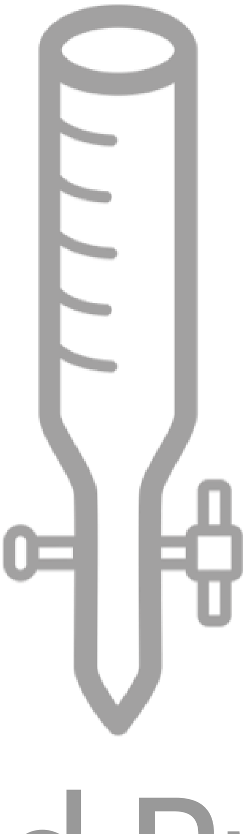
Chemistry

Required Practicals

Paper One



Required Practicals

Making Salts

Key terms

Neutralisation - A chemical reaction of an acid with a base in which a salt and water are formed. If the base is a carbonate or hydrogen carbonate, carbon dioxide is also produced.
Filtration - A technique used to separate an insoluble substance in a solvent.
Evaporation - A technique used to separate a soluble substance in a solvent. It involves the solvent's physical change from a liquid phase to a gas phase.
Crystallisation - A technique which follows evaporation, in which a solid in a crystalline structure forms.

Practice exam question

Outline a safe plan the student could use to make pure, dry, crystals of the soluble salt copper sulfate from the insoluble metal oxide and dilute acid.

Link to video of practical

Method and notes

1. Heat the acid gently until it is almost boiling.
2. Use the spatula to add small amounts of the base.
3. Stop adding the base when some of it remains after stirring.

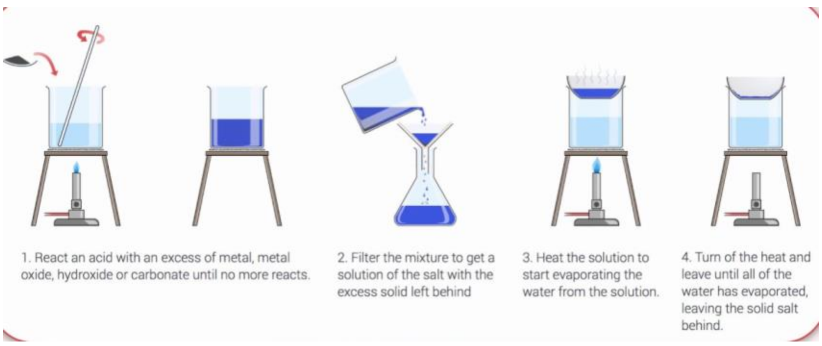
The acid and the base react forming a salt and water. The base is added in excess to ensure all the acid reacts.

4. Filter the contents.
5. Any excess of the base remains as a residue.
6. Pour the contents into an evaporating basin.

The filtrate is an aqueous solution of the salt.

7. Evaporate this gently using a water bath on the tripod and gauze (see diagram). Stop heating once crystals start to form.
8. Transfer the remaining solution to a crystallising dish. Leave this in a cool place for at least 24 hours.

Visual aids:



Temperature Change

Key terms

Exothermic - A reaction that transfers energy to the surroundings.

Endothermic - A reaction that takes in energy from the surroundings.

Practice exam question

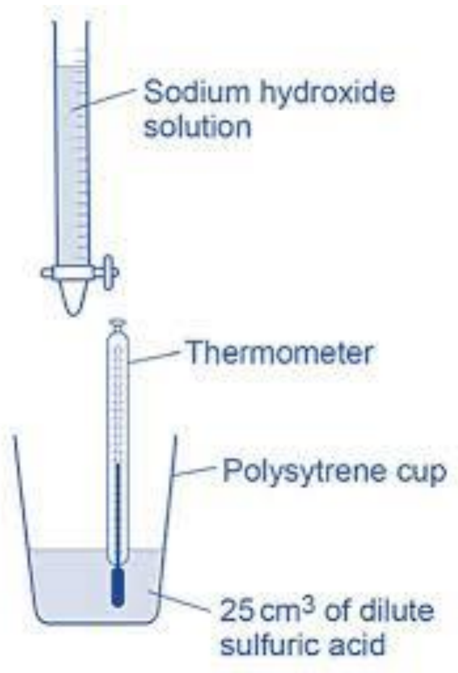
When ammonium chloride is dissolved in water, there is a temperature change.
The water used was at room temperature.
A student added the ammonium chloride to water and stirred the mixture.
a)The water was in a glass beaker. What could the student have used instead of a glass beaker to improve the accuracy? Give a reason why this would improve the accuracy of his results.
b)State two control variables the student should keep the same. Give a reason why changing each of these two control variables would affect the temperature change.

Link to video of practical

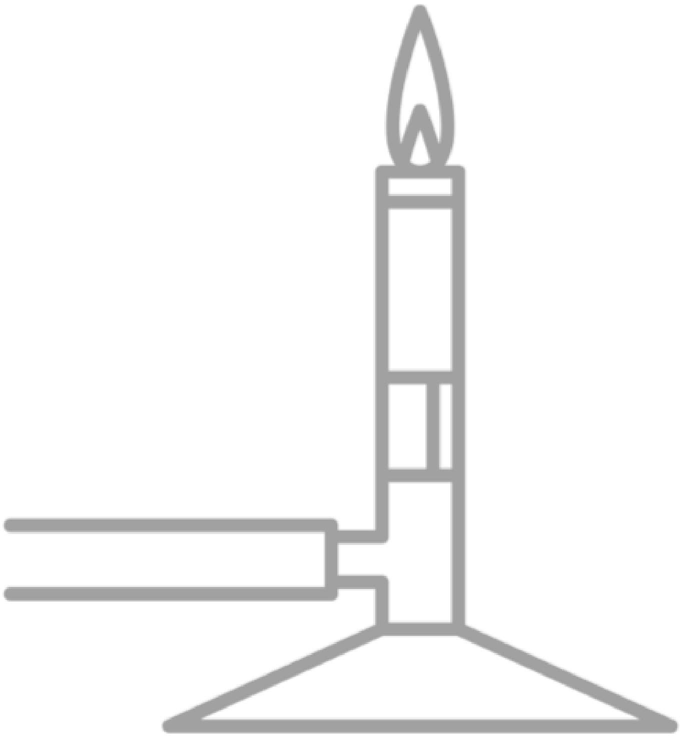
Method and notes

1. Pour the acid into a polystyrene cup.
2. Stand the cup inside the beaker. This will make it more stable.
A polystyrene cup is used because it is an insulator.
3. Use the thermometer to measure the temperature of the acid and record it.
4. Add a certain volume of base into the cup.
5. Fit the lid and gently stir the solution with the thermometer through the hole.
6. When the reading on the thermometer stops changing, record the temperature.
7. Keep on adding the same volume of base until you reach the number of additions the task asks.

Visual aids:



Paper Two



Required Practicals

Rates of Reaction

Key terms

Rate of reaction - how fast a reaction occurs.

Turbidity - The cloudiness of a fluid due to the presence of suspended particles invisible to the naked eye.

Practice exam question

A student investigated the effect of temperature on the rate of a reaction. The student follows the methods described to the left.

a)Explain why the solution goes cloudy.

b)Give two variables the student must control to make the investigation a fair test.

Link to video of practical

Method and notes

1. Add sodium thiosulfate solution into a conical flask and then add water to dilute the solution.
2. Put the conical flask on the black cross.
3. Add hydrochloric acid into the flask.
4. Swirl the flask gently and start the stop clock.
5. Stop the clock when you can no longer see the cross.

The solution becomes cloudy due to the presence of solid sulfur.

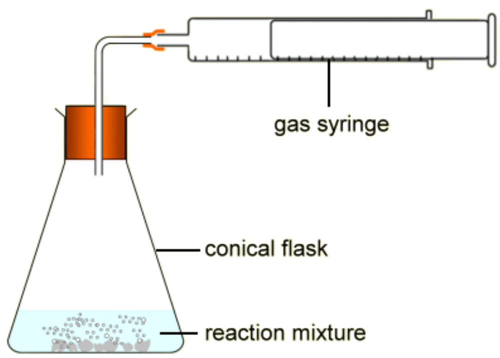
Alternative method:

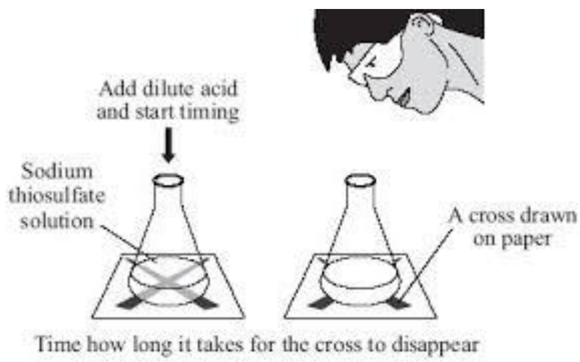
1. Pour the acid into a conical flask.
2. Prepare a water bath using a trough.
3. Fill the other measuring cylinder with water and invert it.

An inverted measuring cylinder or a syringe can be used.

4. Add a strip of magnesium ribbon to the flask, put the bung back into the flask as quickly as you can, and start the stop clock.
5. Record the volume of hydrogen gas given off at suitable intervals and continue until no more gas appears to be given off.

Visual aids:





Chromatography

Key terms

Chromatography - The process whereby small amounts of dissolved substances are separated by running a solvent along a material such as absorbent paper.

Rf (retention factor) - A measurement from chromatography. It is the distance a spot of a substance has been carried above the baseline divided by the distance of the solvent front.

Practice exam question

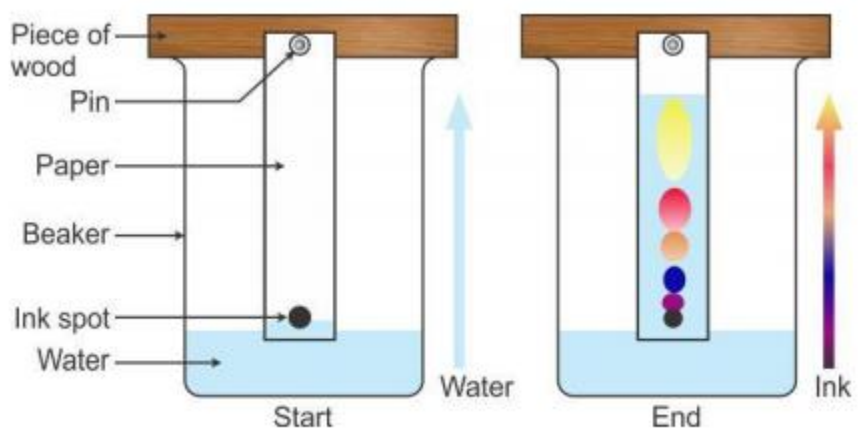
Plan a chromatography experiment to investigate the colours in an ink.

Link to video of practical

Method and notes

1. Use a ruler to draw a horizontal pencil line 2 cm from a short edge of the chromatography paper.
2. Use a glass capillary tube to put a small spot of each of the known colourings (A-D) and then a small spot of the unknown mixture (U).
3. Pour water into the beaker.
Ensure that the pencil line is above the water surface.
4. Tape the edge of the chromatography paper to the glass rod and rest the rod on the top edge of the beaker.
5. Wait for the water solvent to travel at least three quarters of the way up the paper.
6. Remove the paper. Draw another pencil line on the dry part of the paper as close to the wet edge as possible.
This is the distance travelled by the water solvent.

Visual aids:



Paper Chromatography

Water Purification

Key terms

Distillation - Separation of a liquid from a mixture by evaporation followed by condensation.

Potable water - Water that is fit to drink.

Practice exam question

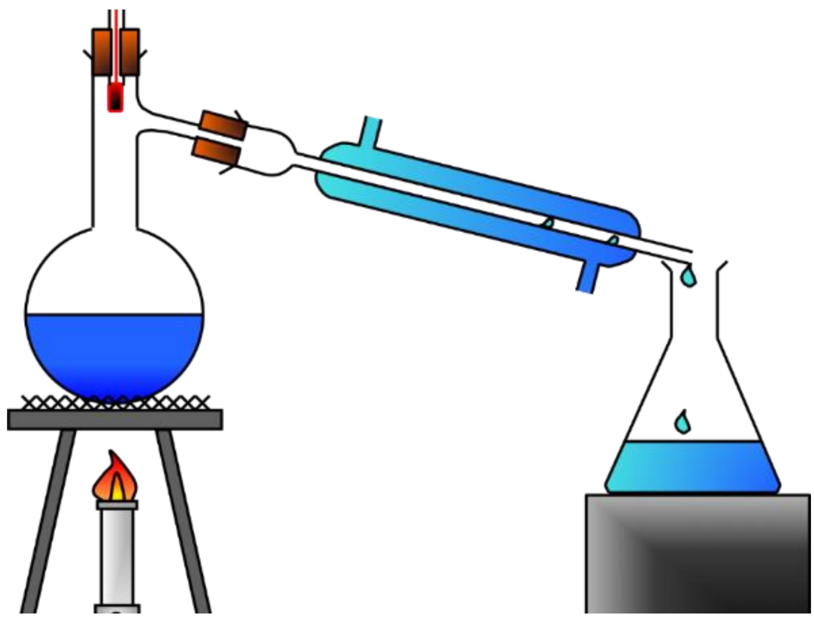
Water in Britain is taken from reservoirs to use as drinking water.
What are the two main steps used to treat water from reservoirs?
Give one reason for each step.

Link to video of practical

Method and notes

Step 1
1. Pour a 1 cm depth of the sea water into a test tube. Add a few drops of UI. Using a pH colour chart, match the colour and record the pH.
Step 2
2. Pour drops of the sea water into a watch glass and place it above a beaker acting as a water bath.
3. Allow all the water to evaporate from the watch glass.
Do not let the water bath boil dry.
Step 3
4. Add seawater in a conical flask and set up the apparatus for distillation.
5. Place a mixture of ice and water in the beaker surrounding the test tube.
6. Heat the sea water with the Bunsen burner until it starts to boil. Then reduce the heat so that the water boils gently.
7. Distilled water will collect in the cooled test tube.

Visual aids:



Key Terms

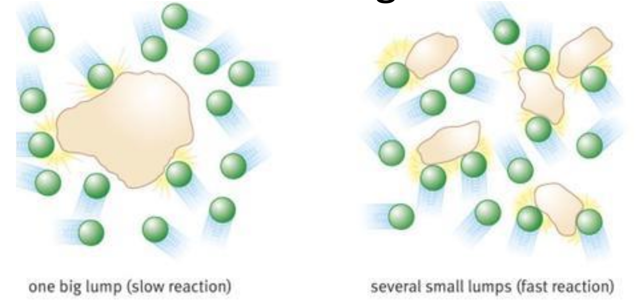
Knowledge Organiser – The Rate and Extent of Chemical Change

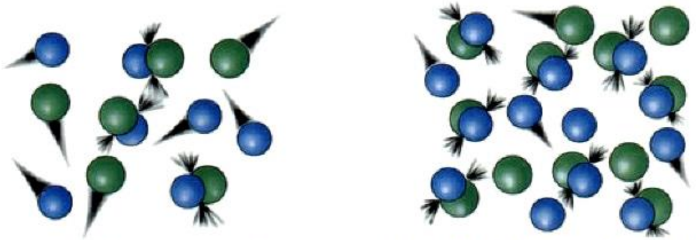
Diagrams

Rate of reaction	The speed at which a reaction takes place. This can be worked out in two ways: Mean rate of reaction = quantity of reactant used ÷ time Mean rate of reaction = quantity of product formed ÷ time
Activation energy	The minimum energy particles must have to react
Catalyst	A substance that speeds up a chemical reaction by lowering the activation energy
Enzymes	Molecules that act as catalysts in biological systems
Closed system	A system where no substances can get in or out
Dynamic equilibrium	System where both the forward and reverse reactions are taking place simultaneously and at the same rate

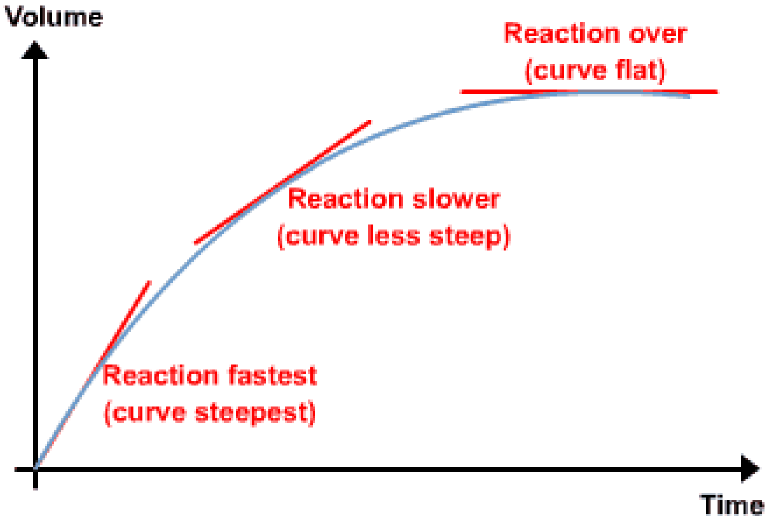
Factors affecting rates of reaction

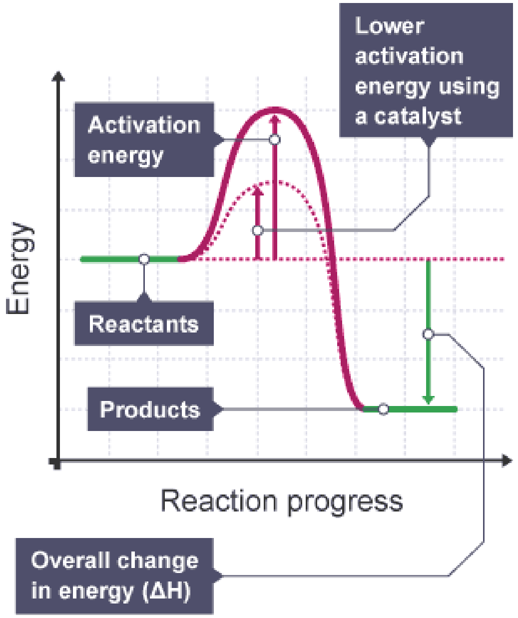
Low surface area High surface area

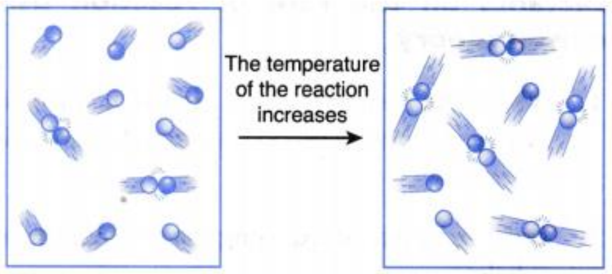




Low concentration = Few collisions High concentration = More collisions



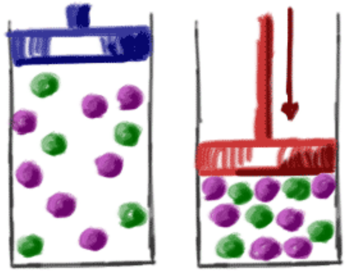




- At a lower temperature, the particles move slower.
- Frequency of collision is lower.

- At a higher temperature, the particles move faster.
- Frequency of collision is higher.

Figure *Temperature of a reaction controls the frequency of collision*



AS PRESSURE INCREASES, THE GAS MOLECULES CAN HAVE MORE COLLISIONS.

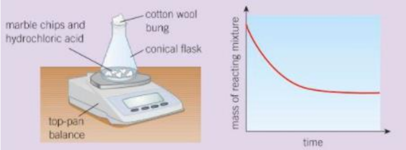
C8 Rates and Equilibrium

Measuring Rate

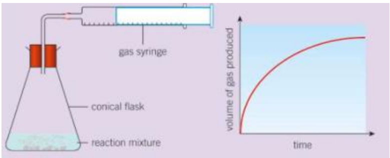
To measure the rate of a reaction you can:

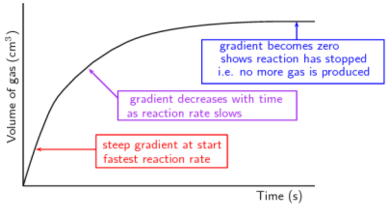
- **Measure how fast the reactants are used up**
- **Measure how fast the products are made**

e.g. Measure mass lost due to gas formed



e.g. Measure volume of gas made

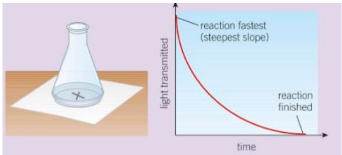




Rate = volume of gas ÷ time

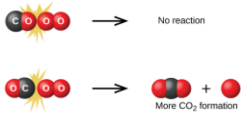
cm^3/s

e.g. Measure time for insoluble product to form



Collision theory

For a reaction to happen reactants must:
collide with **enough energy** (activation energy)



A successful collision is one that leads to a reaction

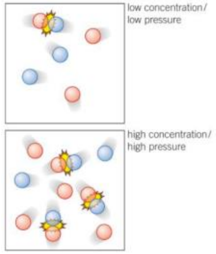
So to increase the rate of a reaction you must either

- **Increase the frequency of collisions**
- **Increase the energy of the collisions**
- **Decrease the energy needed for a collision to be successful**

Factors affecting rate

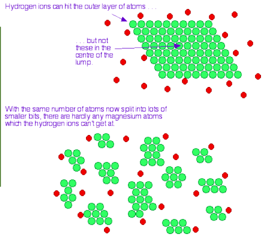
Concentration and Pressure

More particles in the same space.
More frequent collisions



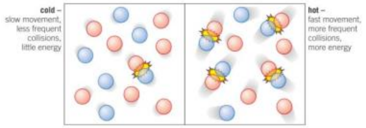
Surface area

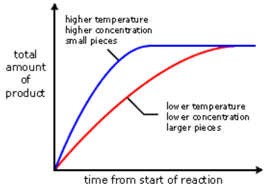
More particles available to react.
More frequent collisions



Temperature

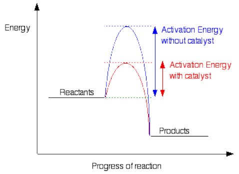
Particles **move faster**.
So they **collide more frequently**.
Particles collide **with more energy**.
So more of the collisions are **successful**.





Catalysts

Lower the energy needed for successful collisions. (Activation energy)
Not used up.
Biological catalysts are called **enzymes**



Reversible reactions

Can go in both directions.

$$A + B \rightleftharpoons C + D$$

If a reaction is exothermic in one direction it is endothermic in the other direction.

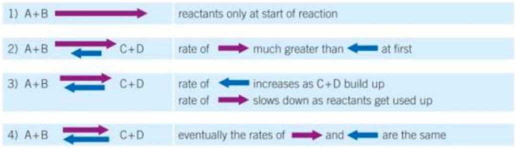
endothermic (in forward reaction)

hydrated copper(II) sulfate (blue) ⇌ anhydrous copper(II) sulfate (white) + water

$$CuSO_4.5H_2O \rightleftharpoons CuSO_4 + 5H_2O$$

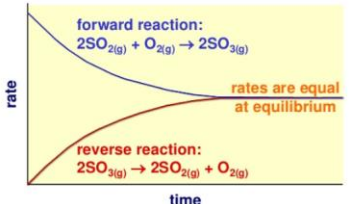
exothermic (in reverse reaction)

In a **closed system** (where nothing can get in or out) an **equilibrium** is reached where the **rate of reaction is the same in both directions.**



At equilibrium:

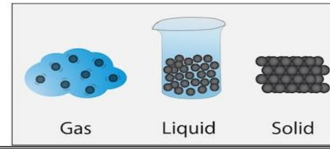
- Rate of forward reaction = rate of reverse reaction.
- Mount of products and reactants don't change.



C2.2 Structure and Bonding

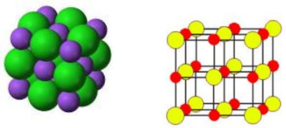
2.1 States of Matter

Solid – (s)
Liquid – (l)
Gas – (g)
Aqueous (aq)



2.3 Properties of Ionic Substances

Ionic compounds have high melting and boiling points **because…**

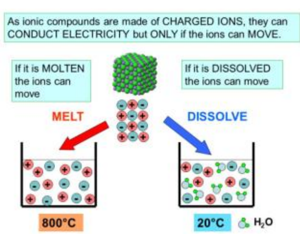


…they are giant structures of atoms (giant ionic lattice) with **strong electrostatic forces** of attraction in **ALL DIRECTIONS** between oppositely charged ions.

A large amount of **energy** is needed to break the many strong bonds.

Only conduct electricity when melted or dissolved in water **because…**

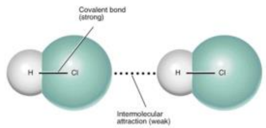
…the **ions are free** to move and so charge can flow.



Properties of Covalent substances

2.4 Small molecules

Small molecules have relatively low melting and boiling points **because…**



…**intermolecular forces** are overcome on melting and boiling and these are weak forces.

The bigger the size of the molecule the higher the melting and boiling point **because…**

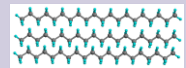
…intermolecular forces increase with the size of the molecules.

Don't conduct electricity **because…**

…the molecules have **no overall electric charge**.

2.6 Giant Structures

Polymers are solids at room temperature **because…**

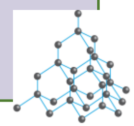


…intermolecular forces increase with the size of the molecules and polymer molecules are **very large**.

Giant covalent compounds have high melting and boiling points **because…**

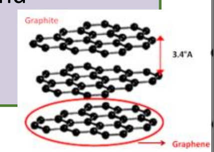
…all of the atoms linked by **strong covalent bonds**.

Diamond is very hard, has a very high melting and boiling point and doesn't conduct electricity **because…**



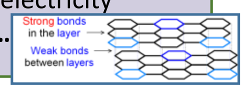
…each carbon is bonded to **4** other carbons by **strong covalent bonds**. There are **no free electrons**.

Graphene is strong, light and an excellent conductor of thermal energy and electricity. **because…**

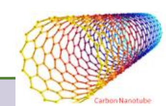


…it is a single layer of graphite so has **free electrons.**

Graphite is very hard, has a very high melting and boiling point and does conduct electricity **because…**



…each carbon is bonded to **3** other carbons by **strong covalent bonds**. It forms **layers** of **hexagonal rings** with no covalent bonds between layers. There are **free electrons**.



Fullerenes (e.g. carbon nanotubes) are extremely strong and are excellent conductors of thermal energy and electricity **because…**

… they have **strong covalent bonds** and **free electrons**.

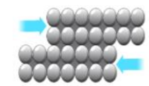
2.7 Properties of Metallic Substances

Metals have high melting and boiling points **because…**

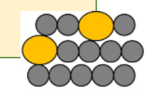
…they are **giant structures** of atoms with **strong metallic bonding**

Can be bent or shaped **because…**

…atoms are arranged in **LAYERS** which can **SLIDE** over each other



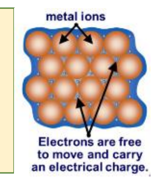
Alloys are harder than pure metals **because…**



Alloys are a mixture of two or more elements, at least one of which is a metal

…the layers are **DISTORTED** so can't slide over each other

Metals are good conductors of electricity and thermal energy **Because…**



…the **electrons are free** to move and carry thermal energy and charge

1.	If the size of a small covalent molecule increases, what effect does this have on the boiling points
2.	Explain why ionic compounds have high melting and boiling points
3.	Explain why ionic compounds can conduct electricity when melted or dissolved in water
4.	What form do ionic compounds take?
5.	How many covalent bonds are their between atoms in diamond and graphite
6.	Why is diamond hard and has very high melting and boiling points?
7.	Name three
8.	Why are pure metals soft?
9.	Why are alloys harder than pure metals?
10.	Why are metals good conductors?
11.	Why do small molecules have low melting and boiling points?

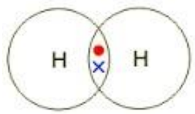
Fullerene	a form of the element carbon that can exist as cage-like or tubular structures based on hexagonal rings of carbon atoms
Delocalised Electron	bonding electron that is no longer associated with any one particular atom
Giant Covalent Structure	a huge 3D network of covalently bonded atoms
intermolecular forces	the attraction between the individual molecules in a covalently bonded substance
Diamond	a form of the element carbon where each carbon atom is covalently bonded to 4 others
Graphite	a form of the element carbon where each carbon atom is covalently bonded to 3 others
Graphene	a form of the element carbon where each atom is covalently bonded to 3 others AND is only one atom thick
Ionic lattice	a giant 3D structure of alternating positive and negative ions, held together by strong electrostatic attraction

Chemistry Knowledge Organiser
C3 - Structure and bonding

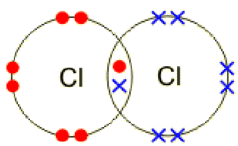
Key Terms	Definitions
Covalent Bonding	Bonding between 2 (or more) atoms where electrons are shared
Molecule	A substance which contains two or more covalently bonded atoms
Lone Pair	A pair of electrons that are not part of the covalent bond

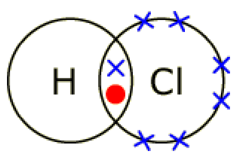
Covalent Bonding

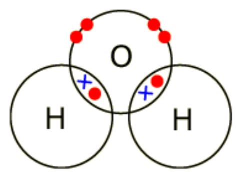
Covalent bonding occurs between non metals. E**lectrons are shared between the atoms**, so that they have a full outer shell. Covalent bonds are strong and require a lot of energy to break. The simplest example is hydrogen: both hydrogen atoms have **one electron in their outer shell. Therefore both hydrogen atoms share one electron each,** to give them both a full outer shell, we can show this bond on a dot and cross diagram.

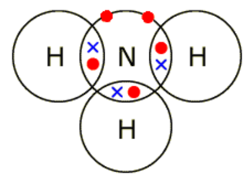


When drawing covalent molecules we use “dot cross diagrams” as we do with ionic compounds. It is important to represent the electrons on one atom with a dot and on the other atom with an X.
The first five examples, **hydrogen, chlorine, water, hydrogen chloride and ammonia (NH_3)** all share one electron per atom in a to make a full outer shell of electrons on each atom.

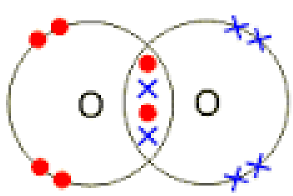


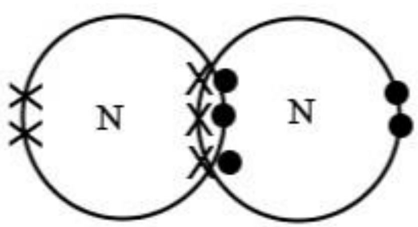






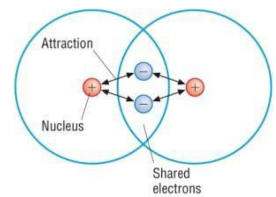
Some atoms need more than one electron to give them a full outer shell, for example oxygen needs 2 electrons to complete its outer shell. Oxygen therefore shares two electrons per atom to **make a double bond.** Nitrogen needs three electrons to complete its outer shell, this forms a triple bond between the two **nitrogen atoms, to make a nitrogen molecule.**





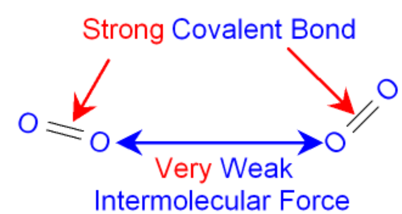
The Nature of a Covalent Bond

Covalent bonds are strong because there is an attraction between the electrons in the covalent bond and the positively charged nucleus. This means a lot of energy is required to break a covalent bond.



Properties of Simple Covalent Compounds

Simple covalent compounds have low melting points and are often gases at room temperature, for **example oxygen and carbon dioxide.** Although the covalent bonds between the atoms are strong, the **intermolecular forces between the molecules are weak. It is very important to remember that covalent bonds are strong but the intermolecular forces are weak .** This means that only a small amount of energy is required to overcome these weak forces.

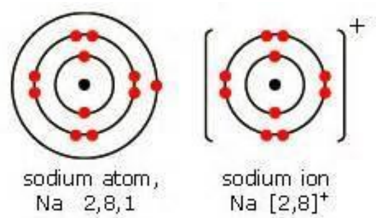


Please see the next page for more properties of covalent compounds.

Chemistry Knowledge Organiser
C3 - Structure and bonding

Ions

All atoms are more stable with a full outer shell of electrons. Some atoms will lose electrons to get a full outer shell: these are metals. Some atoms will gain electrons to get a full outer shell: these **are non metals**. An ion is an atom with a positive or negative charge, these are formed by an atom gaining or losing electrons. For example, sodium has one electron in it's outer shell, it therefore loses one electron to form a Na^{+1} ion. We represent ions with square brackets around the ion and the charge in the top right corner.

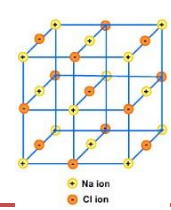


The **group number** indicates how many electrons an atom would have to lose or gain to get a full outer shell of electrons. See below to see what ions different groups form

Group	What happens to the electrons?	Charge on ions
1	Lose 1	+1
2	Lose 2	+2
3	Lose 3	+3
5	Gain 3	-3
6	Gain 2	-2
7	Gain 1	-1

Ionic Lattice

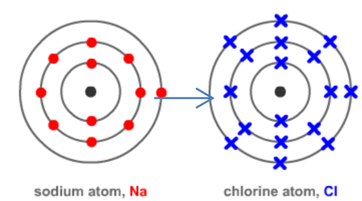
Ionic compounds have **regular structures (giant ionic lattices)** in which there are strong **electrostatic forces** of attraction in all directions between oppositely charged ions.

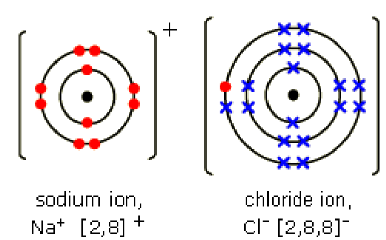


Key Terms	Definitions
Metal	An element which loses electrons to form positive ions
Non Metal	An element which gains electrons to form negative ions
Ion	An atom (or particle) with a positive or negative charge, due to loss or gain of electrons
Ionic Bond	A bond formed by the electrostatic attraction of oppositely charged ion
Electrostatic	The force between a positive and negative charge.

Ionic Bonding

When a metal atom reacts with a non-metal atom electrons in the outer shell of the **metal atom are transferred to the non metal atom**. This means the metal has a positive charge and the non metal has a negative charge. This means there is an **electrostatic attraction** between the two ions, this is what forms an ionic bond. Both atoms will have **a full outer shell** (this is the same as the structure of a noble gas) see example below of sodium chloride.





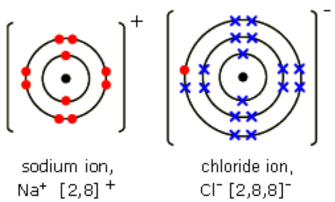
Chemistry Knowledge Organiser
C3 - Structure and bonding

Key Terms	Definitions
Ionic Lattice	The regular 3D arrangement of ions in an ionic compound
Giant	When the arrangement of atoms is repeated may times, with large numbers of atoms or ions
Aqueous	When a substance is dissolved in water
Empirical Formula	The simplest ratio of atoms in a compound

Ionic Bonding- Models

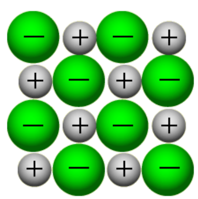
There are a number of ways we can represent ionic bonding all; of these have **advantages and limitations.** For example all the diagrams below show ways we can represent **sodium chloride**

1. **Dot and cross diagrams-** These show clearly how the electrons are transferred. It does not, however, show the 3D lattice structure of an ionic compound or that this is a giant compound.



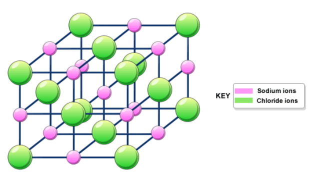
2. **2D ball and stick model of ionic bonding**

This has the advantage of showing that electrostatic forces happen between oppositely charged ions in an ionic compound. However, does not show the 3D structure of an ionic compound.



3. **3D Ball and Stick model of ionic bonding**

This clearly shows the 3D structure of the **ionic lattice** and how different ions interact with other ions **in all directions** to create an ionic lattice.



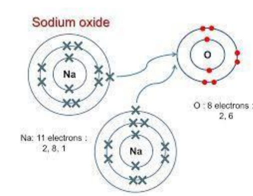
Properties of Ionic compounds

Ionic compounds have **high melting points, due to strong electrostatic forces between the oppositely charged ions.** This means a lot of energy is required to break these bonds. For example the melting point of sodium chloride is 801 °C.

Ionic compounds **do not conduct electricity** as a solid. They **do conduct electricity** if they are dissolved in water (aqueous) or in the liquid state. This is because the ions are free to move, carrying the electric charge.

Empirical Formula of Ionic Compounds

In sodium chloride, 1 sodium atom gives an electron to a chlorine atom, therefore the empirical formula is NaCl. However there are some examples where the ratio of atoms is not 1:1. For example when sodium bonds with oxygen, sodium only wants to lose one electron but oxygen needs to gain two. So you need two sodium atoms for every oxygen so the **empirical formula is Na_2O.**



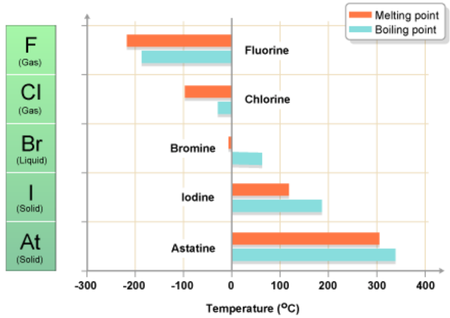
Chemistry Knowledge Organiser
C3 - Structure and bonding

Key Terms	Definitions
Polymer	A very large molecule, made from monomers
Repeating Unit	The shortest repeating section of a polymer
Intermolecular Forces	The force of attraction between two molecules

Properties of Covalent Compounds-Continued

The size of the intermolecular force between molecules increases as the molecules get larger. This is because a force called the van der Waals force increases (you do not need to know that for GCSE). For example as you go down group 7, the boiling points increase because **the molecules get larger**.

As you can see from the graph below, the boiling point of fluorine is -188°C and is therefore a gas at room temperature, whereas the melting point of astatine is 302°C and is therefore a solid at room temperature. This is because the intermolecular forces between the larger astatine molecules are larger than between the **smaller fluorine molecules.**

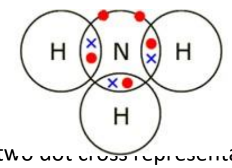


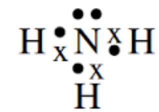
As well as having low melting points, covalent compounds **do not conduct electricity. T**his is because they have no free electrons or ions and therefore there is nothing to carry the electric charge**.** Remember pure water does not conduct electricity, only when it has ions dissolved in it will it conduct.

Representing Covalent Compounds

Like ionic compounds, there are variety of ways that scientists use to represent covalent compounds**.**

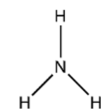
1. **Dot cross diagram**





There are two dot cross representations of ammonia shown above. The advantages of these diagrams are that it is very clear, which electrons are used in bonding and which are lone pairs. However it does not show the 3D structure of the molecule and this can be extremely important for scientists.

2. Ball and stick model

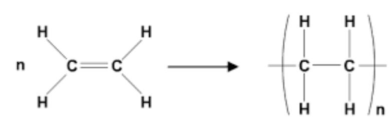


Ammonia (Ball and stick model)

A ball and stick diagram can either be 2D or 3D. While the 2D version clearly shows which atoms are bonded together, the 3D version gives the scientist more information about the 3D shape and the angles between the bonds of the molecule.

Polymers

Polymers are large covalent compounds which can be many thousands of atoms in length. They are made from small molecules known as **monomers.** Rather than drawing out all the atoms in a polymer we draw a **repeating unit** which is the structure of the monomer in square brackets, with a n representing a very large number of atoms. Polymers have higher melting points than smaller covalent compounds like carbon dioxide as the intermolecular bonds are stronger. However the bonds are not as strong as they are in ionic or giant covalent compounds so the melting points are lower than those compounds.



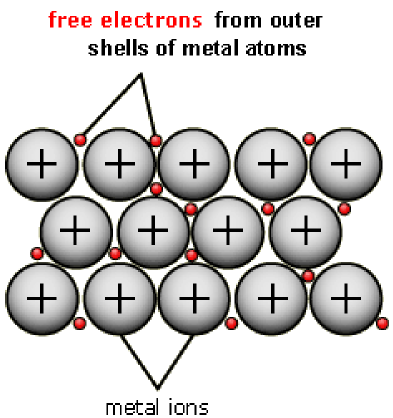
n = a big number of monomers

Chemistry Knowledge Organiser
C3 - Structure and bonding

Key Terms	Definitions
Metallic Bonding	A type of bonding which occurs only in metals
Alloy	A mixture of 2 or elements, one of which is a metal (the other element may be metal or non metal)
Delocalised electron	An electron that is not attached to an atom
Malleable	The ability of a material to be bent into shape.

Metallic Bonding

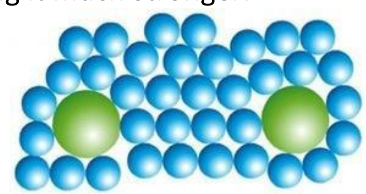
Metals form giant structures. The metal atoms form a regular pattern and the donate their outer electron to the **"sea of delocalised electrons".** These electrons are free to move. The 2D structure of metallic bonding looks like this:



This would be the structure of a group 1 metal like sodium, if it were a group 2 metal like magnesium then the charge on the ions would be Mg^{2+}.

Alloys

Alloys are mixtures of **2 or more elements, one of which is a metal.** Examples of alloys include brass and steel. Metals are alloyed so that the regular structure of metals is changed and the layers of ions can no longer slide over one another; therefore making it much stronger.



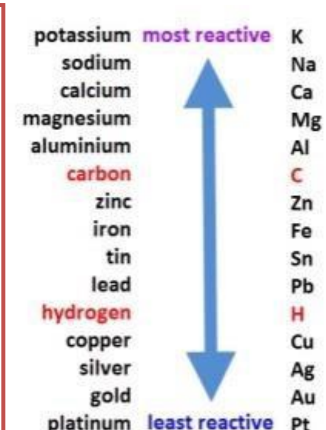
Properties of Metals

Metals are **good conductors of electricity**, due to the delocalised electrons, which can carry the electric charge. Metals are also **good conductors of heat** as the free electrons can transfer the heat energy through the metal.

Metals are also **malleable** (bendy) as the layers of ions can easily slide over one another. This means that many pure metals are too soft for uses such as building.

Reactivity of metals

When a metal reacts it **forms a positive ion.** The easier it is for a metal to form a positive ion, the more reactive it is. This is shown in the reactivity series; you should memorise the position of different elements:



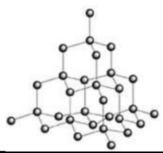
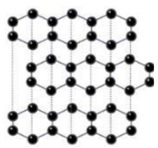
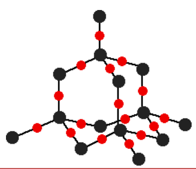
Chemistry Knowledge Organiser
C3 - Structure and bonding

Key Terms	Definitions
Giant Covalent	Giant covalent structures contain a lot of non-metal atoms, each joined to adjacent atoms by covalent bonds
Delocalised electron	An electron that is not attached to an atom
Allotrope	Different forms of the same element for example diamond and graphite are allotropes of carbon
Macromolecule	A molecule which contains many atoms

Giant Covalent Compounds

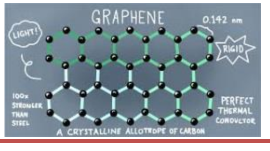
In a giant covalent structure all atoms are bonded to each other by strong covalent bonds. Giant covalent compounds have a **high melting point** because many strong covalent bonds need to be broken and this requires a lot of energy.

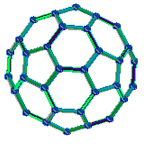
There are three examples you need to know, diamond, graphite and silica (see table below)

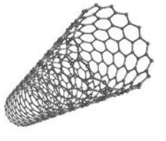
Substance	Diagram	Description	Properties
Diamond		Each carbon is covalently bonded to four other carbons	Very hard, very high melting point, due to strong covalent bonds. Does not conduct electricity – no free electrons/ions.
Graphite		Each carbon is covalently bonded to 3 other carbons, there are weak (non covalent) bonds between the layers.	High melting point, conductor of electricity due to **delocalised electrons which can carry a charge.** Slippery as layers can slide over each other
Silica		Every silicon atom is bonded to 2 oxygen atoms and vice versa	High melting point

Graphene and Fullerenes

There are other forms of carbon which have been discovered recently: **graphene is a single layer of graphite** so it is 1 atom thick. Fullerenes are molecules of carbon with hollow shapes. The most famous example is Buckminsterfullerene (C_{60}). Fullerenes have use in drug delivery and as catalysts. Carbon nanotubes are cylinder shaped fullerenes, these are strong and are excellent conductors of both **heat and electricity.**







Chemistry Knowledge Organiser
C3 - Structure and bonding – triple students only

Key Terms	Definitions
Nanoparticles	A particle between 1nm and 100nm in diameter.
Surface area to volume ratio	The surface area of a substance divided by the volume.
Nanometre	A unit of measurement: $1x10^{-9}$m
Catalyst	A substance which speeds up a reaction, without being used up.

Nanoparticles

Nanoparticles have a diameter **between 1 nm and 100 nm**, this means they are only a few hundred atoms in size. There is a field of Science known as nanoscience which is dedicated to the study of nanoparticles.

Nanoparticles are smaller much **than coarse particles**, which have a diameter between 1×10^{-5} m and 2.5×10^{-6} m. They are also smaller than **fine particles,** which are defined as having a diameter of 100 and 2500 nm.

Nanoparticles have an **extremely large surface area to volume ratio**, this gives them a variety of useful properties.

Uses of nanoparticles

The high surface area to volume ratio means nanoparticles will make excellent catalysts, see more on this in the rate of reaction topic.

Nanoparticles also have many potential applications in medicine for example:

- The targeted delivery of drugs- they are more easily absorbed into the body and therefore could be use to deliver drugs to specific tissues.
- Making synthetic skin

Nanoparticles are also used in the following items:

- Silver nanoparticles have antibacterial properties. These can be used in things like clothing, deodorants and surgical masks.
- Some nanoparticles are electrical conductors, these can be used to make components in very small circuit boards.
- Nanoparticles are also used in cosmetics, to make them less oily
- Nanoparticles are also used in sun creams, they provide better protection from UV than conventional sun creams. They also provide better skin coverage.

Dangers on Nanomaterials

The long term affects of nanomaterials on the body have not been well researched. For example when using suncream, nanoparticles are absorbed through the skin. The affects of long term exposure to these has not been well researched.

Some people believe anything containing nanoparticles should be clearly labelled.

C5 Energy Changes

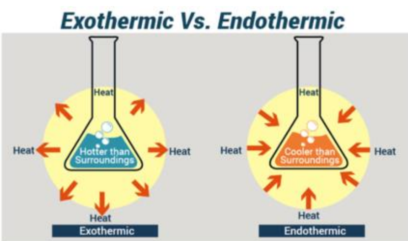
1.1 Exothermic vs Exothermic

Exothermic

In some reactions more energy comes OUT than goes in

The reactants have more energy than the products.

e.g. combustion, oxidation, neutralisation.



Endothermic

In some reactions more energy goes IN than comes out.

The products have more energy than the reactants.

e.g. thermal decomposition

1.1 Uses

Exothermic

Self heating cans, hand warmers

Chemicals react in an exothermic reaction and give OUT heat energy.

Endothermic

Cool packs for sports injuries

Chemicals react in an Endothermic reaction and take IN heat energy – therefore cooling the surroundings.

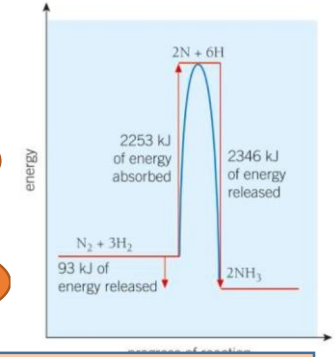
1.3 Bond energy Calculations (HT)

HIGHER ONLY

BINMIX
Bond Breaking Is eNdothermic
Bond Making Is eXothermic

Exothermic – More energy comes OUT making bonds

Endothermic – More energy goes IN breaking bonds



Calculate the energy change in a reaction using average bond energies. **Bond energy** is the amount of energy needed to break one **mole** of a particular bond.

Add together the bond energies for all the bonds in the **reactants** – this is the 'energy in.

Add together the bond energies for all the bonds in the **products** – this is the 'energy out'.

Calculate the energy change = energy in – energy out

Bond	Bond Energy / kJ/mol
H-H	436
Cl-Cl	243
H-Cl	432

Worked example – an exothermic reaction

Hydrogen and chlorine react to form hydrogen chloride gas:

H−H + Cl−Cl → 2 × (H−Cl)

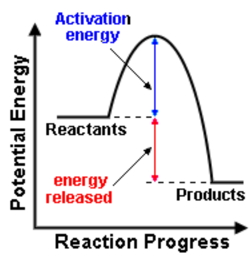
Energy in = 436 + 243 = 679 kJ/mol
Energy out = 2 × 432 = 864 kJ/mol

Energy change = in – out = 679 – 864 = **–185 kJ/mol**

The energy change is negative, showing that energy is released to the surroundings in an **exothermic** reaction.

1.2 Reaction Profiles

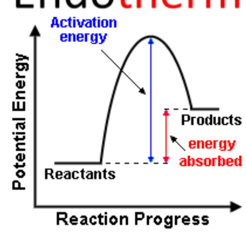
Exothermic



Exothermic reaction

Products at LOWER energy than reactants

Endothermic



Endothermic reaction

Products at HIGHER energy than reactants

Activation Energy is the energy needed to start a reaction.

1.	What is an exothermic reaction?
2.	State examples and uses of exothermic reactions
3.	What is an endothermic reaction?
4.	State examples and uses of endothermic reactions
5.	Define activation energy
6.	Draw and label the reaction profile for an exothermic reaction
7.	Draw and label the reaction profile for an endothermic reaction
8.	*HT ONLY* – Describe an exothermic reaction in terms of bond energies
9.	*HT ONLY* – Describe an endothermic reaction in terms of bond energies

Exothermic reaction	Reaction where thermal energy is transferred from the chemicals to the surroundings and so the temperature increases
Endothermic reaction	Reaction where thermal energy is transferred from the surroundings to the chemicals and so the temperature decreases
Activation energy	The minimum energy particles must have to react – measured on an energy diagram from the level of the reactants to the peak of the line
Reaction Profile	A diagram that shows how the energy changes during a chemical reaction from reactants to products
Energy Change	Always measured on the energy diagram from the level of the reactants to the level of the products

C10 Using Resources

1.1 Finite and Renewable

What do the words mean??

Finite = Will run out eventually
Renewable= We can replace them as we use them
Sustainable = meets the needs of the current generation without compromising he ability of future generations to meet their needs.

What do we use the earth's resources for?
- **Warmth**
- **Shelter**
- **Food**
- **Transport**

We can use them as natural resources or process them.

'Natural resources' + agriculture provides
- **Food**
- **Timber**
- **Clothes**

Finite resources are processed to get us
- **Energy**
- **materials**

e.g. Cotton is natural and we grow cotton plants. OR we can use synthetic materials e.g. nylon

e.g. Coal, oil and gas are used for energy.
e.g. metal ores are mined to get metals.

1.2 & 1.3 Treating water

Potable water must have low levels of SALTS and MICROBES (it isn't PURE water)

Rainwater in lakes, rivers and reservoirs

Filter

Why? To remove insoluble solids

Sterilise – chlorine, ozone, or UV light

Why? To kill microbes

Sewage

Screening and grit removal

Sedimentation

Anaerobic digestion of sludge

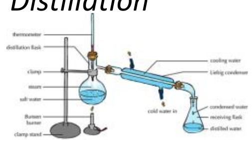
Aerobic treatment of effluent

Industrial and agricultural waste water – remove organic matter and harmful chemicals

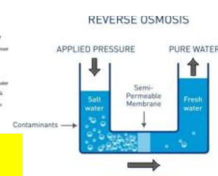
Salt water

Desalination

Distillation



Reverse osmosis



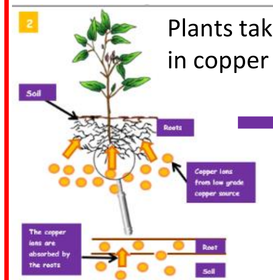
☹ Both use a lot of energy

HIGHER ONLY

1.4 Alternative Metal Extraction

Why bother?
Running out of metal ores

Phytomining



Plants take in copper

- BURN plants
- React ASH with sulphuric acid

Bioleaching

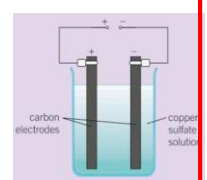
Bacteria feed on metal ore

'leachate solution' contains copper compounds

How to get the copper from the compound

Displacement using scrap iron

Electrolysis



2.1 & 2.2 LCA and RRR

Life Cycle Assessments

Water

Resources

Energy sources

Waste

Reducing use of resources

Why bother?
Reduce...use of limited resources

Why bother?
Reduce...use of energy resources

Why bother?
Reduce...waste and environmental impacts

mining

landfill

1	What is a finite resource? (also called non-renewable resources)
2.	What is a renewable resource?
3	What is potable water?
4	State the sterilising agents use to make water potable in the UK
5	What is desalination?
6	Describe the process of distillation
7	Describe the process of reverse osmosis
8	State the stages in waste water treatment
9	*HT ONLY* – Describe phytomining
10	*HY ONLY* – Describe bioleaching
11	What is a life cycle assessment (LCA)?

Finite resource	A resource that cannot be replaced once it has been used.
Renewable resource	A resource that we can replace once we have used it.
Sustainable development	Using resources to meet the needs of people today without preventing people in the future from meeting theirs.
Life cycle assessment	An examination of the impact of a product on the environment throughout its life.
Value judgement	An assessment of a situation that may be subjective, based on a persons opinion and / or values.
Desalination	Process to remove dissolved substances from sea water to make it potable
Ore	A rock from which a metal can be extracted for profit.
Phytomining	The use of plants to absorb metal compounds from soil as part of metal extraction.
Bioleaching	The use of dilute acid to produce soluble metal compounds from insoluble metal compounds.
Leachate	A solution produced by leaching or bioleaching.
Aerobic digestion	Process used to treat the effluent from sewage – requires oxygen
Potable water	Water that is safe to drink
Reverse osmosis	Process used to remove salt from sea water – uses lots of energy because a high pressure is needed
Distillation	Process used to create potable water from sea water by heating it to evaporate and then condensing it. Uses lots of heat energy.
Anaerobic digestion	Process used to treat the sludge from sewage in the absence of oxygen
Disinfection	Process used to kill microbes to make water safe to drink (UV light or chlorine are often used)
Screening	Removal of large solids
Sedimentation	Removal of small solids by letting them settle

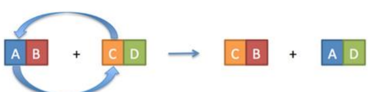
C4 Chemical Changes part 1

1.1 Displacement reactions and metal extraction

potassium	most reactive	K
sodium		Na
calcium		Ca
magnesium		Mg
aluminium		Al
carbon		C
zinc		Zn
iron		Fe
tin		Sn
lead		Pb
hydrogen		H
copper		Cu
silver		Ag
gold		Au
platinum	least reactive	Pt

Reactivity depends on tendency to form metal ion

A more reactive metal will DISPLACE a less reactive metal



HT: OILRIG
Oxidation **I**s **L**oss of electrons
Reduction **I**s **G**ain of electrons

Metal + Oxygen → Metal Oxide

Metal + Water → Metal Hydroxide + hydrogen

Metal + acid → Metal salt + Hydrogen

Unreactive metals can be found in the Earth as themselves (eg Gold)

Carbon Reduction: Extracting metals less reactive than carbon – Oxygen is taken away = **reduction**

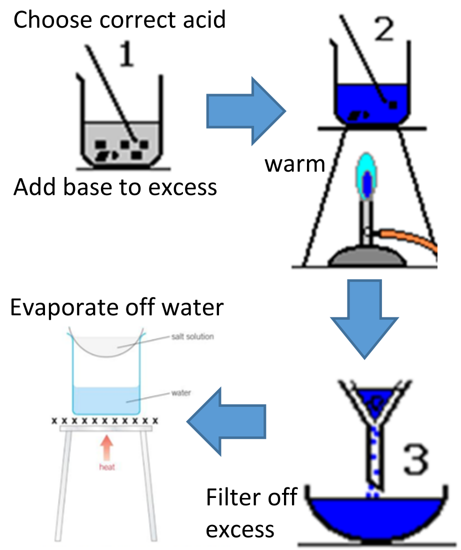
2.1-2.3 Reactions of acids

Acid+metal → salt + hydrogen
Acid + alkali → salt + water
Acid + insoluble base → salt + water
Acid + carbonate → salt + water + carbon dioxide

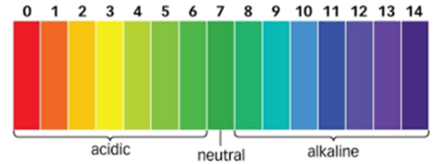
HT: OILRIG
e.g. $2HCl + Mg \rightarrow MgCl_2 + H_2$
Magnesium is oxidised
$Mg \rightarrow Mg^{2+} + 2e^-$

Hydrochloric Acid → Chlorides
HCl
Nitric Acid → Nitrates
HNO_3
Sulphuric Acid → Sulphates
H_2SO_4

RP: Preparation of a dry sample of a soluble salt



2.4 Neutralisation



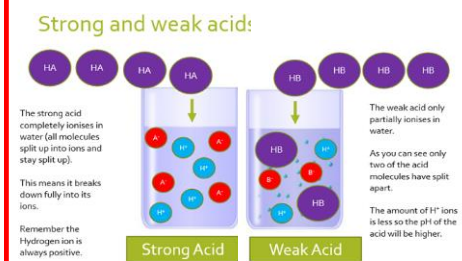
Acids produce H^+ ions
Alkalis produce OH^- ions

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

HIGHER ONLY

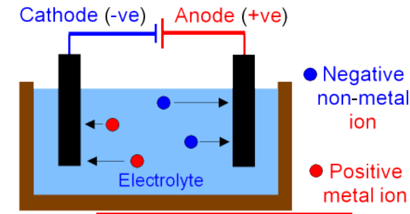
HT: Strong and Weak acids

Concentration of hydrogen ions in mol/dm³	pH
0.10	1.0
0.010	2.0
0.0010	3.0
0.00010	4.0



3.1 Electrolysis

..of molten:



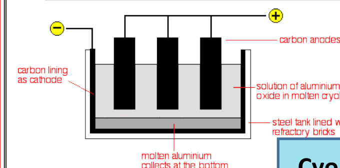
HIGHER ONLY

Higher:
At the cathode
$Pb^{2+} + 2e^- \rightarrow Pb$

Higher:
At the anode
$2Br^- \rightarrow Br_2 + 2e^-$
or
$2Br^- - 2e^- \rightarrow Br_2$

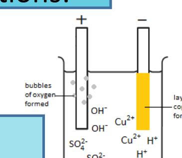
..to extract aluminium:

Oxygen goes to anode → Carbon electrode reacts with it → CO_2 (needs replacing)



Cyolite reduces the melting point

..of solutions:



At the anode:
Halide (Gp7)
Oxygen

At the cathode:
Least reactive

1.	What is produced when metals react with oxygen? Give an example
2.	What happens to a metal when it reacts with different substances?
3.	How are unreactive metals found?
4.	How are metals less reactive than carbon extracted?
5.	How are metals more reactive than carbon extracted?
6.	What are oxidation and reduction reactions?
7.	Describe electrolysis
8.	Which ion moves to which electrode in electrolysis?
9.	Draw an labelled diagram of the set up needed for simple electrolysis
10	What is produced when a metal reacts with an acid?
11	What makes a chemical an acid?
12	State the general equation for the reaction between an acid and an alkali
13	Describe the steps to make a soluble salt

Reactivity series	**An arrangement of metals in order of reactivity**
Displacement reaction	Reaction where a more reactive element takes the place of a less reactive element in a compound
Oxidation	A reaction in which a substance loses electrons (gains oxygen)
Reduction	Reaction in which a substance gains electrons (loses oxygen)
Ore	A rock from which a metal can be extracted for profit
Acid	Solution with a pH less than 7; produces H^+ ions in water
Alkali	Solution with a pH more than 7; produces OH^- ions in water
Aqueous	Dissolved in water
Strong acid	Acid in which all the molecules break into ions in water
Weak acid	Acid in which only a small fraction of the molecules break into ions in water
Dilute	A solution in which there is a small amount of solute dissolved
Concentrated	A solution in which there is a lot of solute dissolved
Neutralisation	A reaction that uses up some or all of the H^+ ions from an acid
Electrolysis	Decomposition of ionic compounds using electricity
Electrolyte	A liquid that conducts electricity
Discharge	Gain or lose electrons to become electrically neutral
Inert electrodes	Electrodes that allow electrolysis to take place but do not react themselves