Whitley Academy

Chemistry Paper 1 (F) Knowledge Recall Booklet

Paper Chemistry 1F 8464/C/1F

For this paper, the following list shows the major focus of the content of the exam:

- 5.1.2 The periodic table
- 5.2.2 How bonding and structure are related to the properties of substances
- 5.2.3 Structure and bonding of carbon
- 5.4.1 Reactivity of metals
- 5.4.2 Reactions of acids
- 5.4.3 Electrolysis

Required practical activities that **will be assessed**:

- Required practical activity 8: preparation of a pure, dry sample of a soluble salt from an insoluble oxide or carbonate, using a Bunsen burner to heat dilute acid and a water bath or electric heater to evaporate the solution.
- Required practical activity 9: investigate what happens when aqueous solutions are electrolysed using inert electrodes. This should be an investigation involving developing a hypothesis.
- Required practical activity 10: investigate the variables that affect temperature changes in reacting solutions such as, eg, acid plus metals, acid plus carbonates, neutralisations, displacement of metals.

Your teacher will tell you specific topic areas to focus on …

Topic	RAG			Revision technique					Date completed	Teacher Signed
	R	A	G	Flashcards	Mindmap	Notes	Video Watched (With notes)	Frog resource		

Required Practical – Making Soluble Salts

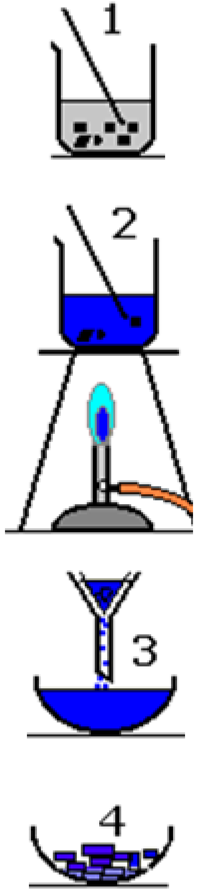
Recall it ...

Use the information in the following page(s) to answer these questions ...

1. Outline all the steps / method for this investigation in sufficient detail.
2. What are the reactants to make a soluble salt?
3. What safety precautions are required?
4. Describe the method for electrolysis?

Required Practical – Preparation of soluble salt

Soluble salts can be made from **acids** by reacting them with **solid insoluble substances**, such as **metals, metal oxides, hydroxides or carbonates**. The solid is added to the acid until no more reacts and the excess solid is **filtered** off to produce a solution of the salt. Salt solutions can be **crystallised** to produce **solid** salts. **You will complete this as a required practical.**



1. Measure the required volume of acid with a measuring cylinder and add the weighed solid (insoluble metal, oxide, hydroxide or carbonate) in small portions with stirring.

2. Safety goggles required - the mixture may be heated to speed up the reaction. When no more of the solid dissolves it means ALL the acid is neutralised and there should be a little excess solid. You should see a residue of the solid (oxide, hydroxide, carbonate) left at the bottom of the beaker.

3. Filter the solution to remove the excess solid metal/oxide/carbonate, into an evaporating dish. On filtration, only a solution of the salt is left.

4. Then hot concentrated solution is left to cool and crystallise.
   After **crystallisation**, you collect and dry the crystals with a filter paper.
   If the solution is heated, the solvent will evaporate faster. Heating a solution until all the solvent has evaporated is known as **heating to dryness**.

Required Practical – Electrolysis

Recall it …

Use the information in the following page(s) to answer these questions …

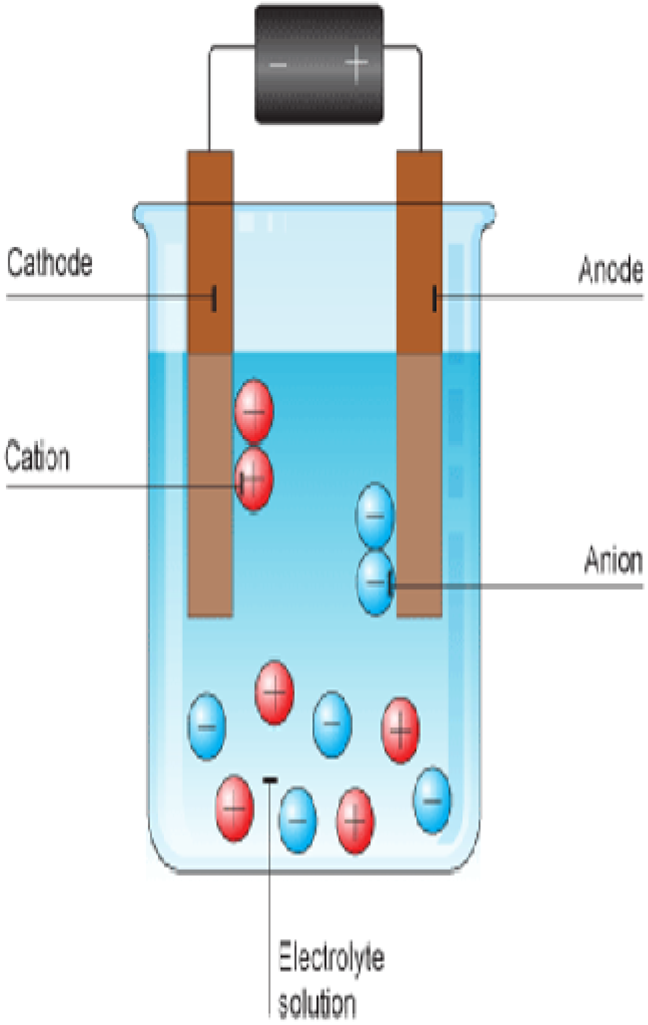
1. What happens in electrolysis?
2. What is the anode? What happens at the anode?
3. What is the cathode? What happens at the cathode?
4. Describe the method for electrolysis?

Required Practical - Electrolysis

3. Electrolysis of Aqueous Solution (aq)

CONTENT LINK

- In electrolysis, the ions move towards the oppositely charged electrodes.
- When you electrolyse aqueous ionic solutions - you get three products..
- When electrolysis happens in aqueous solution, the less reactive element, either hydrogen or the metal, is usually produced at the Cathode. At the Anode, you get either;
  - Oxygen gas given off, from discharged hydroxide ions produced from water.
  - Halogen produced if the electrolyte is a solution of a halide



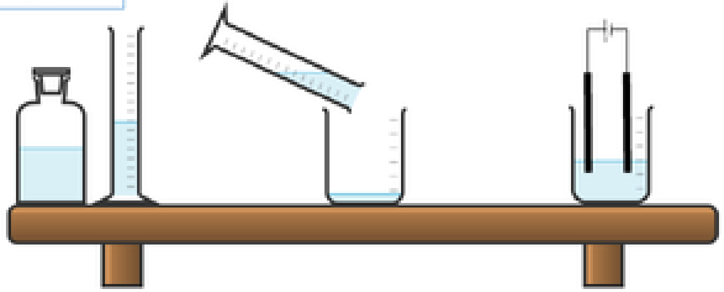
Positive
Anode
Negative
is
Cathode

Electrolysis

1. Measure 50 cm³ of copper (II) chloride solution

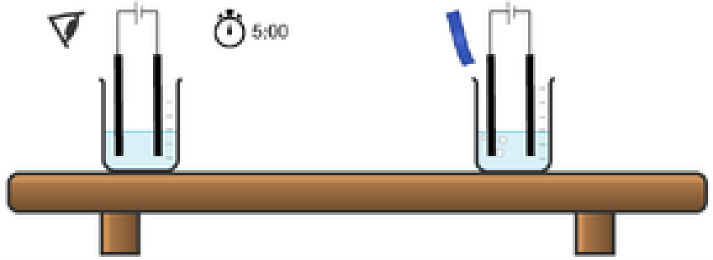
2. Pour into a 100 cm³ beaker

3. Insert the graphite electrodes into the solution. Connect one to the positive outlet of the power pack and the other to the negative.



4. Turn the power pack to 4V and switch on for no longer than 5 minutes. Observe what is happening at each electrode.

5. If you can see bubbles forming at the anode (positive electrode) hold a piece of damp, blue litmus paper next to the bubbles. If the paper turns white, the bubbles are chlorine.



6. Repeat steps 1 to 5 for solutions of copper (II) sulfate, sodium chloride and sodium sulfate.

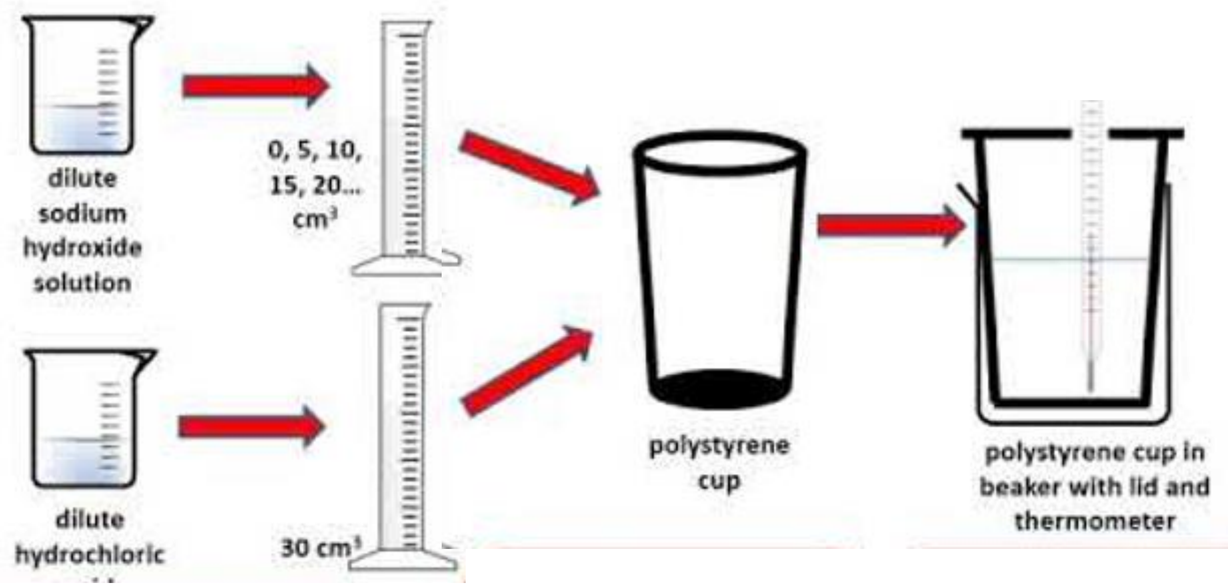
Required Practical – Temperature Change

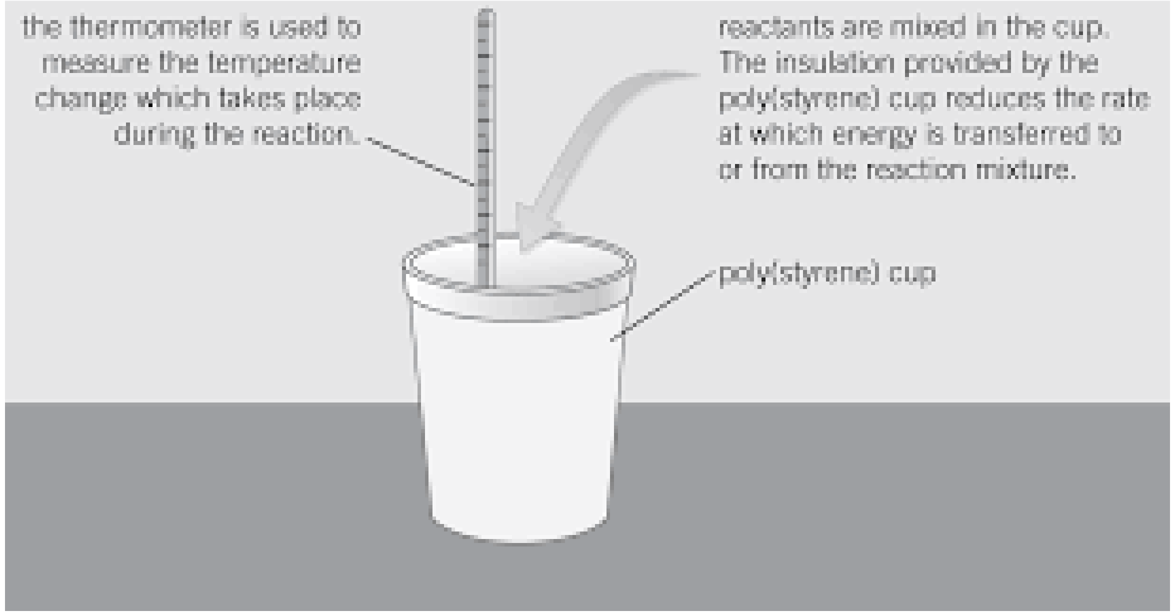
Recall it …

Use the information in the following page(s) to answer these questions …

1. What are the reactants for this experiment?
2. What is the independent, dependent and control variables for this experiment?
3. Describe the method for the experiment?
4. Why is a polystyrene cup used?

Required Practical – Temperature Change





Method:

1. Measure 30cm^3 dilute hydrochloric acid and put it into the polystyrene cup.
2. Stand the cup inside the beaker. This will make it more stable
3. Use the thermometer to measure the temperature of the acid. Record your result in the table below.
4. Measure 5cm^3 sodium hydroxide solution.
5. Pour the sodium hydroxide into the polystyrene cup. Fit the lid and gently stir the solution with the thermometer through the hole.
6. Look carefully at the temperature rise on the thermometer.
7. When the reading on the thermometer stops changing, record the highest temperature reached in the table.
8. Repeat steps 4-7 to add further 5cm^3 amounts of sodium hydroxide to the cup each time, recording your temperature reading in the results table.
9. Repeat until a maximum of 40cm^3 of sodium hydroxide has been added.
10. Wash out all the equipment and repeat the experiment for your second trial.

The Periodic table

Recall it ...

Use the information in the following page(s) to answer these questions …

1. What are the column and rows called?
2. How many groups are there?
3. How is it arranged? In order of ……
4. What does the group number tell you about the number of electrons on the outer shell?
5. What does the period tell you about the number of shells?
6. How is it divided into metals and non-metals?
7. Describe the properties of metals?
8. Describe the properties of non-metals?
9. What did John Dalton, Newlands and Mendeleev do?
10. How did Mendeleev order his periodic table?

Periodic table - PART 1

The elements are arranged in order of increasing atomic number.

1	2											3	4	5	6	7	0
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	?	?	?						

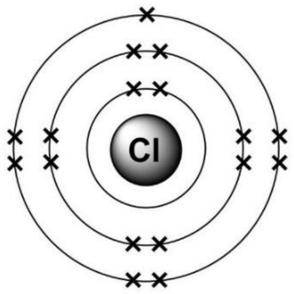
Elements with similar properties are in columns, known as groups.

Elements in the same group have the same number of electrons in their outer shell.

The rows in the table are called periods

It is called a periodic table because similar properties occur at regular intervals

Group = electrons in outer shell
Period = number of shells



Group = 7
Period = 3

Periodic table - PART 1

The elements can be divided into metals and non-metals.

1	2											3	4	5	6	7	0
					H												He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	?	?	?						

Elements that do not form positive ions are non-metals

Elements that tend to form positive ions are metals

Non metals – found towards the right and towards the top of the periodic table

Most elements are metals – found towards the left and towards the bottom of the periodic table

Metals	Non-metals
Shiny	Dull
Mostly solid	Low density
Dense and strong	Weak
Malleable	Brittle
Good heat and electrical conductors	Poor heat and electrical conductors

Periodic table - PART 1

1808

John Dalton published a table of elements that were arranged in order of their atomic weights, which had been measured in various chemical reactions

1864

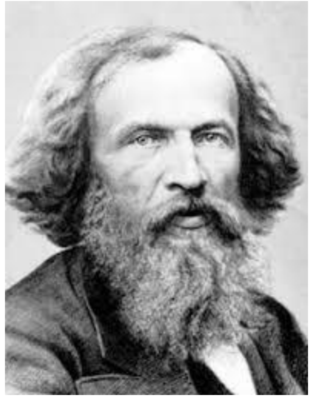
John Newlands published the law of octaves. However the **table was incomplete and elements were placed in inappropriate groups**

Newlands' Arranged Elements in Octaves:

H	F	Cl	Co/Ni	Br	Pd	I	Pt/Ir
Li	Na	K	Cu	Rb	Ag	Cs	Tl
G	Mg	Ca	Zn	Sr	Cd	Ba/V	Pb
Bo	Al	Cr	Y	Ce/La	U	Ta	Th
C	Si	Ti	In	Zn	Sn	W	Hg
N	P	Mn	As	Di/Mo	Sb	Nb	Bi
O	S	Fe	Se	Ro/Ru	Te	Au	Os

1869

Dmitri Mendeleev overcame Dalton's problem by **leaving gaps for the elements that he thought had not been discovered** and **in some places changed the order based on atomic weight** (e.g. Argon and Potassium). Elements with properties predicted by Mendeleev were eventually discovered.



Early20th Century - Scientists began to find out more about the atom and knowledge of **isotopes** explained why the order was not always correct.

Groups 8, 1 and 7

Recall it ...

Use the information in the following page(s) to answer these questions …

For group 8

1) What are they called?
2) Why are they unreactive?
3) Why don't they form molecules?
4) What is the trend in boiling points?

For group 1

1) What are they called?
2) Why are they stored in oil?
3) Are they hard or soft?
4) Why does reactivity increase down group 1?
5) Describe what happens when group 1 metals react with oxygen, water and chlorine?

For group 7

1) What are they called?
2) What are the properties of halogens?
3) They diatomic – what does this mean?
4) Why does reactivity decrease down group 7?
5) What happens why halogens react with hydrogen, metals and other halogens?

Periodic table - PART 2

THE NOBLE GASES

1	2											3	4	5	6	7	0
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	?	?	?						

Elements in **Group 0** of the periodic table are called the **noble gases**. They are unreactive because their atoms have **stable** arrangements of **electrons**. The atoms have **eight** electrons in their outermost shell, apart from helium, which has just **two**, but still has a complete outer shell.

The stable electronic structure explains why they exist as **single atoms**, they have no tendency to react to form **molecules**.

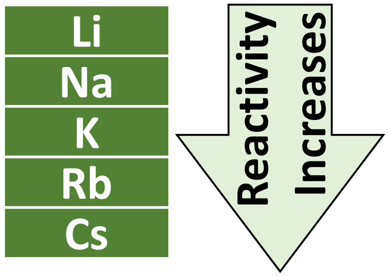
The **boiling points** of the noble gases get **higher** going **down** the group. For example helium boils at -269 °C and radon boils at -62°C.

Periodic table - PART 2

THE ALKALI METALS

1	2											3	4	5	6	7	0
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	?	?	?						

The alkali metals are very reactive. They need to be stored under **oil** to prevent them reacting with **oxygen** and **water** vapour in the **air**. The alkali metals have **low densities**. The metals are very **soft** and can be **cut** with a knife. They also have **low melting** and **boiling points**.
The properties are due to all the atoms having just **one** electron in their outermost shell. They only need to **lose** one electron to get the stable electronic structure of a noble gas.



The atoms get **larger** as you go **down**, so the **single** electron in the outermost shell (highest energy level) is attracted **less** strongly to the positive nucleus. The **electrostatic attraction** with the nucleus gets **weaker** because the **distance** between the outer electron and the nucleus **increases**. Also the outer electron experiences a **shielding effect** from the inner electrons, **reducing** the attraction between the oppositely charged outer electron and the nucleus.

Periodic table - PART 2

The alkali metals have a **silvery**, **shiny** surface when they are first cut. However, this goes **dull** very quickly as the metals reacts with the oxygen in the air.

e.g. sodium + oxygen → sodium oxide

$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

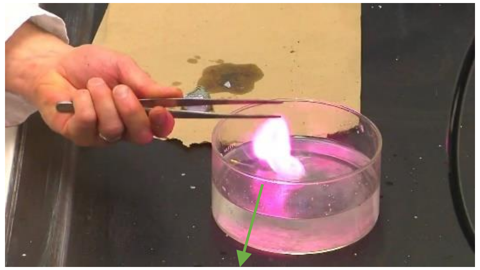
shiny

metal oxide

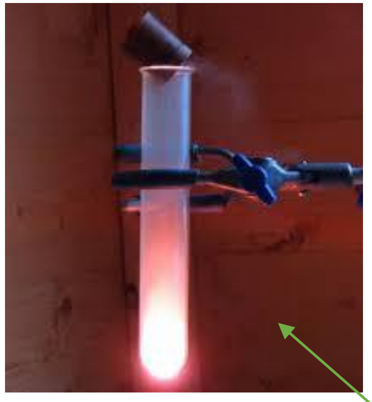
Lithium, sodium and potassium all react **vigorously** with water. When you add them to water, the metal **floats**, **moves** around and **fizzes**.

e.g. potassium + water → potassium hydroxide + hydrogen

$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$



Potassium ignites with a lilac flame



They also react vigorously with non metals, such as group seven. They form **1+** ions in the reactions to make **ionic compounds**. These are generally **white** and **dissolve** in water, **giving colourless solutions**.

e.g. sodium + chlorine → sodium chloride

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

Periodic table - PART 2

1	2											3	4	5	6	7	0
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	?	?	?						

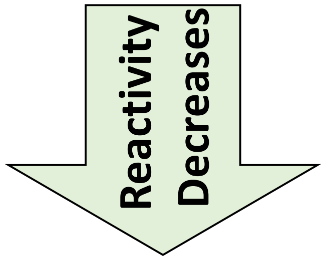
THE HALOGENS

The halogens are a group of **toxic** non-metals that have **coloured** vapours. They have **low melting** and **boiling** points, which **increase** down the group. They are **poor conductors** of **heat** and **electricity**.

As elements, the halogens exist as **molecules** made up of pairs of atoms. These are called **diatomic** molecules **F_2, Cl_2, Br_2, I_2 and At_2**. The halogens have **seven** electrons in their outermost shell and need to **gain** one electron to achieve the stable electronic structure of a noble gas. When they react with non metals, they are joined together by a **covalent** bond.

When Group 7 elements react, the atoms **gain** an electron in their outermost shell. Going down the group, the outermost shell's electrons get **further away** from the attractive force of the nucleus, so it is **harder** to **attract** and **gain** an extra electron. The outer shell will also be **shielded** by more inner shells of electrons, again **reducing** the **electrostatic attraction** of the nucleus for an incoming electron.

F
Cl
Br
I
At



Periodic table - PART 2

The halogens react with **hydrogen**. The reactions with hydrogen become **less** reactive as you go **down** the group.

e.g. fluorine + hydrogen → hydrogen fluoride

$F_2(g) + H_2(g) \rightarrow 2HF(g)$

The halogens also react with **metals**. The halogen atoms **gain** a single electron to give them a stable arrangement of electrons. They form **ionic compound**.

e.g. sodium + chlorine → sodium chloride

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

A more reactive halogen will also **displace** a less reactive halogen from solutions of its **salts**.

e.g. chlorine + potassium bromide → potassium chloride + bromine

$Cl_2(g) + 2KBr(aq) \rightarrow 2KCl(aq) + Br_2(aq)$

The **colour** of the solution after mixing depends on the **less** reactive pair of halogens.

$Cl_2(aq)$ $Br_2(aq)$ $I_2(aq)$

Changes of state and heating curves

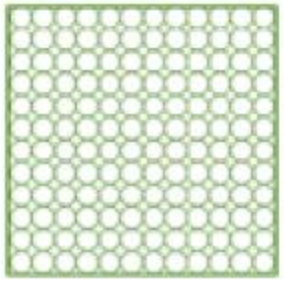
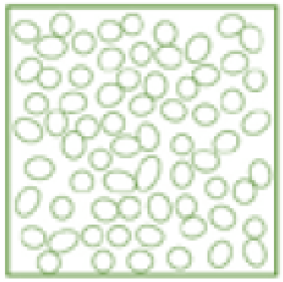
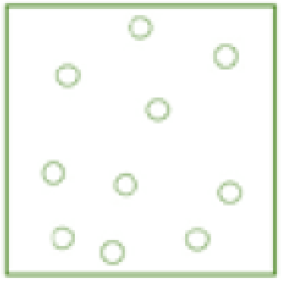
Recall it …

Use the information in the following page(s) to answer these questions …

1) Describe how particles behave in a solid?
2) Describe how particles behave in a liquid?
3) Describe how particles behave in a gas?
4) Describe 4 changes of state?
5) What happens at the melting point and the boiling point?
6) What does the energy needed to change state depend upon?
7) Why may some substances have high melting and boiling points?
8) What happens when the temperature does not change in a heating curve?

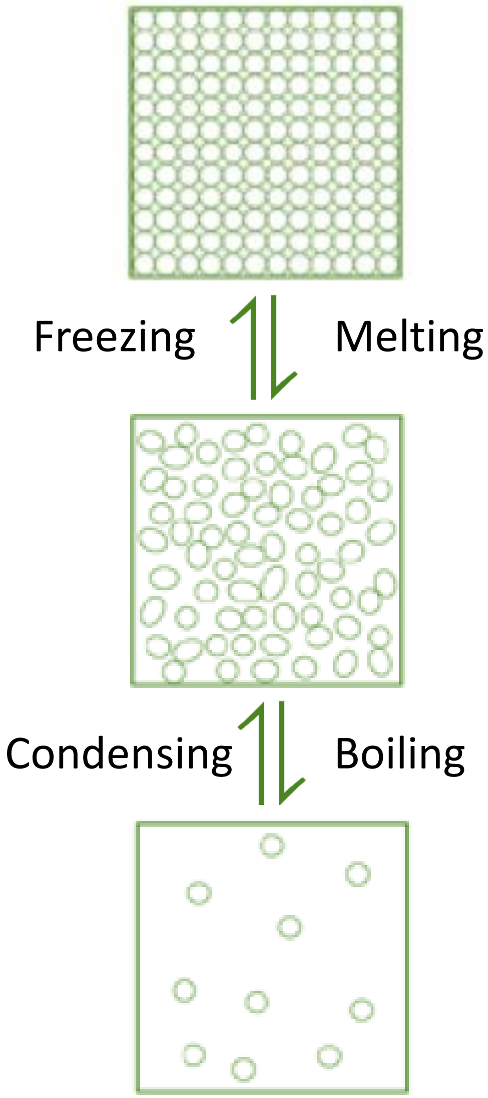
States of matter and state symbols

There are **three** states of matter – **solid, liquid** and **gas**. To explain the properties of the states, the **particle theory** is used. It is based on the fact that all matter is made up of tiny particles and describes the **movement** and **distance** between particles.

Solid	Liquid	Gas
Close together, regular pattern, vibrate on the spot.	Close together, random arrangement, move around each other.	Far apart, random arrangement, move quickly.
		

In chemical equations, the three states are shown as **(s), (l), (g)** and **(aq)** for aqueous solutions.

Changes of state



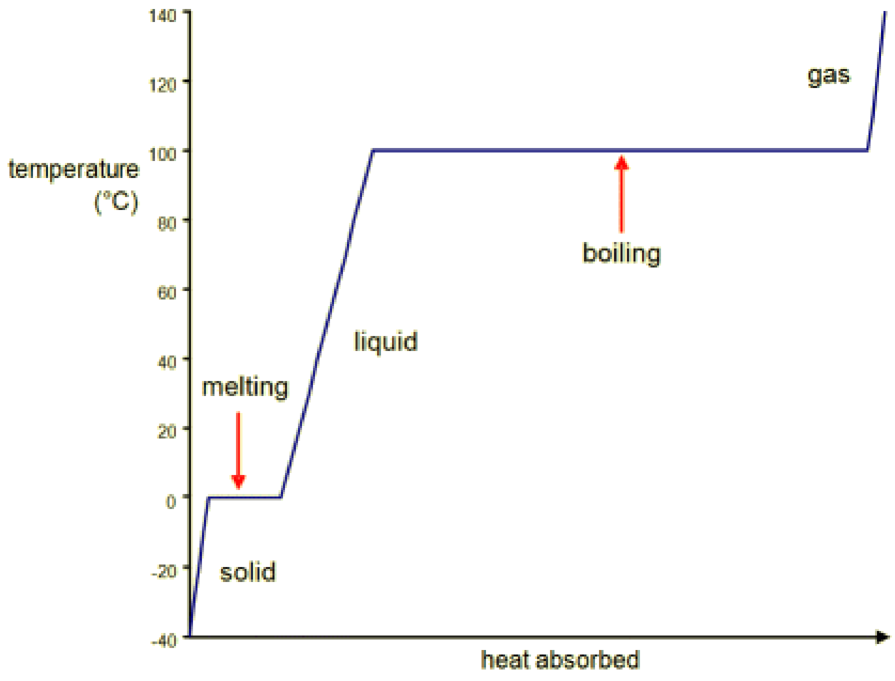
Melting and **freezing** take place at the **melting point**.
Boiling and **condensing** take place at the **boiling point**.

The **amount of energy** required to change the state depends on the **strength of the forces** between the particles of the substance.

The **stronger the forces** between the particles the **higher the melting and boiling point** of the substance.

The type of bonding and the structure of the substance depend on the particles involved.

The graph shows a **heating curve** of a solid, which shows the temperature of a substance plotted against the amount of energy it has absorbed:



A substance must **absorb** heat energy so that it can **melt** or **boil**. The **temperature** of the substance does **not change** during **melting**, **boiling** or **freezing**, even though energy is still being transferred.

Recall it ...

Structures

Use the information in the following page(s) to answer these questions …

1) Describe the structure of sodium chloride (NaCl)
2) Describe and explain the properties of sodium chloride (NaCl)
3) Describe the structure of small molecules?
4) Describe and explain the properties of small molecules?
5) How the size of small molecules affect their melting and boiling points? Explain why?
6) What are polymers? Describe and explain the properties of polymers?
7) Describe and explain the properties of metals?
8) Describe and explain the properties of alloys?
9) Describe and explain the properties of giant covalent structures?
10) Describe and explain the properties of diamond?
11) Describe and explain the properties of graphite?
12) Describe and explain the properties of graphene?
13) Describe and explain the properties of fullerenes?

Properties of ionic compounds

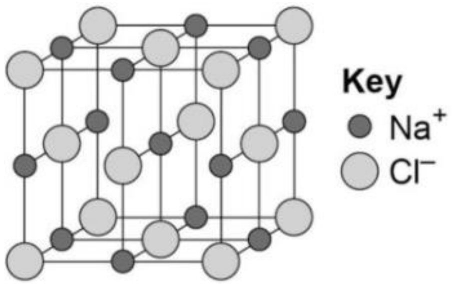
Structure

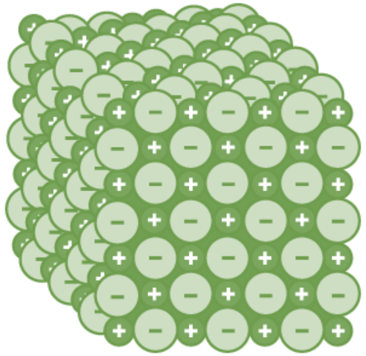
Ionic compounds have **regular structures** called **giant ionic lattices**.

There is **strong electrostatic forces** of attraction in all directions between **oppositely changed ions**.

Properties

- **High melting and boiling points** – large amounts of energy is needed to break the many **strong bonds** and overcome the electrostatic attraction.
- **Conduct electricity** when molten or dissolved in water – **ions are free to move** and can carry charge.





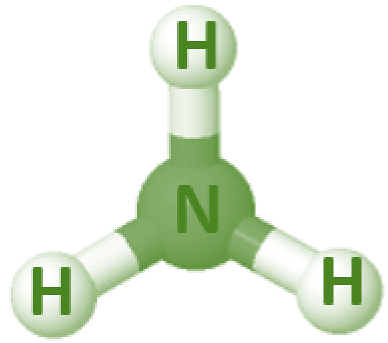
Properties of small molecules

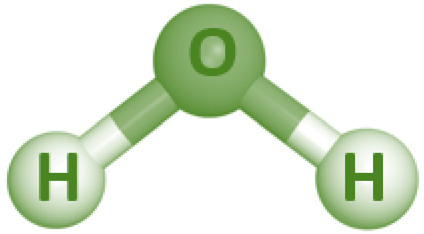
Structure

They have **weak forces between the molecules**. These weak forces are overcome when they change state **not** the strong covalent bonds.

Properties

- **Low melting and boiling points** – small amounts of energy is needed to break the **intermolecular forces**. Most are gases or liquids.
- **Do not conduct electricity** – Particles do not have an overall electric charge.





Intermolecular forces increase with the size of the molecules. So larger molecules have higher melting and boiling points.

Polymers

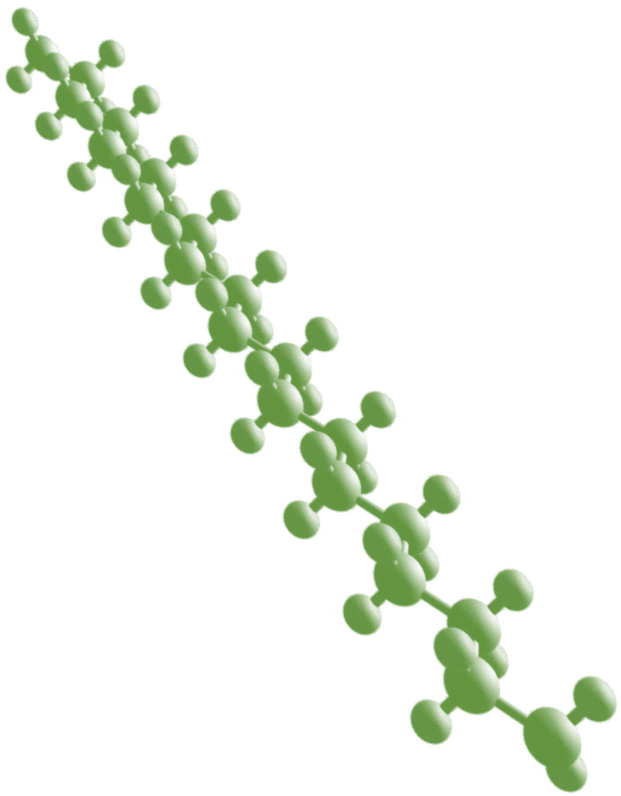
Some covalently bonded substances have very **large** molecules, such as **polymers**.

Structure

Polymers are made up from many small reactive molecules that bond to each other to form **long chains**. The atoms in the polymer molecules are linked to other atoms by **strong covalent bonds**. The **intermolecular forces** between polymer molecules **are relatively strong**.

Properties

- **Solid** at room temperature – **Strong intermolecular forces**.

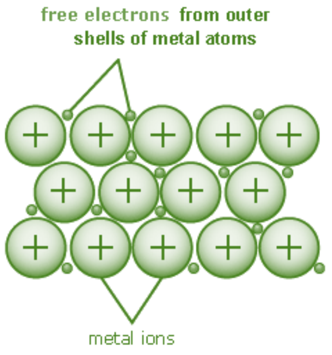


Properties of metals and alloys

The giant structure of atoms with strong metallic bonding gives most metals a **high melting** and **boiling point**.

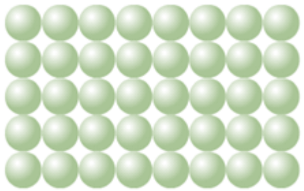
Metals are **malleable** (can be hammered into shape) and **ductile** (can be drawn out into a wire) because the **layers** of atoms (or ions) in a giant metallic structure can **slide** over each other

Delocalised electrons in metals enable **electricity** and **heat** to pass through the metal easily.

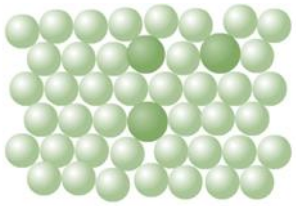


A **metal mixed** with other **elements** is called an **alloy.** Alloys are **harder** than pure metals. Alloys are made from **two or more** different metals.

Pure metal



Alloy



The **different sized** atoms of the metals **distort** the **layers** in the structure, making it more **difficult** for them to **slide** over each other, and so make the **alloys harder** than pure metals.

For example, **gold** is naturally **soft** but adding **copper** to make jewellery **stronger** and last longer.

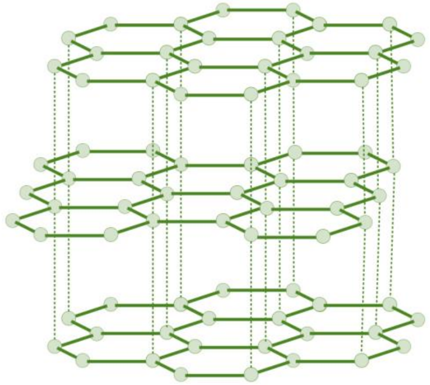
Giant covalent structures

Structure

All atoms within the structure are linked by **strong covalent bonds**. These bonds **must be broken** for a solid to melt or boil.

Properties

- **Very high melting and boiling points** – very large amounts of energy is needed to break the **covalent bonds**.
- **Do not conduct electricity** – Particles do not have an overall electric charge.

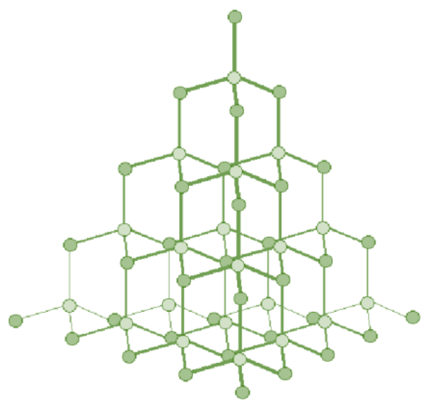


Diamond

Diamond:

In diamond, each carbon atom forms **four covalent bonds** with other carbon atoms in a **giant covalent structure**.

- Diamond is very **hard** – it is the hardest natural substance, so it is often used to make jewellery and cutting tools.
- Diamond has a **very high melting** and **boiling point** – a lot of energy is needed to break the covalent bonds.
- Diamond **cannot conduct electricity** – there are no free electrons or ions to carry a charge.

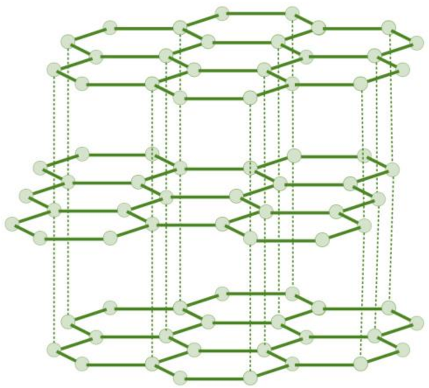


Graphite

Graphite:

In graphite, carbon atom forms **three** covalent bonds with three other carbon atoms, forming layers of **hexagonal rings** which have no covalent bonds between the layers.

- **Graphite is soft and slippery** – layers can easily slide over each other because the weak forces of attraction between the layers are easily broken. This is why graphite is used as a lubricant.
- **Graphite conducts electricity** – the only non-metal to do so. One **electron** from each carbon atom is **delocalised**.

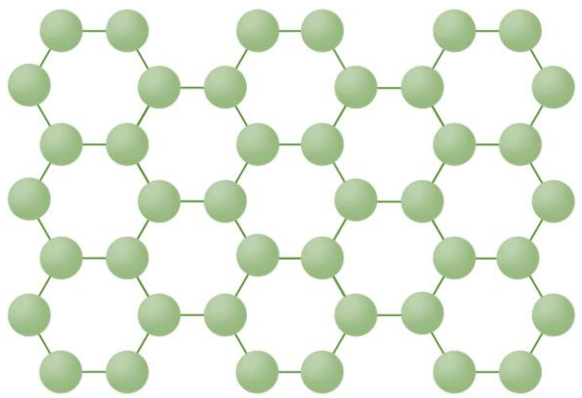


Graphene

Graphene:
This is a **single layer** of graphite – a layer of inter-locking hexagonal rings of carbon atoms **one atom thick**.

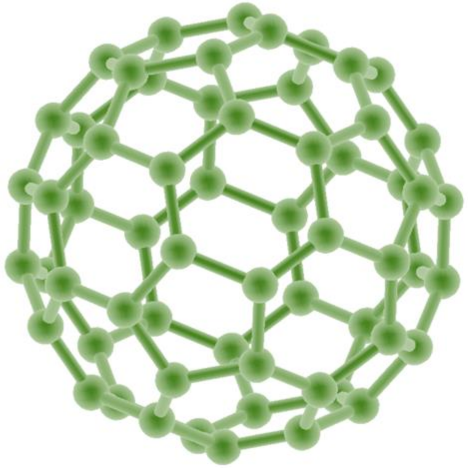
It is an excellent **conductor** of **thermal** energy and **electricity** (even better than graphite), has a very **low density** and is incredibly **strong**.

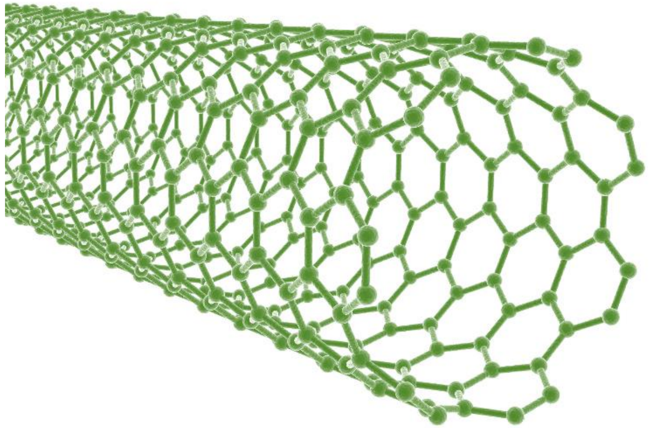
It has many uses in the **electronics industry**.



Fullerenes:

Fullerenes are molecules of carbon with **hollow shapes**. The structure is based on hexagonal rings of carbon atoms, but may have 5 or 7 carbon rings. The first to be discovered was **Buckminsterfullerene** (C_{60}) which is spherical (like a football).





Carbon nanotubes are **cylindrical fullerenes** with very **high length compared to their diameter**. makes them useful for nanotechnology, electronics and materials.

Recall it … Reactivity of Metals

Use the information in the following page(s) to answer these questions …

1) Describe what is an ore? How do we remove metals from their oxides?
2) What is the reactivity series?
3) Describe observations of metal reactivity with water?
4) Describe observations of metal reactivity with acid?
5) Which metals are extracted by electrolysis?
6) Which metals are extracted by reduction with carbon?
7) Which metals are found native?
8) Describe the reaction between metals and acid, giving an example equation?
9) Describe how does the name of the acid change the name of the salt produced?
10) Give the equations for the reactions between

a) Acid and Alkali

b) Acid and Base

c) Acid and Metal carbonate

11) Describe how to work out the chemical formula of salts?
12) What is an indicator? Describe the colour changes and pH values of acids, alkalis and neutral substances?
13) What types of ions are produced by acids and alkalis? What happens to these ions during neutralisation reactions?

Metal Oxides

A **metal compound** within a **rock** is an **ore.** Ores are **mined** and then **purified**.

Whether it is worth extracting a particular metal depends on:

- **How easy it is to extract it from its ore**
- **How much metal the ore contains**
- **The changing demands for a particular metal**

Most **metals** in ores are **chemically bonded** to **other elements** in compounds. Many of these metals have been **oxidised** (have oxygen added) by oxygen in the air to form their oxides.

Iron + oxygen → iron (III) oxide

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

To extract metals from their oxides, the metal oxides must be **reduced** (have oxygen removed).

Reactivity series

Metals can be arranged in order of reactivity in a **reactivity series**.

Order of reactivity	Reaction with water	Reaction with acid
Potassium	Fizz, giving off hydrogen gas and leaving an alkaline solution of metal hydroxide	Reacts violently and explodes
Sodium		
Lithium		
Calcium		Fizz, giving off hydrogen gas and forming a salt
Magnesium	Very slow reaction	
Aluminium		
Zinc		
Iron		
Tin	No reaction with water at room temperature	React slowly with warm acid
Lead		
Copper	No reaction	No reaction
Silver		
Gold		

Reactivity series

Metals can be arranged in order of reactivity in a **reactivity series**.

When metals react with other substances the metal atoms form **positive ions**.

The reactivity of a metal is linked to its tendency to form positive ions.

The **non-metals hydrogen** and **carbon** are often included in the series as they can be used to extract less reactive metals.

Metal + acid → salt + hydrogen

Potassium
Sodium
Lithium
Calcium
Magnesium
CARBON
Zinc
Iron
Lead
HYDROGEN
Copper
Silver
Gold

Increasing reactivity

Extracting metals

The **reactivity** of a metal determines the **method** of **extraction.**

Metals **above** carbon must be extracted from their ores by using **electrolysis.**

Metals **below** carbon can be extracted from their ores by **reduction** using **carbon.**
REDUCTION involves the loss of oxygen.

metal oxide + carbon → metal + carbon dioxide

Gold and **silver** do not need to be extracted. They occur **native** (naturally).

Potassium
Sodium
Calcium
Magnesium
Aluminium
CARBON
Zinc
Iron
Lead
HYDROGEN
Copper
Silver
Gold
Platinum

Increasing reactivity

Reactions of acids - PART 1

Acids react with some **metals** to produce **salts** and **hydrogen**.

Metal + acid → salt + hydrogen

Reactions between metals and acids only occur if the metal is **more reactive** than the **hydrogen** in the acid. If the metal is too reactive, the reaction with acid is **violent**.

The **salt** that is made depends on the **metal** and **acid** used.

Salts made when metals react *nitric acid* are called *nitrates*.

Zinc + Nitric acid → Zinc Nitrate + Hydrogen

Salts made when metals react with *sulfuric acids* are called *sulfates*.

Iron + Sulfuric Acid → Iron Sulfate + Hydrogen

Salts made when metals react with *hydrochloric* acid are called *chlorides*.

Magnesium + Hydrochloric acid → Magnesium Chloride + Hydrogen

Reactions of acids - PART 1

Acids are neutralised by **alkalis** (**eg: soluble metal hydroxides**) and **bases** (**eg: insoluble metal hydroxides and metal oxides**) to produce **salts** and **water** and by **metal carbonates** to produce **salts**, **water** and **carbon dioxide**.

The salt **name** depends on the **acid** used and the **positive ions** in the **alkali, base or carbonate**.

<u>Making Soluble Salts from acids and alkalis</u>
Salts can be made by reacting an acid with an alkali.
Acid + Alkali → Salt + Water

<u>Making Soluble Salts from acids and bases</u>
Salts can be made by reacting an acid with a insoluble base.
Acid + Bases → Salt + Water

<u>Making Soluble Salts from acids and metal carbonates</u>
Salts can be made by reacting an acid with a metal carbonate.
Acid + Metal carbonate → Salt + Water + Carbon dioxide

Salts are made of positive metal ions (or ammonia ions - NH_4^+) and a negative ion from the acid. Like all ionic compounds, salts have **no overall charge**, so once you know the charges on the ions, you can work out the **formula**.
Example: **magnesium sulfate** is **$MgSO_4$**

ion	formula	ion	formula
Group 1	Li^+ Na^+ K^+	Transition metals	Cu^{2+} Fe^{3+}
Group 2	Mg^{2+} Ca^{2+}	Group 7	F^- Cl^- Br^-
Aluminium	Al^{3+}	Nitrate	NO_3^-
Ammonium	NH_4^+	Sulphate	SO_4^{2-}

Reactions of acids - PART 2

Indicators are substances which change **colour** when you add them to acids and alkali.

Litmus goes red in acid and blue in alkali.

Universal indicator, made from many dyes is used to tell you **pH**. The scale runs from 0 (most acidic) to 14 (most alkaline). Aqueous solutions of **acids** have a pH value **less than 7**, and for **alkalis greater than 7** and anything in the middle is **neutral** (pH 7). You can use a pH meter to record the change of a pH over time.

Acids produce **hydrogen ions (H^+)** in aqueous solutions and **alkalis** produce **hydroxide ions (OH^-).** In **neutralisation** reactions between an acid and alkali, hydrogen ions react with hydroxide ions to produce **water**.

Neutralisation symbol equation:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

pH	Examples of solutions
0	Battery acid, strong hydrofluoric acid
1	Hydrochloric acid secreted by stomach lining
2	Lemon juice, gastric acid, vinegar
3	Grapefruit juice, orange juice, soda
4	Tomato juice, acid rain
5	Soft drinking water, black coffee
6	Urine, saliva
7	“Pure” water
8	Sea water
9	Baking soda
10	Great Salt Lake, milk of magnesia
11	Ammonia solution
12	Soapy water
13	Bleach, oven cleaner
14	Liquid drain cleaner

Recall it ...

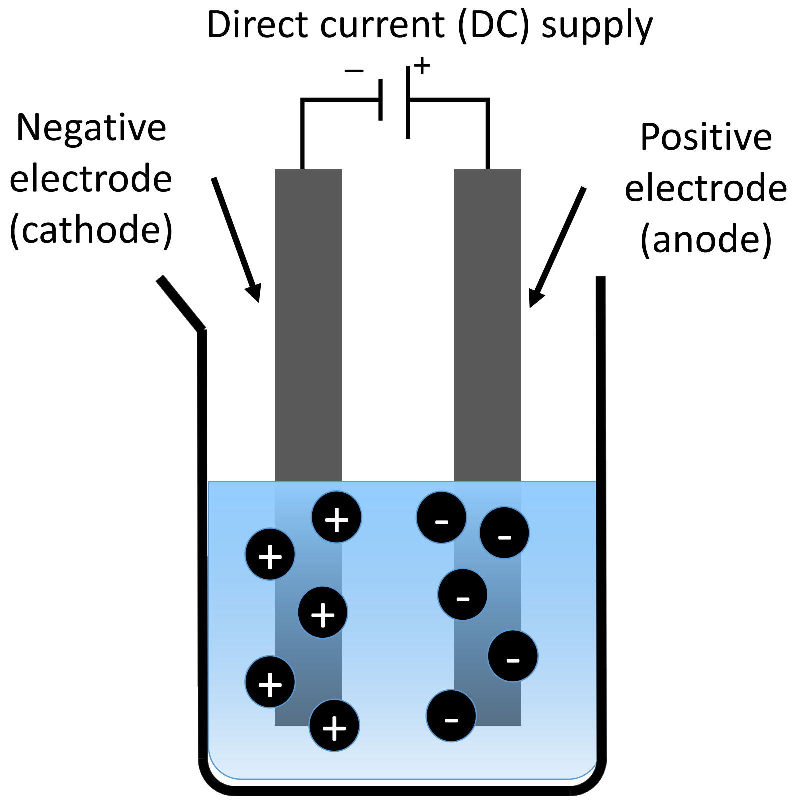
Electrolysis

Use the information in the following page(s) to answer these questions …

1) Describe what is an electrolyte?
2) What is the anode and cathode?
3) Describe what happens at the anode?
4) Describe what happens at the cathode?
5) What happens during electrolysis?
6) Describe what happens at the anode and cathode during the electrolysis of lead bromide?
7) Describe what happens at the anode and cathode during the electrolysis of sodium chloride?
8) What are the uses of the products from the electrolysis of sodium chloride?
9) Why is the electrolysis of aluminium oxide expensive?
10) What is mixed with aluminium to reduce the temperature to heat aluminium oxide?
11) Describe what is produced at the anode and cathode during the electrolysis of aluminium oxide?

Electrolysis - PART 1

When an **ionic compound** is **melted** or **dissolved in water**, the **ions** are **free** to **move** about the liquid or solution. These liquids and solutions are able to **conduct electricity** and are called **electrolytes**. Passing an **electric current** though electrolytes causes the ions to move to the electrodes.



Positive ions go to **negative** electrode (cathode) and are **reduced** (gain of electrons).

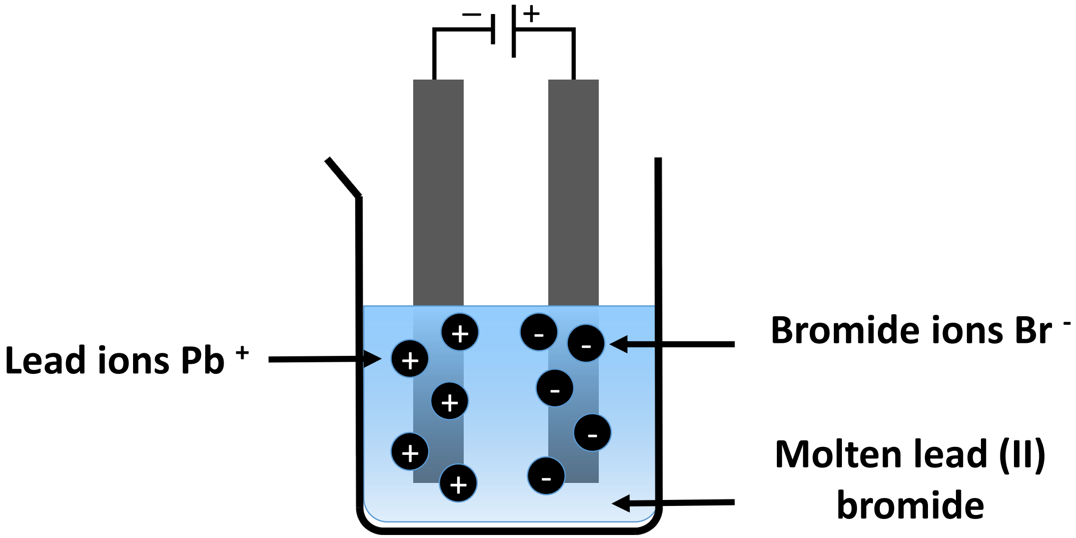
Negative ions go to the **positive** electrode (anode) and are **oxidised** (loss of electrons).

Ions are **discharged** at the electrodes producing **elements**. This is called **electrolysis**.

Electrolysis - PART 1

When an **ionic compound** is electrolysed in a **molten** state using inert electrodes, the **metal** is produced at the **cathode** and the **non-metal** is produced at the **anode**.

lead bromide → lead + bromine



The positively charged lead ions Pb + (cations) are attracted to cathode and the negatively charged bromide ions Br - are attracted to the anode.

Electrolysis - PART 1

The **ions** discharged when an **aqueous solution** is electrolysed using inert electrodes depend on the relative **reactivity** of the elements involved.

At the **negative** electrode:

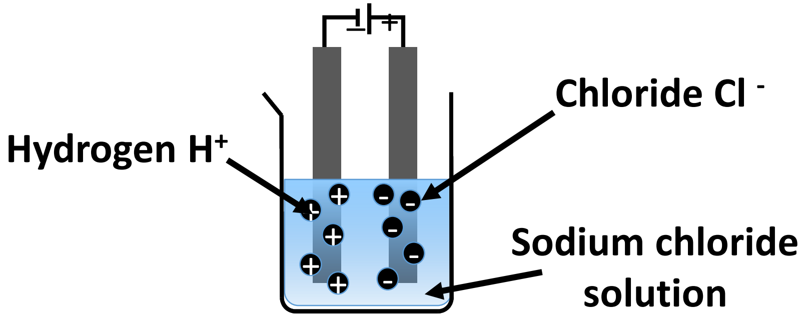
Metal will be produced on the electrode if it is **less** reactive than **hydrogen**.
Hydrogen will be produced if the metal is **more** reactive than hydrogen.

At the **positive** electrode:

Oxygen is formed at positive electrode.

If you have a **halide** ion (Cl^-, I^-, Br^-) then you will get **chlorine, bromine or iodine** formed at that electrode.
This happens because in the aqueous solution, **water molecules** break down producing **hydrogen** ions and **hydroxide** ions that are discharged.

$$H_2O(l) \rightarrow H^+(aq) + OH^-(aq)$$



sodium chloride → hydrogen + chlorine + sodium hydroxide

Uses of the products:
Chlorine: Bleach and PVC
Hydrogen: Margarine
Sodium hydroxide: Bleach and soap

Electrolysis - PART 2

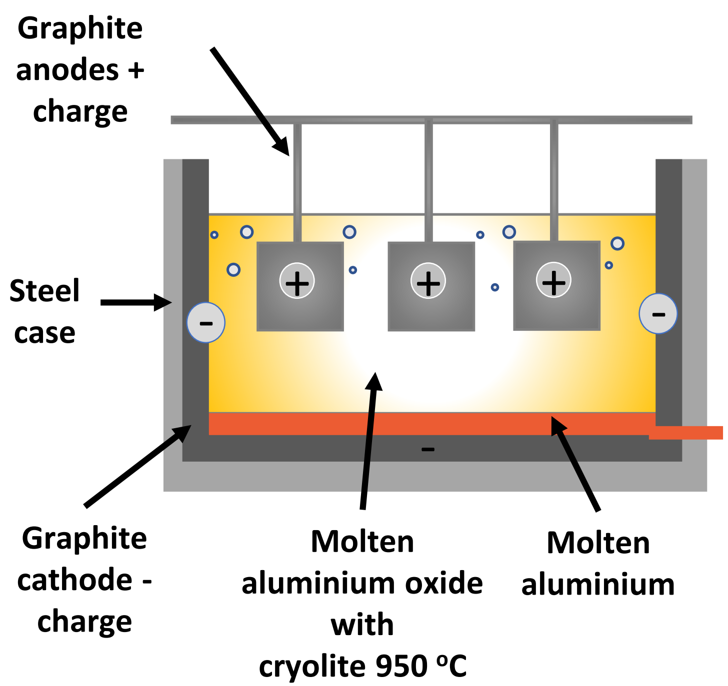
Metals can be extracted from **molten compounds** using electrolysis.

It is used if the metal is too reactive to be extracted by reduction with carbon or if the metal reacts with carbon.

Large amounts of **energy** are used in the extraction process to melt the compounds and to produce the electrical current.

Aluminum is manufactured by electrolysis of molten aluminum oxide.

Aluminium oxide → aluminium + oxygen



Aluminium oxide has a very **high melting point** so is mixed with molten **cryolite** to lower the temperature required to carry out the electrolysis.

Aluminium goes to the negative electrode and sinks to bottom.

Oxygen forms at positive electrodes.

The oxygen reacts with the carbon electrode making carbon dioxide causing damage. The electrode needs **replacing** due to this reaction.

$C + O_2 \rightarrow CO_2$