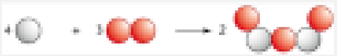
Most metals are found combined as minerals in ores. Many of these minerals are oxides. Methods have been developed to extract them.

Iron (III) oxide

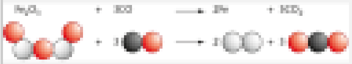
A common ore of iron is haematite, which contains iron oxide.

Iron + Oxygen → Iron oxide

$4Fe + 3O_2 \rightarrow 2Fe_2O_3$



This is an **oxidation** reaction because the metal gains oxygen.
Iron (III) oxide can be **reduced** by carbon monoxide. It is reduced because it loses oxygen.



Aluminium oxide

Bauxite is aluminium oxide.

Aluminium + Oxygen → Aluminium oxide

This prevents aluminium corroding. Aluminium is too reactive to be obtained by reduction by carbon. Electrolysis must be used.

Oxides of Group 1 and Group 2

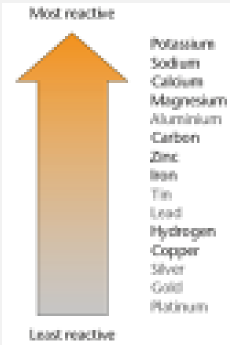
- Sodium oxide is formed when sodium is burned in oxygen
- It is a white powder and is soluble
- During the reaction, Na loses electrons and O gains electrons.

$2Na - 2e^- \rightarrow 2Na^+$ $O_2 + 2e^- \rightarrow O^{2-}$

- This forms a basic oxide
- When dissolved in water, it makes a hydroxide solution (alkaline) and turns universal indicator purple

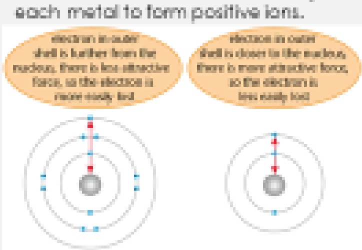
$$Na_2O + H_2O \rightarrow 2NaOH$$

The Reactivity Series



Some metals are so reactive that they react with water. Others are less reactive, but still react with acid. These can be ordered by reactivity.

The differences in reactivity is because of the different **tendency** of each metal to form positive ions.



A more reactive metal can displace a less reactive metal from a solution of a compound of the less reactive metal. This is a **displacement** reaction.

Ion formation

Positive ions are formed by a loss of electrons. ($K - e^- \rightarrow K^+$)
Displacement would be shown by:

$Mg + Fe^{2+} \rightarrow Mg^{2+} + Fe$

Extraction of metals

Reduction

Most metals are found as compounds as they react with other elements

Metals more reactive than carbon need to be extracted by electrolysis.	For example	potassium, sodium, calcium, magnesium, aluminium
Metals less reactive than carbon can be extracted from their oxides by **reduction with carbon**.	For example	zinc iron lead

Reduction involves the **loss of oxygen**

eg. zinc oxide + carbon → zinc + carbon dioxide

Reaction of metals with acids

Metal + Acid → Salt + **Hydrogen**

To make a crystallised salt:

a. excess magnesium needs to be added to the acid
b. the solution needs to be filtered into a crystallising dish
c. the solution needs to be concentrated by evaporation
d. and the solution then left to evaporate, to crystallise.

Forming ions

Metals		Acids			
magnesium	Mg^{2+}	sulfuric acid	H_2SO_4	$2H^+$	SO_4^{2-}
iron	Fe^{2+}	hydrochloric acid	HCl	H^+	Cl^-
zinc	Zn^{2+}				

The charges must cancel each other out and the equation must be balanced

Using half equations

$Mg + 2H^+ \rightarrow Mg^{2+} + H_2$

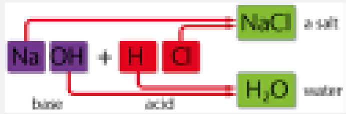
$Mg - 2e^- \rightarrow Mg^{2+}$	$2H^+ + 2e^- \rightarrow H_2$
Electron loss is oxidation	Electron gain is reduction
As both these reactions take place at the same time this is a redox reaction	

Alkalis and Bases

- **A base is any substance that neutralises an acid.**
- Metal oxides and metal hydroxides are **bases**.
- A few bases are soluble in water, these are called **alkalis**.

NEUTRALISATION happens when an acid and base react

Acid + Base → Salt + Water



Metal **carbonates** react with acids to make a metal salt and water, but carbon dioxide is also produced.

Acid + Metal Carbonate → Salt + Water + Carbon Dioxide

Common acid	The salt made
nitric	nitrate
sulfuric	sulfate
hydrochloric	chloride

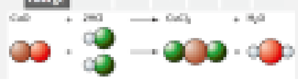
The second part of the salt name is taken from the acid. The first part is taken from the base.

Ions from alkalis and carbonates		Ions from bases and carbonates		Ions from acids	
sodium	Na^+	magnesium	Mg^{2+}	chloride	Cl^-
potassium	K^+	zinc	Zn^{2+}	nitrate	NO_3^-
calcium	Ca^{2+}	copper	Cu^{2+}	sulfate	SO_4^{2-}

eg. $MgSO_4$, K_2SO_4, KCl, $CaCl_2$ → BALANCED CHARGES

Soluble Salts

Single charge	Cl^-	NO_3^-	OH^-	Na^+	K^+	H^+
Double charge	SO_4^{2-}	Ca^{2+}	Mg^{2+}	Zn^{2+}	Fe^{2+}	Cu^{2+}



Acids and alkalis

Acids are substances that produce hydrogen ions in aqueous solution.

Examples of acids are:		Examples of alkalis are:	
hydrochloric acid	HCl	sodium hydroxide	NaOH
nitric acid	HNO_3	potassium hydroxide	KOH
sulfuric acid	H_2SO_4	ammonia	$NH_3(aq)$
ethanoic acid	CH_3COOH		
citric acid	$C_6H_8O_7$		

The hydrogen ions they produce have the symbol H^+. They are ions with a positive charge.	Alkalis are substances that make hydroxide ions in aqueous solution. These hydroxide ions have the symbol OH^-. They are ions with a negative charge.

The pH Scale

When **universal indicator** (UI) is added to solutions, it changes colour.

1 = very acidic, 7 = neutral, 14 = very alkaline

The higher the concentration of acid, the lower the pH

Neutralisation

If an acid is added to an alkali, neutralisation takes place

An acid solution has a low pH. If an alkali is added slowly to an acid, the pH number of the acid will gradually increase. When it gets to pH 7 the acid is neutralised.	An alkaline solution has a high pH. If acid is slowly added to an alkali, the pH number will gradually decrease. When it gets to pH 7 the alkali has been neutralised.

$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

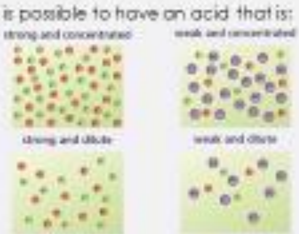
Neutralisation leaves no free H^+ ions.

Strong and Weak Acids

Strong: In water all of the acid molecules, HA, become ions (H^+ and A^-).	Weak: In water only a few of the acid molecules, HA, become ions (H^+ and A^-), most stay as molecules.
Strong acids ionise completely in water. $HCl \rightarrow H^+ + Cl^-$	Weak acids do not ionise fully. The equilibrium lies to the left. $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$
A high concentration of H^+ means that the pH is low.	A low concentration of H^+ means that the pH is higher.

strong acids	weak acids
hydrochloric nitric sulfuric	ethanoic citric carbonic

It is possible to have an acid that is:



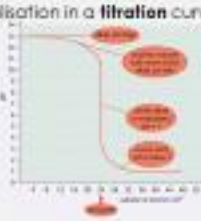
pH, Neutralisation and Titration Curves

- The pH scale is related to the concentration of H^+ ions
- Strong acids have a lower pH

Low pH = higher concentration H^+	High pH = lower concentration H^+

- As the pH decreases by one unit, the H^+ concentration increases by a factor of 10

Neutralisation in a **titration** curve:



Electrolysis

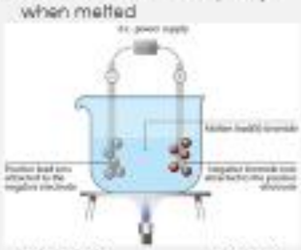
Electrolysis is the process of passing direct current (d.c.) through a solution or melted ionic compound to move ions apart and so break the compound down and discharge some of the elements at the electrodes.

The cathode is the negative electrode.	The anode is the positive electrode.
positive ions are attracted to the cathode and are called **cations**	negative ions are attracted to the anode and are called **anions**

Ions discharging:
The half equation for the discharge of a sodium ion would be:
$Na^+ + e^- \rightarrow Na$
The half equation for the discharge of a chlorine ion would be:
$Cl^- - e^- \rightarrow Cl$
But chlorine is a molecule of 2 atoms, so:
$2Cl^- - 2e^- \rightarrow Cl_2$

Simple Binary Electrolysis

- Is made up of two ions (eg. lead bromide)
- Can conduct electricity **only** when melted



At the cathode	At the anode
Electrons are gained	Electrons are given up

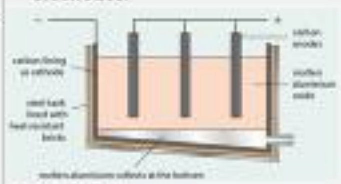
Electrolyte	Half equation at cathode	Half equation at anode
KCl	$2K^+ + 2e^- \rightarrow 2K$	$2Cl^- - 2e^- \rightarrow Cl_2$
$CaCl_2$	$Ca^{2+} + 2e^- \rightarrow Ca$	$2Cl^- - 2e^- \rightarrow Cl_2$
PbI_2	$Pb^{2+} + 2e^- \rightarrow Pb$	$2I^- - 2e^- \rightarrow I_2$
Al_2O_3	$2Al^{3+} + 6e^- \rightarrow 2Al$	$6O^{2-} - 12e^- \rightarrow 3O_2$

Electrolysis to extract metals

Metals found towards the top of the reactivity series are too reactive to be extracted by carbon, so must be extracted by electrolysis. Large amounts of energy are used to melt the compounds and to produce an electric current, so these metals can be expensive. **Aluminium** is extracted in this way. It's very expensive to extract which is why aluminium **recycling** is encouraged.

Electrolysis of aluminium

- Aluminium ore is **bauxite** (Al_2O_3).
- It is mixed with **cryolite** to lower its melting temperature.
- The electrolysis cell has **carbon electrodes**.



Half equations

$2Al_2O_3 \rightarrow 4Al + 3O_2$

At the cathode:

- $4Al^{3+} + 12e^- \rightarrow 4Al$ ($Al^{3+} + 3e^- \rightarrow Al$ – per ion)

At the anode:

- $6O^{2-} + 12e^- \rightarrow 3O_2$

The oxygen reacts with the electrodes:

- $3C + 3O_2 \rightarrow 3CO_2$

Preferential discharge of ions

At the cathode	At the anode
H^+ is discharged in preference to Zn^{2+}, Al^{3+}, Mg^{2+}, Ca^{2+}, Na^+ and K^+	OH^- is discharged in preference to SO_4^{2-}
Cu^{2+} and Ag^+ are discharged in preference to H^+ (as Cu^{2+} and Ag^+ are less reactive than H^+)	Cl^- and Br^- are discharged in preference to OH^-
H^+ produces hydrogen gas	OH^- produces oxygen gas

The electrolysis of copper sulphate

At the cathode	At the anode
Copper ions (Cu^{2+}) Hydrogen ions (H^+) (from the water)	Sulfate ions (SO_4^{2-}) Hydroxide ions (OH^-) (from the water)
Hydrogen ions stay in the electrolyte. Copper ions are discharged in preference and make copper metal.	Sulfate ions stay in the electrolyte as hydroxide ions are more easily discharged. Hydroxide ions discharge in preference and form oxygen gas.