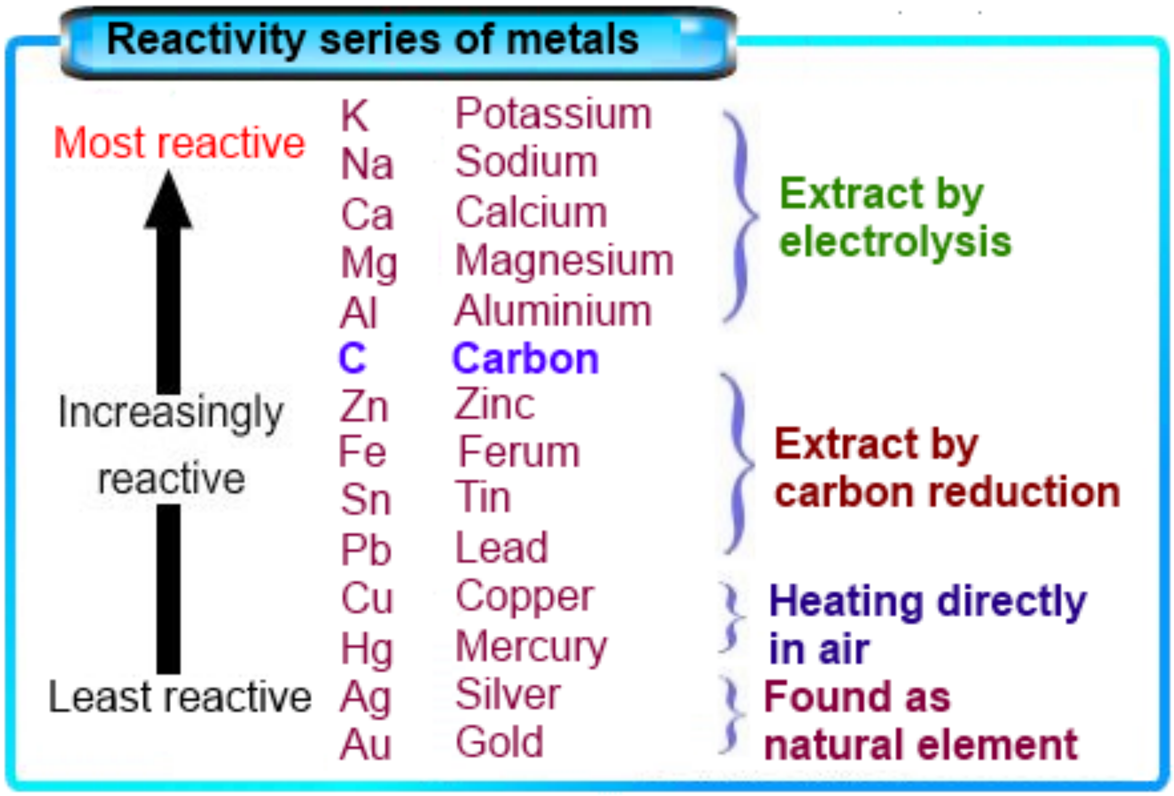
Reactivity Series



Observations of the way that these elements react with water, Acids and steam enable us to put them into this series.

The tables show how the elements react with water and dilute acids:

Element	Reaction with water
Potassium	Violently
Sodium	Very quickly
Lithium	Quickly
Calcium	More slowly

Element	Reaction with dilute acids
Calcium	Very quickly
Magnesium	Quickly
Zinc	More slowly
Iron	More slowly than zinc
Copper	Very slowly
Silver	Barely reacts
Gold	Does not react

Note that **aluminium** can be difficult to place in the correct position in the reactivity series during these experiments. This is because its protective aluminium oxide layer makes it appear to be less reactive than it really is. When this layer is removed, the observations are more reliable.

It is useful to place **carbon** and **hydrogen** into the reactivity series because these elements can be used to extract metals.

Displacement reactions of metal oxides

A more reactive metal will **displace** a less reactive metal from a **compound**. The **thermite reaction** is a good example of this. It is used to produce white hot molten (liquid) iron in remote locations for welding. A lot of heat is needed to start the reaction, but then it releases an incredible amount of heat, enough to melt the iron.

aluminium + iron(III) oxide → iron + aluminium oxide
$2Al + Fe_2O_3 \rightarrow 2Fe + Al_2O_3$

Because aluminium is more reactive than iron, it displaces iron from iron(III) oxide. The aluminium removes oxygen from the iron(III) oxide:

- iron is **reduced**
- aluminium is **oxidised**

Reactions between metals and metal oxides allow us to put a selection of metals into a reactivity series. Using metals A, B and C:

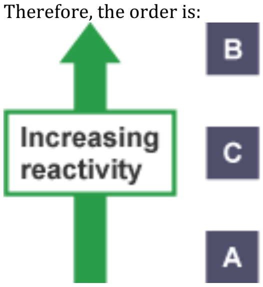
	Metal A	Metal B	Metal C
A oxide	X	Displaces A	Displaces A
B oxide	No reaction	X	No reaction
C oxide	No reaction	Displaces C	X

Metal A cannot displace either B or C - so it must be the **least reactive**and be at the bottom of this reactivity series.
Metal B displaces both A and C - so it must be the **most reactive** and be at the top of this reactivity series.
Metal C displaces A but cannot displace B - so it must be more reactive than A but less reactive than B, and be in between them in this reactivity series.

In general, the greater the difference in reactivity between two metals in a displacement reaction, the greater the amount of energy released.

Aluminium is much higher than iron in the reactivity series, so the thermite reaction releases a lot of energy. Magnesium is very high in the reactivity series, and copper is very low - so the reaction between magnesium and copper oxide is more violent.

Therefore, the order is:



Oxidation and reduction

Oxidation is the loss of **electrons** from a substance. It is also the gain of oxygen by a substance. For example, magnesium is oxidised when it reacts with oxygen to form magnesium oxide:

magnesium + oxygen → magnesium oxide
$2Mg + O_2 \rightarrow 2MgO$

Reduction is the gain of electrons by a substance. It is also the loss of oxygen from a substance. For example, copper(II) oxide can be reduced to form copper when it reacts with hydrogen:

copper(II) oxide + hydrogen → copper + water
$CuO + H_2 \rightarrow Cu + H_2O$
Usually, oxidation and reduction take place at the same time in a reaction. We call this type of reaction a **redox reaction**.

Note that:
- the **oxidising agent** is the chemical that causes oxidation
- the **reducing agent** causes the other chemical to be reduced

Take a look at the following **thermite reaction**:

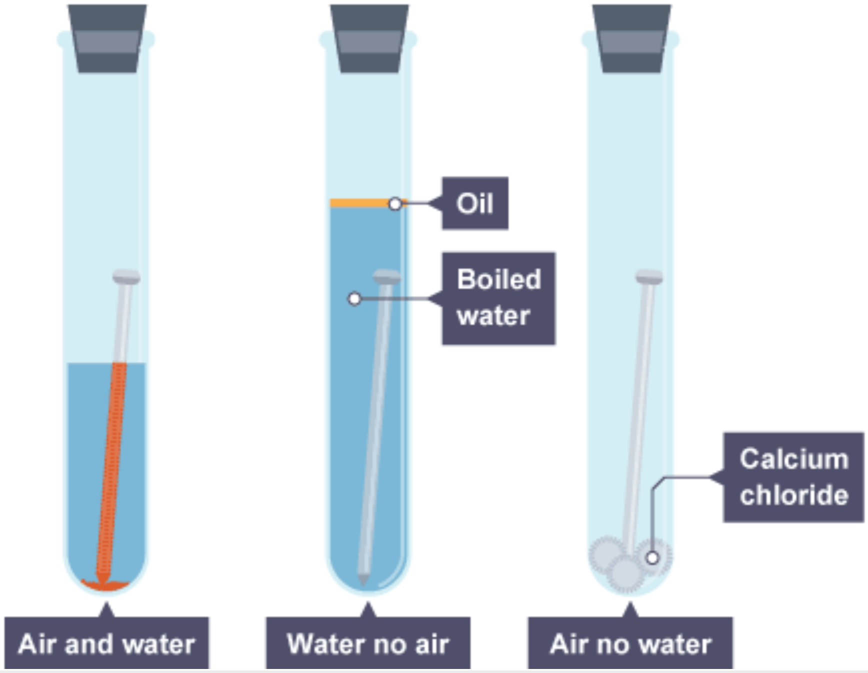
aluminium + iron(III) oxide → iron + aluminium oxide
It is easy to see that the aluminium has been oxidised. This means that the iron oxide is the **oxidising agent**. We can also see that the iron oxide has been reduced. This means that the aluminium is the **reducing agent**.

Rusting

Rusting is an **oxidation** reaction. The iron reacts with water and oxygen to form hydrated iron(III) oxide, which we see as rust. Here is the word equation for the reaction:

iron + water + oxygen → hydrated iron(III) oxide

Iron and steel rust when they come into contact with water and oxygen. Both water and oxygen are needed for rusting to occur. In the experiment below, the nail does not rust when air (containing oxygen) or water is not present:



Calcium chloride absorbs water in the right-hand test tube

Salt dissolved in water does not cause rusting - but it does speed it up, as does **acid rain**.

Aluminium does not rust (corrode) because its surface is protected by a natural layer of aluminium oxide. This prevents the metal below from coming into contact with air (containing oxygen).

Unlike rust, which can flake off the surface of iron and steel objects, the layer of aluminium oxide does not flake off.

Preventing rusting

There are several ways to prevent iron and steel rusting. Some of these work because they stop oxygen or water reaching the surface of the metal:

- oiling - for example, bicycle chains
- greasing - for example, nut and bolts
- painting - for example, car body panels
- coating with a thin layer of plastic

Iron and steel objects may also be covered with a layer of metal. Food cans, for example, are plated with a thin layer of tin.

Galvanising

Galvanising is a method of rust prevention. The iron or steel object is coated in a thin layer of zinc. This stops oxygen and water reaching the metal underneath - but the zinc also acts as a **sacrificial metal**. Zinc is more reactive than iron, so it **oxidises** in preference to the iron object.

Sacrificial protection

A reactivity series lists metals in order of how reactive they are.

Magnesium and zinc are often used as **sacrificial metals**. They are more reactive than iron and lose their electrons in preference to iron. This prevents iron from losing its electrons and becoming oxidised