**. This is neutralisation and therefore **not redox**.

- **Look for elements** in the equation as this should indicate redox.

For example:

$$Mg(NO_3)_2 \rightarrow MgO + 2NO_2 + \tfrac{1}{2} O_2$$

Not redox as it is neutralisation →

$$HCl + NaOH \rightarrow NaCl + H_2O$$

$$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$$

$$Cl_2 + 2Br^- \rightarrow Br_2 + 2Cl^-$$

✓ Note the circled elements, which again indicates redox.

In part 2 we will look at forming half-equations with a view to constructing ionic equations, which are discussed in the final part.

half-equations
ionic equations