GCSE Chemistry Paper 1
Revision Notes

Name: ……………………….

Class: ……..

Atoms, Elements and Compounds

All substances are made of atoms.
An **atom is the smallest part of an element** that can exist.
Atoms of each element are **represented by a chemical symbol**,
E.g. O represents an atom of oxygen, Na represents an atom of sodium.
There **are about 100 different elements**.
Elements are shown in the periodic table.

1	2											3	4	5	6	7	0
							1 **H** hydrogen 1										4 **He** helium 2
7 **Li** lithium 3	9 **Be** beryllium 4			**Key** relative atomic mass **atomic symbol** name atomic (proton) number								11 **B** boron 5	12 **C** carbon 6	14 **N** nitrogen 7	16 **O** oxygen 8	19 **F** fluorine 9	20 **Ne** neon 10
23 **Na** sodium 11	24 **Mg** magnesium 12											27 **Al** aluminium 13	28 **Si** silicon 14	31 **P** phosphorus 15	32 **S** sulfur 16	35.5 **Cl** chlorine 17	40 **Ar** argon 18
39 **K** potassium 19	40 **Ca** calcium 20	45 **Sc** scandium 21	48 **Ti** titanium 22	51 **V** vanadium 23	52 **Cr** chromium 24	55 **Mn** manganese 25	56 **Fe** iron 26	59 **Co** cobalt 27	59 **Ni** nickel 28	63.5 **Cu** copper 29	65 **Zn** zinc 30	70 **Ga** gallium 31	73 **Ge** germanium 32	75 **As** arsenic 33	79 **Se** selenium 34	80 **Br** bromine 35	84 **Kr** krypton 36
85 **Rb** rubidium 37	88 **Sr** strontium 38	89 **Y** yttrium 39	91 **Zr** zirconium 40	93 **Nb** niobium 41	96 **Mo** molybdenum 42	[98] **Tc** technetium 43	101 **Ru** ruthenium 44	103 **Rh** rhodium 45	106 **Pd** palladium 46	108 **Ag** silver 47	112 **Cd** cadmium 48	115 **In** indium 49	119 **Sn** tin 50	122 **Sb** antimony 51	128 **Te** tellurium 52	127 **I** iodine 53	131 **Xe** xenon 54
133 **Cs** caesium 55	137 **Ba** barium 56	139 **La*** lanthanum 57	178 **Hf** hafnium 72	181 **Ta** tantalum 73	184 **W** tungsten 74	186 **Re** rhenium 75	190 **Os** osmium 76	192 **Ir** iridium 77	195 **Pt** platinum 78	197 **Au** gold 79	201 **Hg** mercury 80	204 **Tl** thallium 81	207 **Pb** lead 82	209 **Bi** bismuth 83	[209] **Po** polonium 84	[210] **At** astatine 85	[222] **Rn** radon 86
[223] **Fr** francium 87	[226] **Ra** radium 88	[227] **Ac*** actinium 89	[261] **Rf** rutherfordium 104	[262] **Db** dubnium 105	[266] **Sg** seaborgium 106	[264] **Bh** bohrium 107	[277] **Hs** hassium 108	[268] **Mt** meitnerium 109	[271] **Ds** darmstadtium 110	[272] **Rg** roentgenium 111	Elements with atomic numbers 112 – 116 have been reported but not fully authenticated						

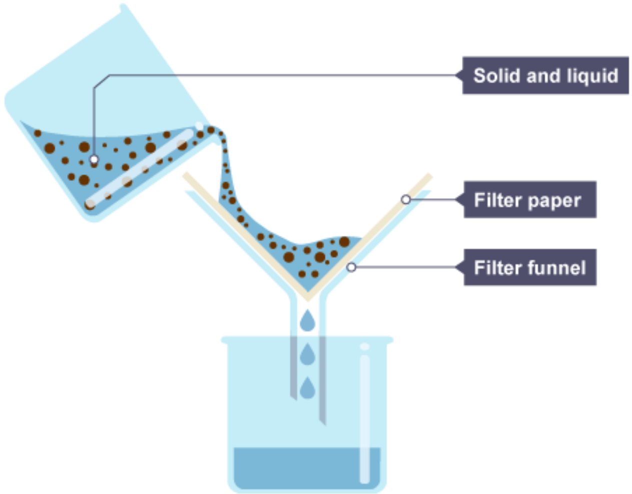
Compounds are formed from elements by chemical reactions.
Chemical reactions always involve the formation of one or more new substances, and often involve a detectable energy change.
Compounds contain two or more elements chemically combined in fixed proportions (a certain amount of one element with a certain amount of another element) and can be **represented by formulae** using the symbols of the atoms from which they were formed.
Compounds can only be separated into elements by chemical reactions.
Chemical reactions can be represented by word equations or equations using symbols and formulae.

Mixtures

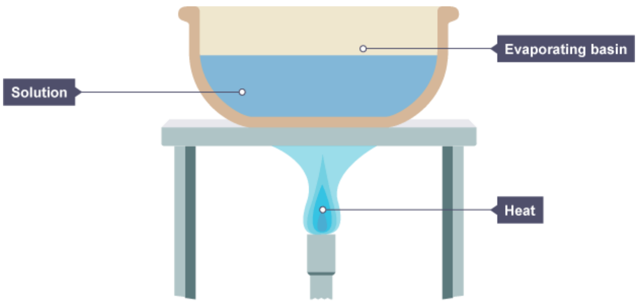
A mixture consists of **two or more elements or compounds not chemically combined** together.
The **chemical properties of each substance** in the mixture **are unchanged**.

Mixtures **can be separated by physical processes** such as filtration, crystallisation, simple distillation, fractional distillation and chromatography. These physical processes **do not involve chemical reactions** and **no new substances are made**.

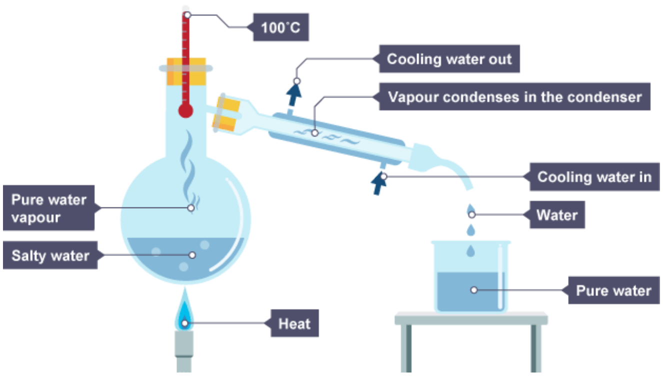
Filtration



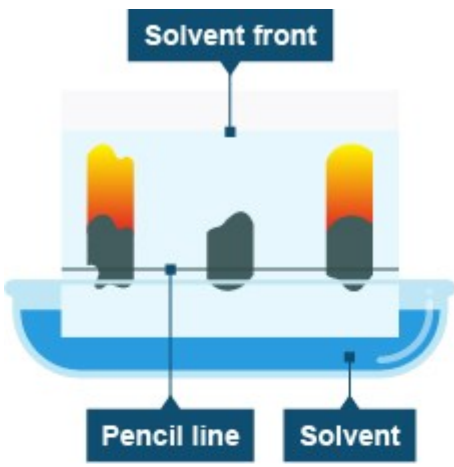
Crystallisation (Evaporation)



Simple distillation



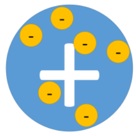
Chromatography



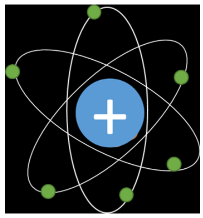
The development of the model of the atom

New experimental evidence may lead to a scientific model being changed or replaced.

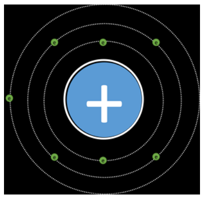
Before the discovery of the electron, **atoms were thought to be tiny spheres that could not be divided**.

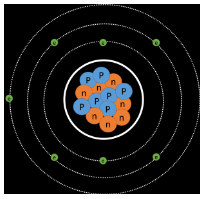


The discovery of the electron led to the **plum pudding model** of the atom. The plum pudding model suggested that the atom is a **ball of positive charge** with **negative electrons** embedded in it.



The results from the **alpha particle scattering experiment** led to the conclusion that the **mass of an atom was concentrated at the centre** (nucleus) and that the **nucleus was positively charged**. This nuclear model replaced the plum pudding model.





Niels Bohr adapted the nuclear model by suggesting that **electrons orbit the nucleus at specific distances**. The theoretical calculations of Bohr agreed with experimental observations.

Subatomic Particles

The relative electrical charges of the particles in atoms are:

Name of particle	Relative charge
Proton	+1
Neutron	0
Electron	–1

In an atom, the **number of electrons is equal to the number of protons** in the nucleus.
Atoms have **no overall electrical charge**.

The **number of protons** in an atom of an element is its **atomic number**.
All **atoms** of a **particular element** have the **same number of protons**.
Atoms of **different elements** have **different numbers of protons**.

Atoms are very small, having a radius of about 0.1 nm (1×10^{-10} m).
The radius of a nucleus is less than 1/10 000 of that of the atom (about 1×10^{-14} m).
Almost all of the mass of an atom is in the nucleus.
The relative masses of protons, neutrons and electrons are:

Name of particle	Relative mass
Proton	1
Neutron	1
Electron	Very small

The **sum of the protons and neutrons** in an atom is its **mass number**.

Atoms of the **same element can have different numbers of neutrons**; these atoms are **called isotopes** of that element.

Atoms can be represented as shown in this example:

(Mass number) (Atomic number) $^{23}_{11}Na$

An **ion** is an atom which has **lost or gained electrons**.

Relative atomic mass

The relative atomic mass of an element is an average value that takes account of the abundance of the isotopes of the element.
(You're figuring out the average mass of an atom)

Example:

75% of chlorine atoms are ^{35}Cl
25% of chlorine atoms are ^{37}Cl

In 100 chlorine atoms 75 will be ^{35}Cl and 25 will be ^{37}Cl

The total A_r (Relative atomic mass) for these chlorine atoms will be:

$(75 \times 35) + (25 \times 37) = 2625 + 925 = 3550.$

So the average A_r for chlorine is $3550 \div 100 = 35.5$.

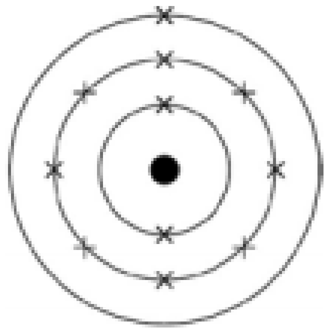
Electronic Structure

The **electrons** in an atom **occupy** the **lowest available energy levels** (innermost available shells).

The electronic structure of an atom can be represented by numbers or by a diagram.

The first energy level can hold 2 electrons, the second 8 electrons, the third 8 electrons and the fourth 2 electrons.

For example, the electronic structure of sodium is 2,8,1 or



The Periodic Table

The **elements in the periodic table** are **arranged in order of atomic** (proton) **number** and so that **elements with similar properties are in columns**, known as groups.

The table is called a periodic table because **similar properties occur at regular intervals**.

Elements in the **same group** in the periodic table have the same **number of electrons in their outer shell** (outer electrons) and this gives them **similar chemical properties**.

Development of the Periodic Table

Before the discovery of protons, neutrons and electrons, scientists attempted to **classify the elements by arranging them in order of their atomic weights**.

The **early periodic tables were incomplete** and some elements were placed in inappropriate groups if the strict order of atomic weights was followed.

Mendeleev overcame some of the problems by leaving **gaps for elements that he thought had not been discovered** and in some places changed the order based on atomic weights.

Elements with properties **predicted by Mendeleev were discovered** and filled the gaps.

Knowledge of **isotopes** (elements with different numbers of neutrons) **made it possible to explain why the order based on atomic weights was not always correct**.

Metals and Non-Metals

Elements that **react** to **form positive ions** are **metals**.
Elements that **do not form positive ions** are **non-metals**.

The majority of elements are metals.

Metals are found to the left and towards the bottom of the periodic table.
Non-metals are found towards the right and top of the periodic table.

Metal	Metalloid	Nonmetal

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-Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac-Lr															

La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

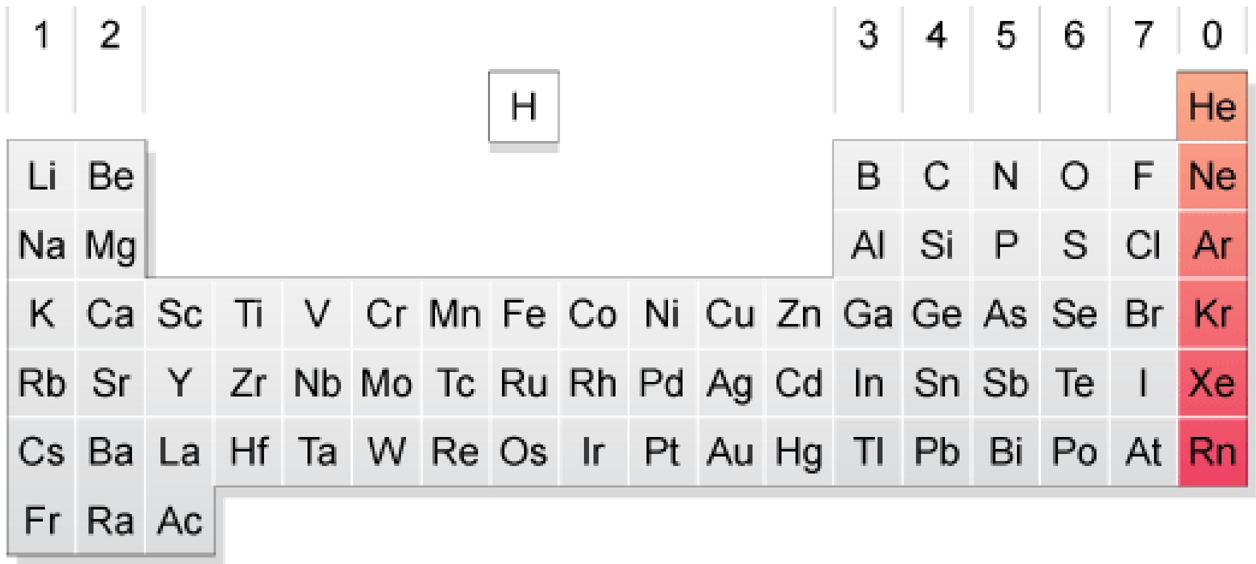
Group 0

The elements in Group 0 of the periodic table are called the **noble gases**.

They are **unreactive** and do not easily form molecules because their atoms have **stable arrangements of electrons**.

The **noble gases have eight electrons in their outer shell**, except for helium, which has only two electrons.

The **boiling points of the noble gases increase** with **increasing relative atomic mass** (going down the group).

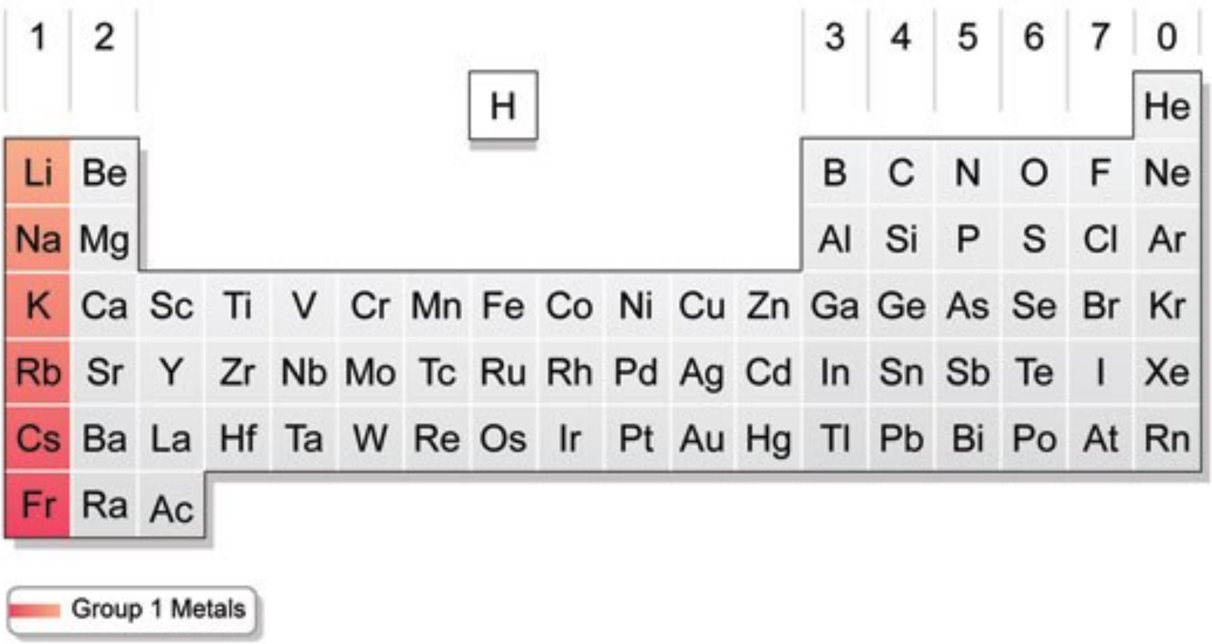


Group 1

The elements in Group 1 of the periodic table are known as the **alkali metals** and have characteristic properties because of the **single electron in their outer shell**.

In Group 1, the **reactivity of the elements increases going down the group**.

	Reaction with...		
	oxygen	**chlorine**	**water**
Li	Burns with a red flame.	White powder	Fizzes steadily, gradually disappears
Na	Burns with a yellow-orange flame.	Bright yellow flame. Gives white powder.	Fizzes rapidly, melts into a ball and disappears quickly
K	Burns with a lilac flame.	Reacts more violently than Sodium.	Ignites with sparks and a lilac flame, disappears very quickly



Group 7

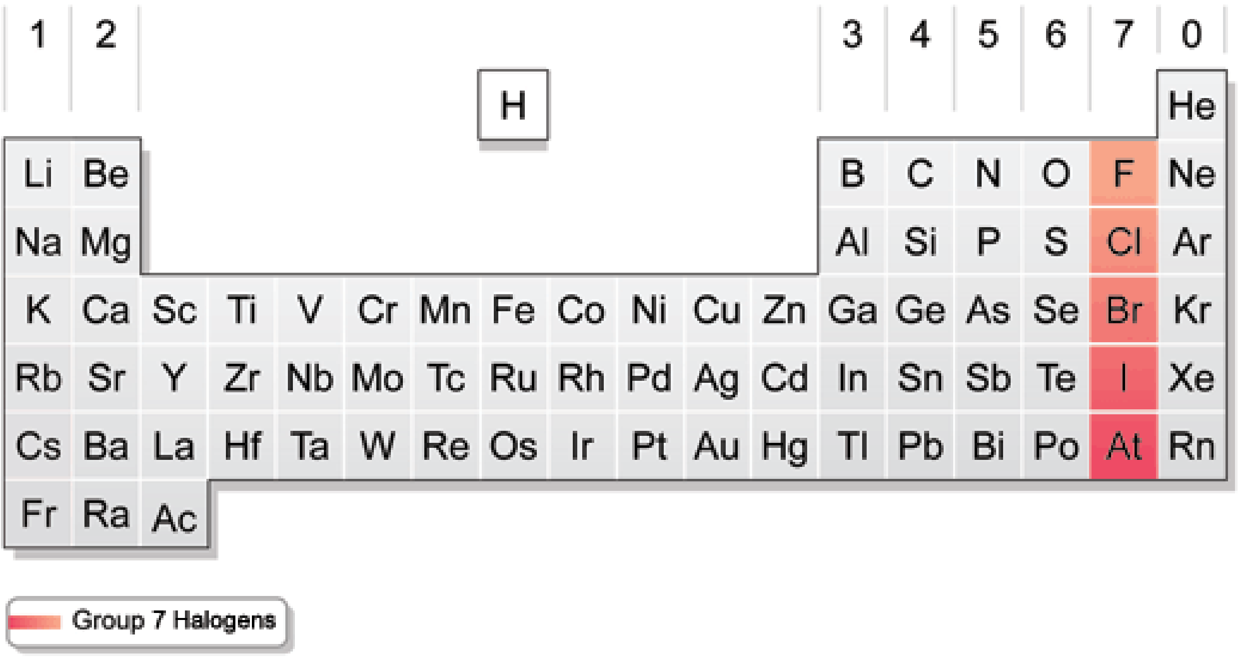
The elements in Group 7 of the periodic table are known as the **halogens** and have similar reactions because they all have **seven electrons** in their outer shell.

The halogens are **non-metals** and consist of molecules made of **pairs of atoms**.

In Group 7, the **further down the group** an element is the **higher** its relative molecular mass, **melting point and boiling point**.

In Group 7, the **reactivity of the elements decreases going down the group**.

A more reactive halogen can displace a less reactive halogen from an aqueous solution of its salt (the more reactive element joins with the 'chloride' 'bromide' etc. leaving the less reactive element on its own).



Properties of Transition metals

Comparison with Group 1 metals

The **transition elements** are **metals with similar properties** which are different from those of the elements in Group 1.

Properties of Group 1 and transition metals:

Group 1	**Group 1 and transition**	**Transition elements**
Low density	Electrical conductors	Coloured compounds
Low melting point	Thermal conductors	Good Catalysts
Soft	Ductile, malleable	Ions with different charges
Make 1+ ions	Lustrous	Hard, strong
Very reactive with water and oxygen	Sonorous	React little with water
	React with halogens	Some react with O_2

Typical properties

Many transition elements have ions with different charges, form coloured compounds and are useful as catalysts.

Transition metal or transition metal compound	**Use**	**Why the transition metals or its compound is used in this way**
Chromium	Chromium plating on cars	It is less chemically reactive so stays shiny for longer.
Manganese compounds	To make panels for stained-glass windows	It forms coloured compounds
Iron	In the Haber process to make ammonia	It is a catalyst
Cobalt	Aircraft engines	It is hard
Nickel	To make coins	It is hard and doesn't get worn away
Copper	To make water pipes	It is unreactive

Chemical Bonds

There are three types of strong chemical bonds: ionic, covalent and metallic.

For **ionic bonding** the particles are **oppositely charged ions**. Ionic bonding occurs in compounds formed from **metals combined with non-metals**.

For **covalent bonding** the particles are atoms which **share pairs of electrons**. Covalent bonding occurs in most **non-metallic elements** and in compounds **of non-metals.**

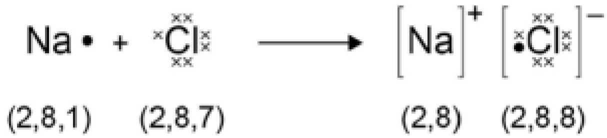
For **metallic bonding** the particles are atoms which **share delocalised electrons**. Metallic bonding occurs in **metallic elements** and alloys.

Ionic Bonding

When a **metal atom reacts with a non-metal atom electrons** in the outer shell of the metal atom **are transferred**.
Metal atoms lose electrons to become **positively charged** ions. **Non-metal atoms** gain electrons to become **negatively charged** ions.

The **ions produced by metals in Groups 1 and 2 and by non-metals in Groups 6 and 7** have the electronic structure of a noble gas (Group 0) (**a full outer shell**).

The electron transfer during the formation of an **ionic compound** can be represented by a **dot and cross diagram**,
E.g. for sodium chloride.

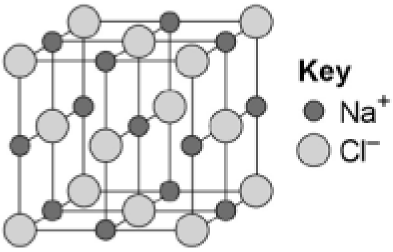


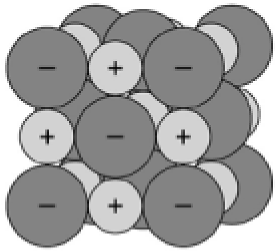
The charge on the ions produced by metals in Groups 1 and 2 and by non-metals in Groups 6 and 7 relates to the group number of the element in the periodic table.
Group 1 non metals form +1 ions
Group 2 non metals form +2 ions
Group 6 metals form -2 ions
Group 7 metals form -1 ions

Ionic Compounds

An **ionic compound is a giant structure of ions**.
Ionic compounds are held together by strong **electrostatic forces of attraction** between **oppositely charged ions**. These **forces act in all directions** in the lattice and this is called ionic bonding.

The structure of sodium chloride can be represented in the following forms:





Covalent Bonding

When atoms **share pairs of electrons**, they **form covalent bonds**.
These bonds between atoms are strong.

Covalently bonded substances may consist of small molecules.

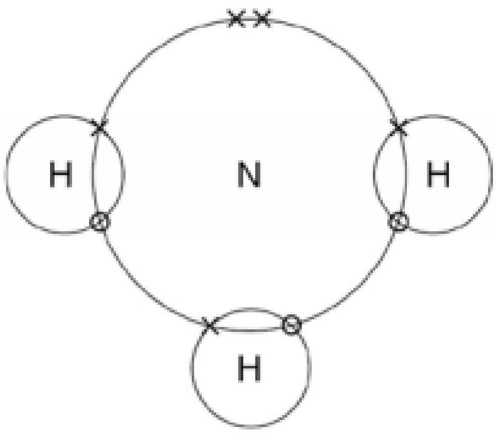
Some covalently bonded substances consist of small molecules e.g. O_2

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

Some covalently bonded substances have giant covalent structures, such as diamond and silicon dioxide.

The covalent bonds in molecules and giant structures can be represented in the following forms:

For ammonia (NH_3)



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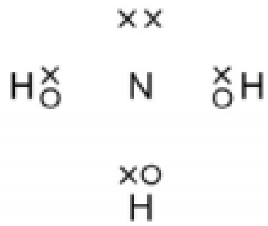

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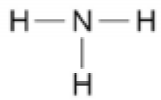

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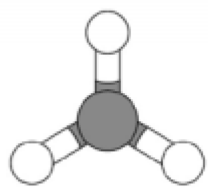
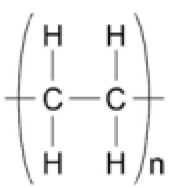


Polymers can be represented in the form:



poly(ethene)

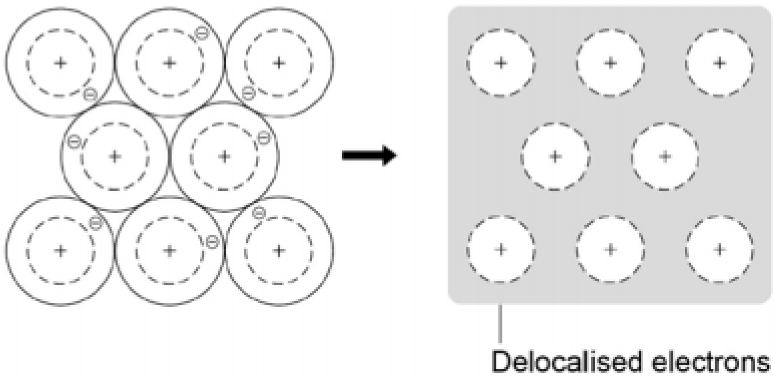
where n is a large number.

Metallic Bonding

Metals consist of **giant structures** of atoms arranged in a **regular pattern**.

The **electrons** in the outer shell of metal atoms are **delocalised** and so are **free to move** through the whole structure. The **sharing of delocalised electrons** gives rise to **strong metallic bonds**.

The bonding in metals may be represented in the following form:



States of Matter

The three states of matter are **solid, liquid and gas**.
Melting and freezing take place at the melting point, boiling and condensing take place at the boiling point.

The three states of matter can be represented by a simple model. In this model, particles are represented by small solid spheres.

Particle theory can help to explain melting, boiling, freezing and condensing. The **amount of energy** needed to change state from solid to liquid and from liquid to gas **depends on the strength of the forces between the particles of the substance**. The **stronger the forces** between the particles the higher **the melting point and boiling point** of the substance.

In chemical equations, the three states of matter are shown as (s) solid, (l) liquid and (g) gas, with (aq) for aqueous solutions.

Properties of ionic compounds

Ionic compounds have **regular structures** (giant ionic lattices) in which there are **strong electrostatic forces** of attraction in all directions between **oppositely charged ions**.

These compounds have **high melting points and high boiling points** because of the **large amounts of energy needed to break the many strong bonds**.

When **melted or dissolved in water**, ionic compounds **conduct electricity** because the **ions are free to move** and so charge can flow.

Properties of Small Molecules

Substances that consist of **small molecules are usually gases or liquids** that have **relatively low melting points and boiling points**.

These substances have only **weak forces** between the molecules (intermolecular forces). It is these intermolecular forces that are overcome, not the covalent bonds, when the substance melts or boils.

The **intermolecular forces increase with the size of the molecules**, so larger **molecules have higher melting and boiling points**.

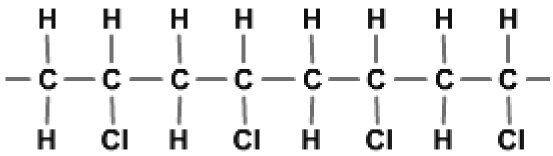
These substances **do not conduct electricity** because the molecules do not have an overall electric charge.

Polymers

Polymers have **very large molecules**. The atoms in the polymer molecules are **linked to other atoms by strong covalent bonds**.

The **intermolecular forces** between polymer molecules are **relatively strong** and so these substances are **solids at room temperature**. Students should be able to recognise polymers from diagrams showing their bonding and structure.

A polymer is a long chain of repeating molecules (monomers).

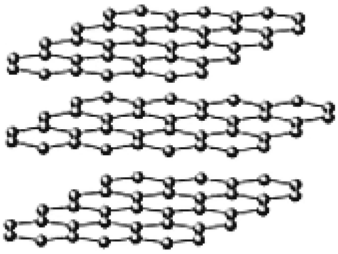


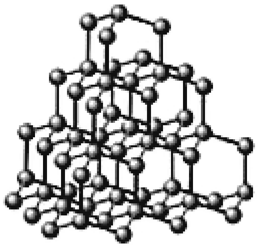
a section of poly (chloroethene) PVC

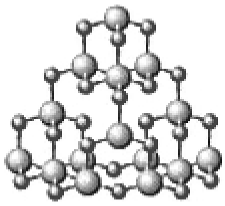
Giant Covalent Structures

Substances that consist of giant covalent structures are **solids** with **very high melting points**. All of the atoms in these structures are **linked to other atoms by strong covalent bonds**. These **bonds must be overcome to melt or boil these substances.**

Diamond and graphite (forms of carbon) and silicon dioxide (silica) are examples of giant covalent structures.







Properties of Metals and Alloys

Metals have **giant structures** of atoms with **strong metallic bonding**.

This means that most metals have **high melting and boiling points**.

In pure metals, **atoms are arranged in layers**, which **allows metals to be bent and shaped**.

Pure metals are too soft for many uses and so are **mixed with other metals** to make **alloys which are harder**.

Metal element structure

Alloy structure – harder than original metal

Metals as conductors

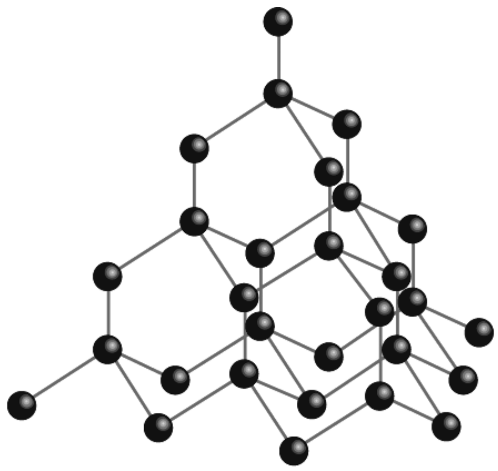
Metals are **good conductors of electricity** because the **delocalised electrons** in the metal **carry electrical charge** through the metal.

Metals are **good conductors of thermal energy** because **energy is transferred** by the **delocalised electrons.**

Structure and bonding of carbon

Diamond

In diamond, each **carbon atom forms four covalent bonds with other carbon atoms** in a **giant covalent structure**, so diamond is **very hard**, has a **very high melting point** and **does not conduct electricity**.

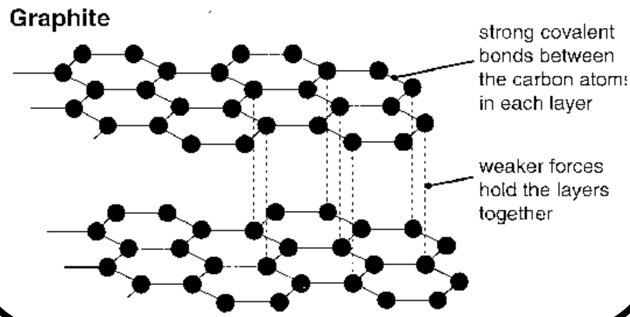


Graphite

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

In graphite, one electron from each carbon atom is delocalised (meaning it **can conduct electricity**).

The layers can separate easily due to the weak bonds between layers.



Fullerenes

Fullerenes are **molecules of carbon atoms with hollow shapes**.

The structure of fullerenes is based on **hexagonal rings of carbon atoms** but they may also contain rings with five or seven carbon atoms.

The first fullerene to be discovered was Buckminsterfullerene (C60) which has a spherical shape.

Carbon nanotubes are cylindrical fullerenes with very high length to diameter ratios.

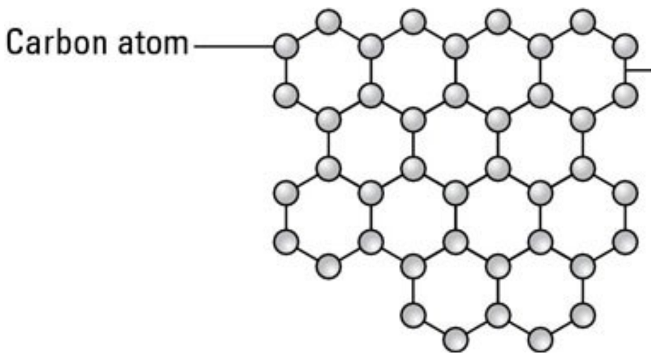
Their properties make them useful for nanotechnology, electronics and materials.

Graphene

Graphene is a **single layer of graphite** and has properties that make it useful in electronics and composites.

It can conduct electricity.

It is very strong.



Uses of Nanoparticles

Nanoparticles are very small particles that are between 1 and 100 nm in size. They can also be described as being ultra-fine particles.

Fine particles have diameters between 100 and 2500nm.

Coarse particles have diameters between 10μm and 2.5μm (2500nm), coarse particles are also sometimes called dust.

Remember **1nm = 0.000,000,001m = 1 x 10^{-9}m**

Atoms have a radius of 1×10^{-10}m. **Atoms are smaller than nanoparticles.**

Silver nanoparticles are added to the fabric used to make socks. These nanoparticles **kill bacteria** that could cause unpleasant odours.

Nanoparticles of iron oxide are used to **remove poisonous chemicals** including arsenic from water wells.

Nanoparticles of zinc oxide are used in some **sun-screens** and are also used to **protect** textiles and wood **from exposure to harmful ultraviolet rays**.

In the future, **nanoparticles could be developed** to **deliver medicines** to tumours inside the body or to break up clusters of bacteria that are causing chronic bacterial infections.

Conservation of mass

The law of conservation of mass states that no **atoms are lost or made during a chemical reaction** so **the mass of the products equals the mass of the reactants**. This means that **chemical reactions can be represented by symbol equations** which are **balanced in terms of the numbers of atoms of each element involved on both sides** of the equation.

H_2O means there are 2 atoms of hydrogen and 1 atom of oxygen.

$C_6H_{12}O_6$ means there are 6 atoms of carbon, 12 atoms of hydrogen and 6 atoms of oxygen.

3 MgO means there are 3 whole molecules of magnesium oxide.

When balancing equations **we can change the large number in front**, but **not the small number after the symbol**.

Some reactions may appear to involve a change in mass but this can usually be explained because **a reactant or product is a gas** and its mass has not been taken into account.

For example:

- when a metal reacts with oxygen the mass of the oxide produced is greater than the mass of the metal
- in thermal decompositions of metal carbonates carbon dioxide is produced and escapes into the atmosphere leaving the metal oxide as the only solid product.

Relative Formula Mass (M_r)

The **relative formula mass** (M_r) of a compound is the sum **of the relative atomic masses (A_r) of the atoms** in the numbers shown in the formula.

In a balanced chemical equation, the **sum of the relative formula masses of the reactants** in the quantities shown equals the **sum of the relative formula masses of the products** in the quantities shown.

Example:

MgO

Relative atomic mass of Magnesium = 24

Relative atomic mass of Oxygen = 16

Relative formula mass = 16 + 24 = 40

Example:

$C_6H_{12}O_6$

Relative atomic mass of Carbon = 12

Relative atomic mass of Hydrogen = 1

Relative atomic mass of Oxygen = 16

Relative formula mass = (12 x 6) + (1 x 12) + (16 x 6) = 180

Moles

Chemical amounts are measured in moles.
The symbol for the unit mole is mol.
The **mass of one mole** of a substance **in grams** is the **same** number as its **relative formula mass**.

One mole of a substance contains the same number of the stated particles, atoms, molecules or ions as one mole of any other substance. The number of atoms, molecules or ions in a mole of a given substance is the Avogadro constant. The value of the Avogadro constant is 6.02×10^{23} per mole.

One mole of any substance contains 6.02×10^{23} particles, atoms, molecules or ions.

Using moles to balance equations

Example: 48g of Magnesium is reacted with 32g of Oxygen to give 80g of Magnesium oxide.

Step 1: Write out equation

$$Mg + O_2 \rightarrow MgO$$

Step 2: Calculate the mass of 1 mole of each molecule, then the number of moles, by dividing the given mass by the mass of a mole

Mass of 1 mole of Mg = 24g
Mass of 1 mole of O_2 = 16 x 2 = 32g
Mass of 1 mole of MgO = 24 + 16 = 40g

Step 3: Simplify the ratio of the number of moles.

$$(48 \div 24):(32 \div 32):(80 \div 40)$$

$$2:1:2$$

Step 4: Put numbers from ratio into equation.

$$2Mg + O_2 \rightarrow 2MgO$$

Limiting Reactants

In a chemical reaction involving two reactants, it is common to use an excess (more than we need) of one of the reactants to ensure that all of the other reactant is used.

The **reactant that is completely used up is called the limiting reactant** because it **limits the amount of products**.

Step 1: Balance the equation.
Step 2: Find mass of a mole of the reactants.
Step 3: Divide mass of reactant (from question) by mass of a mole (from step 2)
Step 4: Multiply all mass of all moles by number found in step 3.

Example:
If you have 4.8g of magnesium how much hydrochloric acid can it be reacted with to make the maximum possible amount of magnesium chloride and hydrogen?

Step 1:Balance the equation

$$2Mg + 2HCl \rightarrow 2MgCl + H_2$$

Step 2: Find the mass of a mole of the reactants.

Mass of 1 mole of 2Mg = 2 x 24 = 48g

Mass of 1 mole of 2HCl = 2 x (1 + 35.5) = 73g

Step 3: Divide mass of reactant (from question) by mass of a mole (from step 2)

4.8g ÷ 48g = 0.1

Step 4: Multiply all mass of all moles by number found in step 3.

Mass of Magnesium = 48 x 0.1 = 4.8g

Mass of Hydrochloric acid = 73 x 0.1 = 7.3g

Concentration of Solutions

$$\text{Concentration (g/m}^3\text{)} = \frac{\text{Mass of solute (g)}}{\text{Volume of solvent (m}^3\text{)}}$$

(Higher: Sometimes we replace mass in the equation above with number of moles)

A dm^3 is a volume 10cm x 10cm x 10cm.

Example:

50cm^3 of vinegar contains 1.25g of ethanoic acid. Give its concentration in g/cm^3

Concentration = 1.25 ÷ 50 = 0.025g/m^3

Example:

What mass of ethanoic acid would there be in 2m^3 of this vinegar?

Mass = concentration x volume

Mass = 0.025 x 2 = 0.05 g

Percentage Yield

Even though no atoms are gained or lost in a chemical reaction, it is not always possible to obtain the calculated amount of a product because sometimes in reactions:

- Some product may be left in containers
- Some of the reactants may react in unexpected ways.

The equation to calculate the percentage yield is:

$$\%\ \text{yield} = \frac{\text{mass of product made}}{\text{Max possible mass of product}} \times 100$$

Example:

In a reaction 82g of calcium oxide is made, the maximum that could possibly be made is 112g. Calculate the percentage yield.

% yield = 82 ÷ 112 = 73%

Atom Economy

Atom economy is a measure of how much of the starting materials (reactants) end up as the product that we want.

$$\frac{M_r \text{ of wanted product}}{M_r \text{ of all reactants}} \times 100$$

Example:

Calcium carbonate ($CaCO_3$) decomposes to make CaO and CO_2, if calcium oxide is the desired product, what is the atom economy?

$$CaCO_3 \rightarrow CaO + CO_2$$

Relative formula mass of Calcium oxide = 40 + 16 = 56

Relative formula mass of Calcium carbonate = 40 + 12 + (16 x 3) = 100

(56 ÷ 100) x 100 = 56%

Amounts of substances and Volumes of Gases

Equal amounts in moles of gases occupy the same volume under the same conditions of temperature and pressure.

The **volume** of **one mole** of **any gas at room temperature and pressure** (20°C and 1 atmosphere pressure) is **24 dm^3** .

The volumes of gaseous reactants and products can be calculated from the balanced equation for the reaction, because **the volume of 1 mole is 24 dm^3 (at room temperature and pressure).**

Metal oxides

Metals react with oxygen to produce metal oxides. The reactions are oxidation reactions because the metals gain oxygen.

E.g. Magnesium + Oxygen → Magnesium oxide

OIL RIG

Oxidation is loss (of oxygen) **Reduction is gain (of oxygen)**

Oxidation and Reduction in terms of electrons (HT)

Oxidation is the **loss of electrons.**

Reduction is the **gain of electrons**.

The Reactivity Series

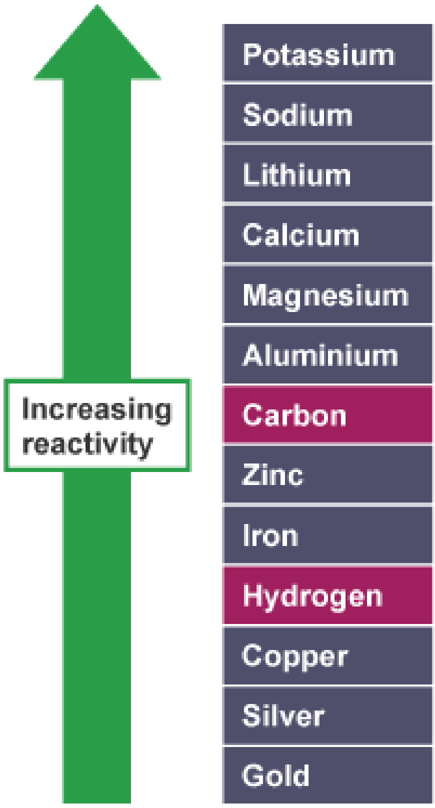
When metals react with other substances the metal atoms form positive ions.

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

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

The metals potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper can be put in order of their reactivity from their reactions with water and dilute acids.

The non-metals hydrogen and carbon are often included in the reactivity series.



A **more reactive metal** can **displace** a **less reactive** metal from a compound.

The more reactive metal can 'steal' the other part of the compound from a less reactive metal.

E.g. Zinc chloride + Aluminium → Aluminium chloride + Zinc

Extraction of metals and reduction

Unreactive metals such as gold are found in the Earth as the metal itself but **most metals are found as compounds that require chemical reactions to extract the metal**.

Metals less reactive than carbon can be **extracted** from their oxides **by reduction with carbon.**

Reduction involves the loss of oxygen.

Reactions of acids

Acids react with some metals to produce a salt and hydrogen.

Acid + Metal → Salt + Hydrogen

Hydrochloric acid makes **chloride** salts
Sulphuric acid makes **sulphate** salts

Example:

Magnesium + hydrochloric acid → Magnesium chloride + Hydrogen

Example:

Zinc + Sulphuric acid → Zinc sulphate + Hydrogen

(HT) In these reactions electrons are lost or gained, therefore these can be called redox reactions.

Acids are neutralised by alkalis (soluble metal hydroxides) and bases (insoluble metal hydroxides and metal oxides). When acids are neutralised they produce a salt and water.

Acid + Alkali → Salt + Water
Acid + Base → Salt + Water

Acids can also be neutralised by metal carbonates to produce a salt, water and carbon dioxide.

Acid + Metal carbonate → Salt + Water + Carbon dioxide

As we said before:
Hydrochloric acid makes **chloride** salts
Sulphuric acid makes **sulphate** salts
Nitric acid makes **nitrate** salts

Soluble Salts

Soluble salts (salts that will dissolve in water) can be made from **acids** by **reacting** them with solid insoluble substances, such as **metals, metal oxides, hydroxides or carbonates**.

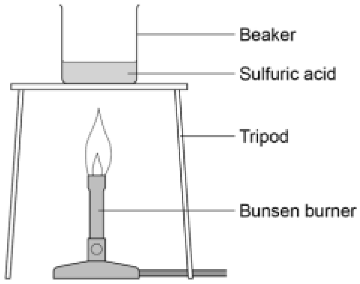
The **solid is added to the acid** until **no more reacts** and the **excess solid is filtered off** to produce a solution of the salt.

Salt solutions can be crystallised (evaporated) to produce solid salts.

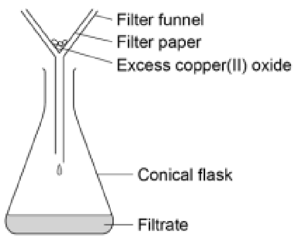
Required Practical 1: Making Salts

Method

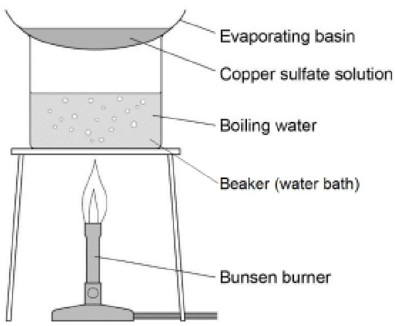
1. Measure 40 cm3 sulfuric acid and put it into the 100 cm^3 beaker.
2. Set up the Bunsen burner, tripod, gauze and heatproof mat. Put the beaker on the gauze and heat the acid gently until it is almost boiling. Turn off the Bunsen burner.



3. Remove the glass beaker from the tripod. Use the spatula to add a small amount of copper oxide powder to the hot acid. Stir with the glass rod. The copper oxide will disappear and the solution will turn clear blue.
4. Add some more copper oxide and stir again.
5. Keep adding the copper oxide until some of it remains after stirring.
6. Allow the apparatus to cool completely.
7. Set up the filter funnel and paper over the conical flask. Filter the contents of the beaker.
8. Pour the filtrate from the conical flask into the evaporating basin.



9. Set up a water bath using the 250 cm3 beaker on the tripod and gauze.
10. Evaporate the filtrate gently using the water bath.



11. When crystals start to form, stop heating the water bath.
12. Pour the remaining solution into the crystallising dish.
13. Leave the crystallising dish in a cool place for at least 24 hours.

The pH scale and neutralisation

Acids produce **hydrogen ions (**H^+) in aqueous solutions.
Aqueous solutions of alkalis contain **hydroxide ions** (OH^-).

The **pH scale, from 0 to 14, is a measure of the acidity or alkalinity of a solution**, and can be measured using universal indicator or a pH probe.

A solution with **pH 7 is neutral**.

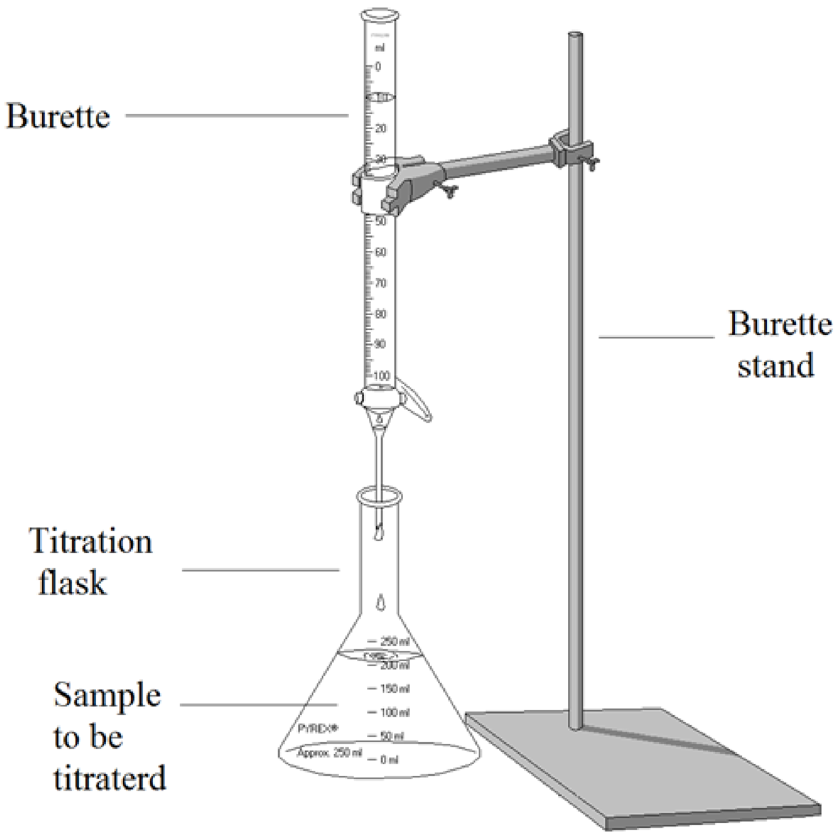
Aqueous solutions of acids have pH values of less than 7 and aqueous solutions of alkalis have pH values greater than 7.

In neutralisation reactions between an acid and an alkali, **hydrogen ions react with hydroxide ions to produce water.**
This reaction can be represented by the equation:

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

Titrations

1. Use the pipette and pipette filler to put exactly 25 cm^3 sodium hydroxide solution into the conical flask.
2. Put the flask on a white tile.
3. Clamp the burette vertically in the clamp stand. There should be just enough room underneath for the conical flask and tile.
4. Close the burette tap. 5
5. Use the small funnel to carefully fill the burette with dilute sulfuric acid. Before it completely fills put a small beaker underneath the tap, gently open it to allow acid to fill the tap, before closing again and filling the burette to the 0.00 cm3 line. Remove the funnel.
6. Put 5–10 drops of phenolpthalein indicator into the conical flask. Swirl the flask to mix and put under the burette on top of the tile. The contents of the flask will go pink.
7. Carefully open the burette tap so that 10 cm^3 sulfuric acid slowly flows into the flask. Constantly swirl the flask when adding the acid. Then add the acid drop by drop until you see a permanent colour change from pink to colourless in the flask. You need to be able to shut the tap immediately after a single drop of acid causes the colour to become permanently colourless.
8. Read the burette scale carefully and record the volume of acid you added
9. Repeat steps 1-8 twice more recording your results each time. From your 3 results then calculate an average.



Strong and Weak Acids

A **strong acid** is **completely ionised** in aqueous solution.
Examples of strong acids are hydrochloric, nitric and sulfuric acids.

A **weak acid** is only **partially ionised** in aqueous solution.
Examples of weak acids are ethanoic, citric and carbonic acids.

For a given concentration of aqueous solutions, the **stronger an acid, the lower the pH**.

As the **pH decreases by one unit**, the hydrogen ion concentration of the solution increases by a factor of 10. (**There are 10x more hydrogen ions**)

Electrolysis

When an ionic compound is melted or dissolved in water, the **ions are free to move about within the liquid or solution**.

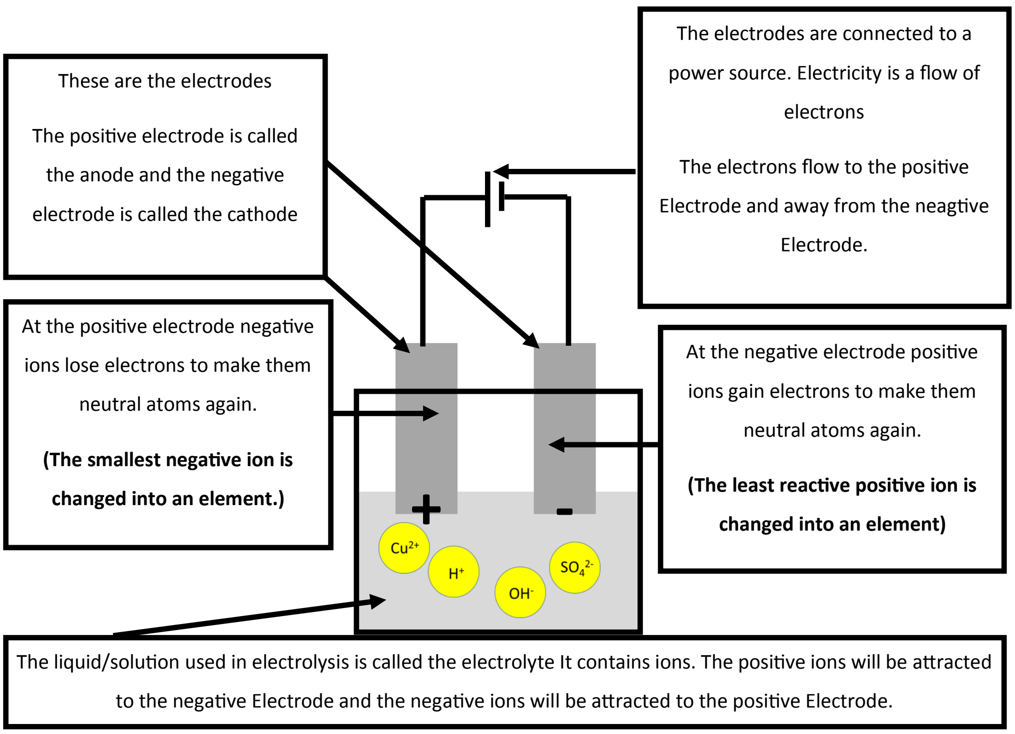
These liquids and solutions are **able to conduct electricity** and are **called electrolytes**.

Passing an **electric current through electrolytes causes the ions to move to the electrodes**.

Positively charged ions move to the **negative electrode** (the cathode), and negatively **charged ions** move to the **positive electrode** (the anode).

Ions are discharged at the electrodes (gain or lose electrons) **producing elements**.

This process is called electrolysis.



At the **negative electrode** the **least reactive of hydrogen or the metal** is produced.

At the **positive electrode** the **smallest negative ion** is turned into elements. (Hydroxide ions are turned into oxygen gas)

Electrolysis of molten ionic compounds

When a simple ionic compound (eg lead bromide) is electrolysed in the molten state using inert electrodes, the **metal** (lead) is **produced at the cathode** and the **non-metal** (bromine) is **produced at the anode**.

Electrolysis of aqueous solutions

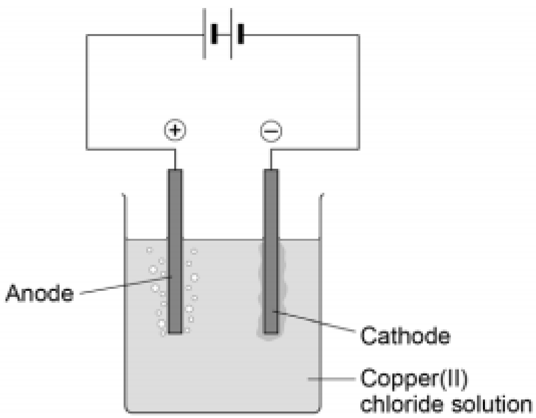
The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved.

At the negative electrode (cathode), hydrogen is produced if the metal is more reactive than hydrogen.

At the positive electrode (anode), the smallest negative ion is turned into a n element.

Required Practical: Electrolysis

1. Pour approximately 50cm^3 copper chloride solution into the beaker.
2. Add the petri dish lid and insert the carbon rods through the holes. The rods must not touch each other.
3. Attach crocodile leads to the rods. Connect the rods to the dc (red and black) terminals of a low voltage power supply.



4. Select 4 V on the power supply and switch on.
5. Look at both electrodes and record your initial observations
6. Get a fresh set of electrodes and replace your copper chloride solution with sodium chloride and again record your observations.

Oberavtions:

Solution	Positive electrode (anode)			Negative electrode (cathode)		
	Observations	Element formed	State	Observations	Element formed	State
Copper (II) chloride	Bubbles of gas Bleaches blue litmus white	Chlorine	gas	Brown/red solid coating on rod	Copper	solid
Sodium chloride	Bubbles of gas Bleaches blue litmus white	Chlorine	gas	Bubbles of gas (more rapid production)	Hydrogen	gas

Reactions at electrodes as half equations (HT)

During electrolysis:

- At the cathode (**negative electrode), positively charged ions gain electrons** and so the reactions are reductions.
- At the anode (**positive electrode**), **negatively charged ions lose electrons** and so the reactions are oxidations.

Reactions at electrodes can be represented by half equations.
This is simply saying what happens to the ion to turn it into an element.
For example:

$$2H^+ + 2e^- \rightarrow H_2$$

and

$$4OH^- \rightarrow O_2 + 2H_2O + 4e^- \text{ or } 4OH^- - 4e^- \rightarrow O_2 + 2H_2O$$

Exothermic and Endothermic Reactions

Energy is conserved in chemical reactions.
The amount of energy in the universe at the end of a chemical reaction is the same as before the reaction takes place.

If a reaction transfers energy to the surroundings the product molecules must have less energy than the reactants, by the amount transferred.

An **exothermic** reaction is **one that transfers energy to the surroundings** so the temperature of the surroundings increases.
Exothermic reactions include combustion, many oxidation reactions and neutralisation.
Everyday uses of exothermic reactions include self-heating cans and hand warmers.

An **endothermic** reaction is **one that takes in energy from the surroundings** so the temperature of the surroundings decreases.
Endothermic reactions include thermal decompositions and the reaction of citric acid and sodium hydrogencarbonate.
Some sports injury packs are based on endothermic reactions.

Required Practical: Temperature Changes

1. Measure 30 cm^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 a table.
4. Measure 5 cm^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 5 cm^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.

Total volume of sodium hydroxide added in cm^3	Maximum temperature in °C		
	First trial	Second trial	Mean
0	20.0	21.0	
5	24.0	24.6	
10	26.8	27.6	
15	28.6	29.6	
20	30.8	31.3	
25	31.8	32.8	
30	32.0	32.6	
35	31.6	31.8	
40	30.6	31.0	

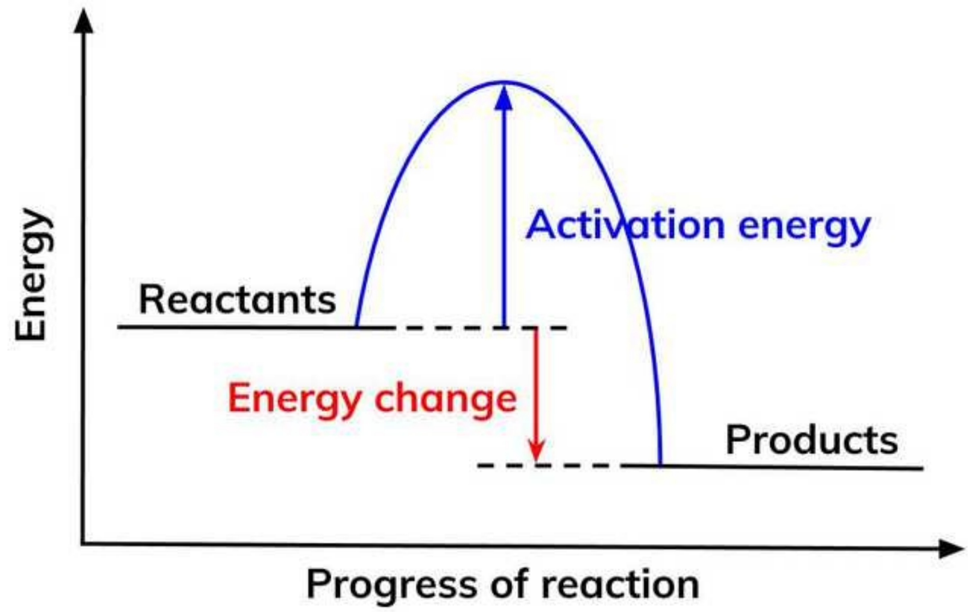
Reaction Profiles

Chemical reactions can **occur only when reacting particles collide with each other and with sufficient energy**.

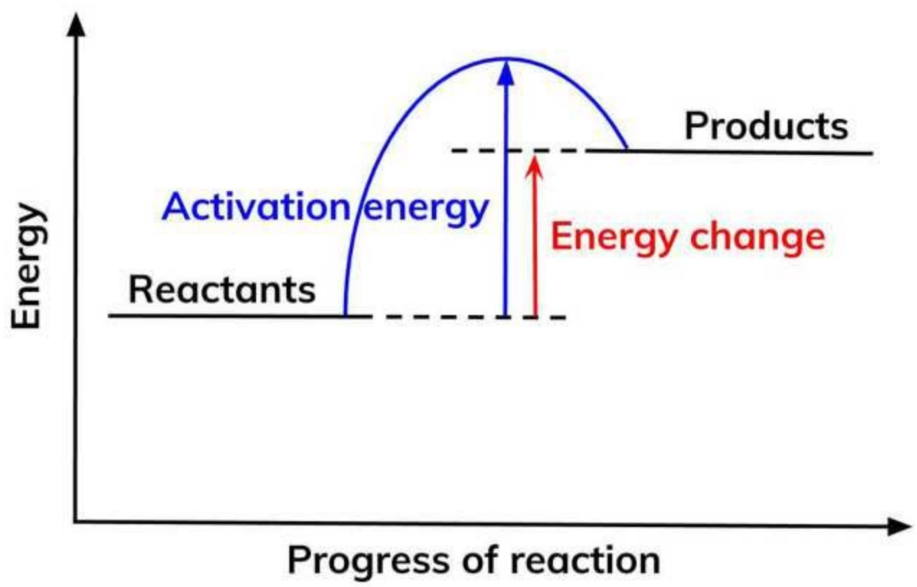
The **minimum amount of energy that particles must have to react is called the activation energy**.

Reaction profiles can be used to show the relative energies of reactants and products, the activation energy and the overall energy change of a reaction.

Exothermic reaction



Endothermic reaction



Energy Changes in a Reaction (HT)

During a chemical reaction:

- energy must be supplied to break bonds in the reactants
- energy is released when bonds in the products are formed.

The energy needed to break bonds and the energy released when bonds are formed can be calculated from bond energies.

The difference between the sum of the energy needed to break bonds in the reactants and the sum of the energy released when bonds in the products are formed is the overall energy change of the reaction.

In an **exothermic reaction**, the **energy released from forming new bonds is greater than the energy needed to break existing bonds**.
In an **endothermic reaction**, the **energy needed to break existing bonds is greater than the energy released from forming new bonds**.

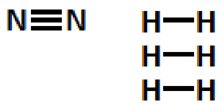
Example:
In a reaction ammonia is made from hydrogen and nitrogen.

$$N_2 + 3H_2 \rightarrow 2NH_3$$

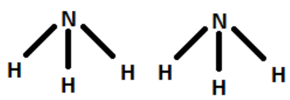
The energy of the bonds are:
$N \equiv N$ has an energy of 945kJ/mol
H - H has an energy of 436 kJ/mol
N - H has an energy of 391 kJ/mol
Calculate the overall energy change for the reaction.

Overall energy change of reaction = energy to break bonds in reactants - energy released when bonds are formed in products

Energy to break bonds in reactants = 945 + 3 x (436) = 2253 kJ



Energy released when bonds are formed in products = 2 x (3 x 391) = 2346 kJ



Overall energy change in reaction = 2253 - 2346 = - 93kJ
93kJ of energy is transferred to the surroundings in this reaction, therefore this is an exothermic reaction.

Chemical Cells and Fuel Cells

Cells and batteries

Cells contain **chemicals which react to produce electricity**.

The **voltage** produced by a cell is **dependent** upon a number of factors including the **type of electrode and electrolyte**.

A simple cell can be made by connecting two different metals in contact with an electrolyte (a liquid which conducts electricity).

Batteries consist of two or more cells connected together in series to provide a greater voltage.

In **non-rechargeable** cells and batteries the **chemical reactions stop when one of the reactants has been used up**. Alkaline batteries are non-rechargeable.

Rechargeable cells and batteries can be recharged because the **chemical reactions are reversed** when an external electrical current is supplied.

Fuel cells

Scientists are trying to make more environmentally friendly sources of energy.
Fuel cells are supplied by an external source of fuel (e.g. hydrogen) and oxygen or air.

The reaction for hydrogen is:
Hydrogen + Oxygen → Water
$2\,H_2 + O_2 \rightarrow 2\,H_2O$

Half equations for the reactions at the electrodes in the fuel cell are:

$$2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$$
$$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$$

The key point is that no pollutants are made, just water. An issue with using hydrogen is that to get hydrogen to use as fuel we **have to use electrolysis** which requires **energy which may come from a non-renewable power plant**.
The **hydrogen is oxidised** within the fuel cell to **produce a potential difference**.
The overall reaction in a hydrogen fuel cell involves the oxidation of hydrogen to produce water.

Hydrogen fuel cells offer a potential alternative to rechargeable cells and batteries.

Advantages:
- Don't need to be recharged
- No pollutants made
- Can be a range of sizes for different uses

Disadvantages:
- Hydrogen in very flammable
- Hydrogen has to be produced using energy
- Hydrogen is difficult to store.