



AQA GCSE CHEMISTRY

Knowledge Organiser

and

Required Practicals

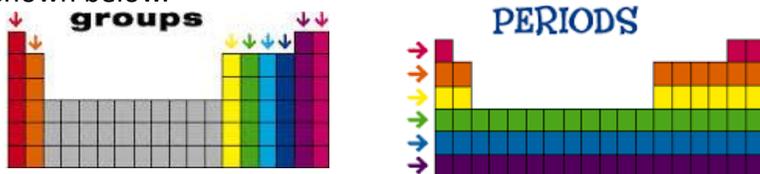


Chemistry Knowledge Organiser

C1 - Atomic structure

Elements

- An **element** contains only one type of atom. All elements are given a symbol and are found on the periodic table. You need to learn the symbols for the first 20.
- The Periodic Table is arranged into groups (columns) and periods (rows), as shown below.



Elements in the same group have:

- The same number of electrons in their outer shell
- Similar properties

Elements in the same period have:

- The same number of electron shells

Compounds

- Compounds are 2 or more elements that are chemically bonded
- These are made in chemical reactions.
- Compounds are given a formula for example carbon dioxide is CO_2 means 1 carbon atom and 2 oxygen atoms.
- Another example is calcium hydroxide $\text{Ca}(\text{OH})_2$ which means 1 calcium, 2 oxygen atoms and 2 hydrogen atoms

Chemical Reactions

- In some chemical reactions it may appear that there are less products than there were reactants; however this is often because a gas has been made and this has escaped into the atmosphere.



Key Terms	Definitions
Element	A substance that contains only one type of atoms
Mixture	A mixture is two or more different atoms which are not chemically bonded
Compound	Two or more elements that are chemically bonded
Group	The columns on the Periodic Table
Period	The rows on the Periodic Table
Reactant	What you start with in a chemical reaction
Product	What is made in a chemical reaction

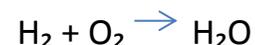
The Conservation of Mass

- In a chemical reaction, chemical bonds are broken the atoms are rearranged and the chemical bonds are made again.
- In a chemical reaction, **mass is never lost**, you must start and finish with the same mass.

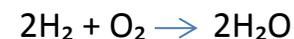


Balancing Equations

- We need to write balanced chemical equations represent chemical reactions and the conservation of mass.
- For example: The equation below shows hydrogen and oxygen making water but there are more oxygen atoms on the right than the left.



- In the equation below there are 4 hydrogen atoms on the left and right of the equation and 2 oxygen atoms on each side



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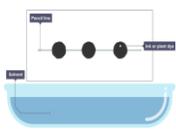
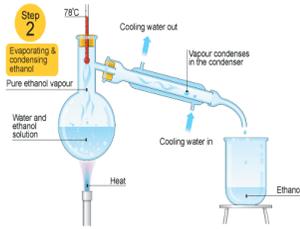
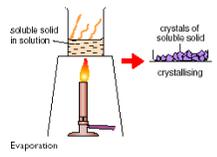
C1 - Atomic structure

Pure and Impure Substances

- A pure substance contains only one type of **element** or **compound**.
- **An impure substance** contains more than one type of element or compound in a mixture, for example salt water contains NaCl and H₂O. All mixtures are impure substances.
- Mixtures are much easier to separate than elements or compounds as they are not chemically bonded
- There are a variety of ways that mixtures can be separated and they are outlined below. Remember that these are all physical changes and chemical bonds are not broken during any of these processes.

Key Terms	Definitions
Pure	A substance made of only ONE type of element or compound
Impure	A mixture of elements and/or compounds
Chromatography	A technique where mixtures can be separated based on their solubility.
Distillation	A separation technique which means a mixture of two liquids is heated
Crystallisation	Method of mixture separation where a solvent is evaporated, leaving the solute behind.

Separating Impure Substance

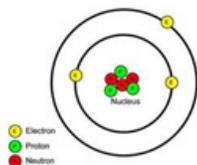
Name	Diagram	Explanation
Chromatography		<ul style="list-style-type: none"> • Different substances travel different distances up the paper depending on their solubility in the solvent used (it is often water but not always). The more soluble, the further it moves up the paper • Line must be drawn with pencil because pencil will not run. • Artificial colours in foods can be identified using chromatography. Additives do not necessarily have a colour and therefore are identified using chemical analysis.
Distillation		<ul style="list-style-type: none"> • Distillation is when two liquids with <i>different boiling points</i> are separated • For example ethanol (alcohol) boils at 78 °C and water boils at 100 °C • If you heat a mixture of water and ethanol to 80°C the ethanol will evaporate but the water will not. • You then condense the ethanol and collect the pure ethanol
Crystallisation		<ul style="list-style-type: none"> • Crystallisation is when a solvent is evaporated from a solute.

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C1 - Atomic structure

The structure of the Atom

- All matter is made from atoms. Atoms are very small. The radius of atom is about 1×10^{-10} m (this is also known as 0.1 nanometres).
- The central part of the atom is known as the nucleus. It is only 1×10^{-14} m across, which is 10,000 times smaller than the total atom.
- An atom is made up of three subatomic particles: **protons**, **electrons** and **neutrons**.
- Protons and neutrons are found in the nucleus
- Electrons are found orbiting the nucleus in shells (also known as *energy levels*).

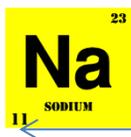


- The mass and charges of the sub atomic particles is shown below:

	Mass	Charge
Proton	1	+1
Neutron	1	0
Electron	0	-1

- Atoms have **no overall charge** because they have the same number of positive protons as negative electrons.

Atomic Number and Mass Number



Mass number: This is the total of protons+neutrons

Atomic number: This is the number of protons

Therefore sodium has 11 protons, 11 electrons and $23-11=12$ neutrons

Key Terms	Definitions
Atom	The particles that make up all substances with mass, they contain protons, neutrons and electrons.
Nucleus	The centre of an atom, it contains protons and neutrons.
Nanometre	A unit of measurement: 1×10^{-9} m
Proton	A sub atomic particle found in the nucleus, it has a charge of +1 and a relative mass of 1.
Electron	A sub atomic particle found in the shells of an atom, it has a charge of -1 and a negligible mass
Subatomic	These are the smaller particles that make up an atom
Neutron	A sub atomic particle found in the nucleus of an atom, it has a charge of 0 and a mass of 1
Atomic Number	The number of protons in an atom.
Mass Number	The total of protons and neutrons in an atom.

Electron Configuration

There are very strict rules about how electrons fill up the electron shells, the inner shell is always filled first. Each shell has a maximum number of electrons it can take.

Shell 1: maximum 2 electrons

Shell 2: maximum 8 electrons

Shell 3: maximum 8 electrons

Example:



The electronic configuration of Sodium (Na) can also be written like this 2,8,1. This shows there is 2 electrons in the 1st shell, 8 electrons in the second shell and 1 electron in the 3rd shell.

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C2 - The periodic table

The History of the Periodic Table

- Throughout history scientists have tried to classify substances and many scientists attempted to construct a Periodic Table.
- Before the knowledge of protons, neutrons and electrons, scientists arranged the Periodic table by **atomic weight**. This meant the groups were not always correct.
- In 1869 Dimitri **Mendeleev**, a Russian Scientist, published his Periodic Table. It was slightly different to those that had been before. He still arranged elements by atomic weight but he also left gaps for where he predicted elements would be.
- He very accurately predicted the properties of elements that were not discovered until many years later; for example, Gallium.
- Mendeleev's Periodic Table is still different from the modern one as some of his masses were wrong due to the existence of **isotopes**
- Isotopes are elements with same number of protons and electrons but a different number of neutrons and therefore different atomic weights.

Isotopes of Carbon

		
^{12}C	^{13}C	^{14}C
Carbon-12	Carbon-13	Carbon-14
6 protons 6 neutrons	6 protons 7 neutrons	6 protons 8 neutrons

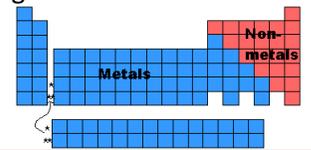
Mendeleev's Periodic Table

			Tl = 50	Zr = 90	? = 180
			V = 51	Nb = 94	Ta = 182
			Cr = 52	Mo = 96	W = 186
			Mn = 55	Rh = 104,4	Pt = 197,4
			Fe = 56	Ru = 104,4	Ir = 198
			Ni = 59	Pd = 106,6	Os = 199
			Cu = 63,4	Ag = 108	Hg = 200
			Zn = 65,2	Cd = 112	
			? = 68	U = 116	Au = 197?
			? = 70	Sr = 118	
			As = 75	Sb = 122	Bi = 210?
			Se = 79,4	Te = 128?	
			Br = 80	J = 127	
			K = 39	Rb = 85,4	Cs = 133
			Ca = 40	Sr = 87,6	Ba = 137
			? = 45	Ce = 92	Pb = 207
			? = 56	La = 94	
			? = 60	Di = 95	
			Zn = 75,6	Th = 118?	

Key Terms	Definitions
Dimitri Mendeleev	A Russian Chemist, who in 1869 published a Periodic Table containing gaps.
Periodic Table	The table which organises the 118 elements based on atomic structure
Isotope	Two atoms with the same number of protons and electrons but a different number of neutrons
Metal	An element which loses electrons to form a positive charge
Non Metal	An element which gains electrons to form a negative charge
Ion	An element with a positive or negative charge

Metals and Non-Metals

- Metals are found on the left hand side of the Periodic Table, the majority of elements are metals.
- When metals react, they lose an electrons to form positive ions.
- Non metals gain electrons to form a negative charge.



Groups in the Periodic Table

	Physical properties	Chemical Properties	Equation	Trends/Explanation
Group 1 (Alkali metals)	Soft, low density	React vigorously with water releasing hydrogen	Sodium + Water → Sodium Hydroxide + Hydrogen	More reactive as you go down, outermost electron further from the nucleus so it's easier to lose
Group 7 (Halogens)	Low melting point, exist as pair (Cl ₂)	React with group 1 metals to form compounds. Can carry out displacement reactions	Sodium + Chlorine → Sodium Chloride Sodium Bromide + Chlorine → Sodium Chloride + Bromine	Higher melting point as you go down the group (higher molecular mass). Less reactive as you go down the group.
Group 0 (Noble Gases)	Low melting point/boiling point Eight electrons in outer shell (except helium)	Unreactive, as they have a full outer shell	N/A	Higher melting point and boiling point as you go down the group (due to increase in density)

Chemistry Knowledge Organiser

C2 - The periodic table

Transition Metals Continued

- Transition metals also differ from group 1 elements as they can form multiple different ions (sometimes called oxidation states). Elements in group 1 can only form a +1 ion.
- The ability to form different ions, gives transition metals other properties, firstly it makes them good catalysts in chemical reactions, see more on this in the rate of reaction topic.
- Different ions also form different coloured compounds for example if vanadium forms a +3 ion it is green, +4 is blue and +5 is yellow. This means transition metals are often used in paints.
- The table below shows the ions that different period 4 transition metals can form. You are not expected to memorise this table:

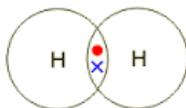
Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
	+2	+2	+2	+2	+2	+2	+2	+2	+2
+3	+3	+3	+3	+3	+3	+3	+3	+3	
	+4	+4	+4	+4	+4	+4	+4		
	+5	+5	+5	+5	+5	+5			
			+6	+6	+6				
				+7					

Chemistry Knowledge Organiser

C3 - Structure and bonding

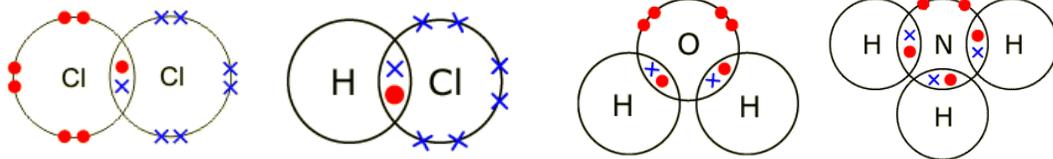
Covalent Bonding

Covalent bonding occurs between non metals. **Electrons 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.

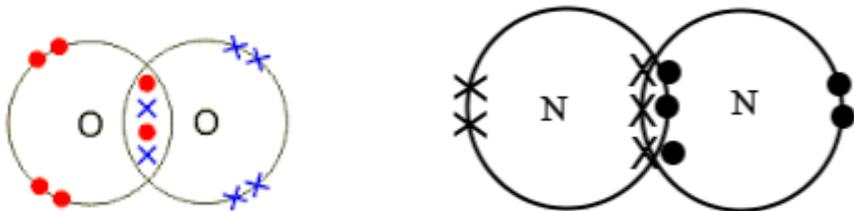


When drawing covalent molecules we use “dot cross diagrams” as we do with ionic compounds. It is important to represent the electrons on one atom with a dot and on the other atom with an X.

The first five examples, **hydrogen, chlorine, water, hydrogen chloride and ammonia (NH₃)** all share one electron per atom in a to make a full outer shell of electrons on each atom.



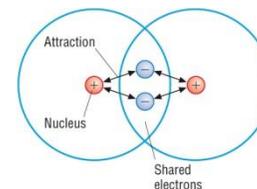
Some atoms need more than one electron to give them a full outer shell, for example oxygen needs 2 electrons to complete its outer shell. Oxygen therefore shares two electrons per atom to **make a double bond**. Nitrogen needs three electrons to complete its outer shell, this forms a triple bond between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



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

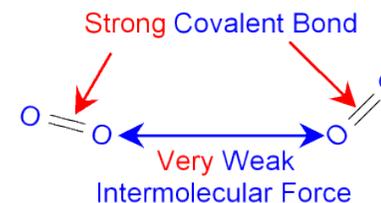
The Nature of a Covalent Bond

Covalent bonds are strong because there is electrostatic 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.



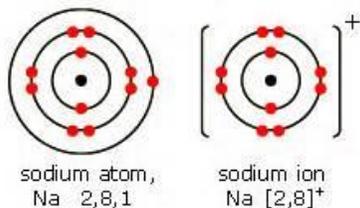
Please see the next page for more properties of covalent compounds.

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 its 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.

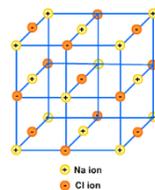


The **group number** indicates how many electrons an atom would have to lose or gain to get a full outer shell of electrons. See below to see what ions different groups form

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

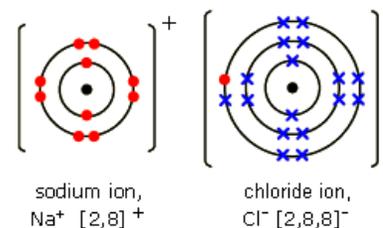
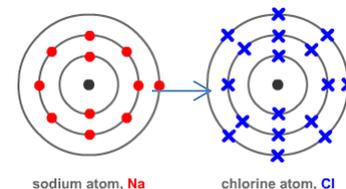
Ionic compounds have **regular structures (giant ionic lattices)** in which there are strong **electrostatic forces** of attraction in all directions between oppositely charged ions.



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.



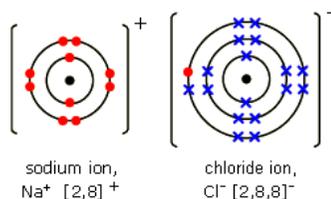
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C3 - Structure and bonding

Ionic Bonding- Models

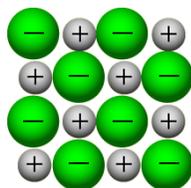
There are a number of ways we can represent ionic bonding all; of these have **advantages and limitations**. For example all the diagrams below show ways we can represent **sodium chloride**

- 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.



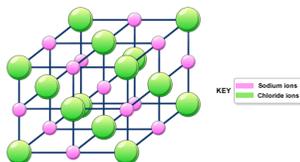
- 2. 2D ball and stick model of ionic bonding**

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.



Key Terms	Definitions
Ionic Lattice	The regular 3D arrangement of ions in an ionic compound
Giant	When the arrangement of atoms is repeated many 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

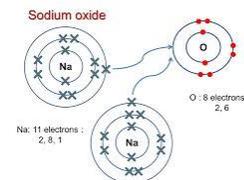
Properties of ionic compounds

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 Na₂O**.



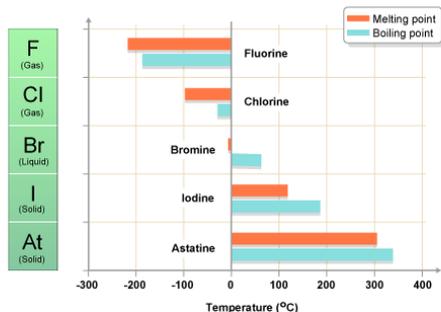
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C3 - Structure and bonding

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**.



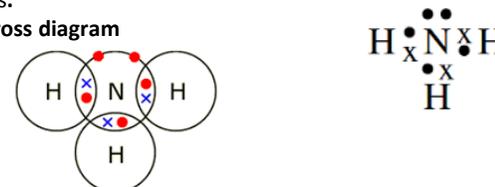
As well as having low melting points, covalent compounds **do not conduct electricity**. This is because they have no free electrons or ions and therefore there is nothing to carry the electric charge. Remember pure water does not conduct electricity, only when it has ions dissolved in it will it conduct.

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

Representing Covalent Compounds

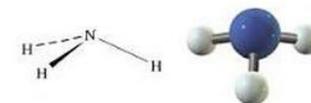
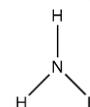
Like ionic compounds, there are a variety of ways that scientists use to represent covalent compounds.

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

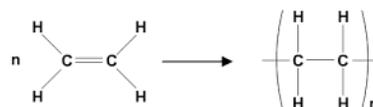


Ammonia (Ball and stick model)

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.



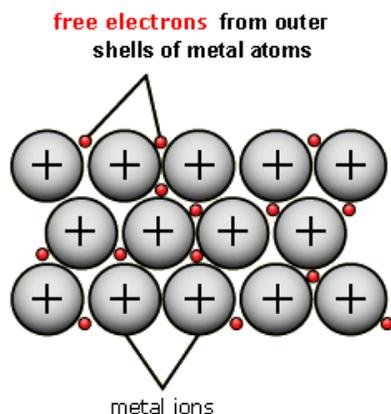
n = a big number of monomers

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C3 - Structure and bonding

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+} .

Properties of Metals

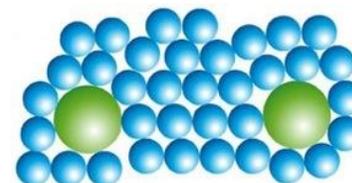
Metals are **good conductors of electricity**, due to the delocalised electrons, which can carry the electric charge. Metals are also **good conductors of heat** as the free electrons can transfer the heat energy through the metal.

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.

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.

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.



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:

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

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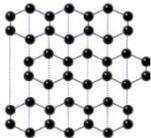
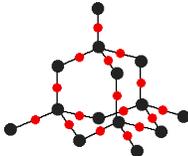
C3 - Structure and bonding

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.

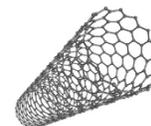
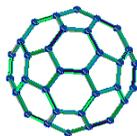
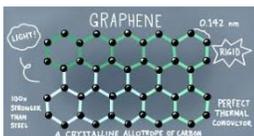
There are three examples you need to know, diamond, graphite and silica (see table below)

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

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.
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 . 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

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 used 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.

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: 1×10^{-9} m
Catalyst	A substance which speeds up a reaction, without being used up.

Dangers on Nanomaterials

The long term affects of nanomaterials on the body have not been well researched. For example when using sunscreen, 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.

Chemistry Knowledge Organiser

C4 - Chemical calculations

Relative formula mass (M_r):

This is the mass in grams of 1 mole of the substance. To calculate it you need to add up the atomic mass (bigger number) of all of the atoms in the molecule.

e.g 1. $NaCl = Na + Cl = 23 + 35.5 = 58.5$

e.g 2. $MgF_2 = Mg + (2 \times F) = 24 + (2 \times 19) = 62$

The Mole

A mole of an element is simply **6.02×10^{23} atoms (this number is known as Avogadro's number)**. Obviously, if the atoms are larger than 1 mole of that atom will be heavier. For example, one mole of hydrogen atoms weighs 1 gram but 1 mole of carbon weighs 12 grams. To calculate the number of moles in an element you need to divide the mass by the relative atomic mass:

For example, how many moles are there in 6 grams of carbon?

$$6/12 = 0.5$$

To work out the number of moles in a compound you divide the mass of the compound by the relative formula mass, for example how many moles in 30 grams of magnesium oxide (MgO)?

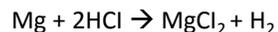
$$M_r \text{ of } MgO = 24 + 16 = 40$$

$$\text{Moles} = 30/40 = 0.75 \text{ Moles}$$

HT: Calculating Masses in Reactions

An understanding of the mole will allow to calculate the mass made in a chemical reaction.

Take the chemical reaction below:



This equation shows that one mole of magnesium reacts with two moles of hydrochloric acid to produce one mole of magnesium chloride and one mole of hydrogen gas. Suppose you started with 5 grams of magnesium, how much magnesium chloride would you make?

Step 1: Calculate the moles of the element or compound you were given in the equation:

$$5/24 = 0.21 \text{ moles of magnesium}$$

Step 2: Look at the balanced equation, you must therefore have 0.21 moles of magnesium chloride, as the ratio between magnesium and magnesium chloride is 1 to 1.

Step 3: Calculate the M_r of the relevant product: what you want to find is the M_r of magnesium chloride:

$$M_r \text{ of } MgCl_2 = 24 + 35.5 + 35.5 = 94$$

Step 4: Now find the mass of that number of moles of the product

$$\text{Mass} = \text{moles} \times M_r, \text{ so } 0.21 \times 94 = 19.7 \text{ grams}$$

Key Terms	Definitions
Mole	6.02×10^{23} atoms of an element or molecules in a compound
Avogadro's number	6.02×10^{23}
Relative Formula Mass	The total atomic mass of elements in compound

Equation	Meanings of terms in equation
$\text{moles} = \frac{\text{mass}}{M_r}$	<i>Mass is the mass of the substance in grams M_r is the relative formula mass of the compound (or use the relative atomic mass if it is an element)</i>

Calculating moles from masses -Higher Tier

If you know the mass of each reactant and product you can calculate a balanced equation from the masses, for example: Calculate the balanced equation when 12 grams of magnesium reacts completely with 19.25g of HCl, to make 99 grams of $MgCl_2$ and 1 gram of H_2



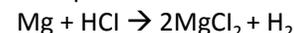
Step 1: work out the moles of each reactant and product.

$$Mg = 12/24 = 0.5 \quad HCl = 19.25/38.5 = 0.5 \quad MgCl_2 = 99/99 = 1 \quad H_2 \frac{1}{2} = 0.5$$

Step 2 divide through by the smallest number

$$Mg = 0.5/0.5 = 1 \quad HCl = 0.5/0.5 = 1 \quad MgCl_2 = 1/0.5 = 2 \quad H_2 \frac{1}{2} = 0.5/0.5 = 1$$

Step 3 write the balanced equation:



Chemistry Knowledge Organiser

C4 - Chemical calculations

% Yield

In reactions in chemistry it is very rare that we make the exact mass predicted by calculation, for a variety of reasons we often make a lot less. The equation to calculate percentage yield is outlined in the equation box. **A percentage yield is always 100% or less**, the law of conservation of mass states that we cannot make mass in a chemical reaction.

It is extremely rare that the yield of a chemical reaction is 100% reasons for this are:

- The reaction is reversible and may not go to completion
- There may be side reactions
- Some may be lost when the product is transferred from the reaction vessel

Atom Economy

Some reactions make more than one product, atom of these products will be waste products. The atom economy is a measure of the atoms that form useful products. Like percentage yields we express atom economy as a percentage so that comparisons can be easily made between reactions.

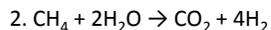
For example below two ways of making hydrogen are outlined:



M_r of $\text{H}_2 = 1 + 1 = 2$

M_r of $\text{ZnCl}_2 = 65 + 35.5 + 35.5 = 136$

Atom economy = $\frac{2}{136 + 2} \times 100 = \frac{2}{138} \times 100 = 1.45\%$ Very low atom economy



M_r of $\text{H}_2 = 1 + 1 = 2$

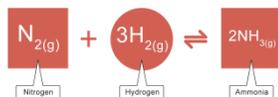
M_r of $\text{CO}_2 = 12 + 16 + 16 = 44$

Atom economy = $\frac{4 \times 2}{44 + (4 \times 2)} \times 100 = \frac{8}{52} \times 100 = 15.4\%$ Higher atom economy

In the second example the atom economy is higher, therefore in terms of atom economy reaction 2 is better. Chemists often need to balance atom economy and percentage yield. A poor atom economy is bad for a number of reasons:

1. A lot of reactant is wasted, this costs money.
2. The waste products have to be disposed of, this can be expensive. Some companies try to get around this problem by reusing the waste product.

The best reactions in terms of atom economy are those that only make one product, for example the Haber process, the atom economy here is 100%:



Key Terms	Definitions
Yield	The amount of product made in a chemical reaction
Atom Economy	The percentage of atoms that form useful products
Limiting Reagent	The reagent which is used up first in a chemical reaction.

Equation	Meanings of terms in equation
$\% \text{Yield} = \frac{\text{Mass of products}}{\text{Mass of theoretical products}} \times 100$	<i>Mass of products is the mass made in a chemical reaction</i> <i>Mass of theoretical products is the mass we expected to make based on calculations</i> <i>Both these masses must be in the same unit.</i>
$\text{Atom economy} = \frac{\text{Relative formula mass of DESIRED products}}{\text{Total relative formula mass of products}} \times 100$	<i>See the previous page for how to calculate relative formula mass</i>

Limiting Reagent

When a chemical reaction is carried out, one or more reagents are in excess and one reagent is the limiting reagent. The **limiting reagent** is the reagent which is used up first in a chemical reaction, if all of this reagent is used up the reaction can no longer continue, for example, if a tiny amount of sodium is dropped into a large bowl of water there are a lot more water particles than there are sodium atoms. We therefore say that the sodium is the **limiting reagent** and the water is in **excess**.

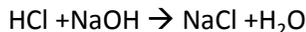
The amount of product formed is **directly proportional** to the amount of limiting reagent. Therefore if you double the amount of limiting reagent you will get double the amount of product.

Chemistry Knowledge Organiser

C4 - Chemical calculations

Titration Calculations

We can use the information that we get from a titration to work out the concentration of an alkali or acid. For example a titration was carried out using hydrochloric acid and sodium hydroxide, the equation for this reaction is:



This means that one mole of hydrochloric acid will neutralise 1 mole of sodium hydroxide.

Therefore we can calculate the following:

27.5 cm³ of 0.2 mol/dm³ hydrochloric acid is needed to titrate 25.0 cm³ of sodium hydroxide solution. What is the concentration of the sodium hydroxide solution?

Step 1: Convert all volumes to dm³

$$27.5 \text{ cm}^3 = 27.5 \div 1000 = 0.0275 \text{ dm}^3$$

$$25.0 \text{ cm}^3 = 25.0 \div 1000 = 0.025 \text{ dm}^3$$

Step 2: Calculate the number of moles of the substance where the volume and concentration are known

$$\text{number of moles} = \text{concentration} \times \text{volume}$$

$$\text{number of moles of hydrochloric acid} = 0.2 \times 0.0275 = 0.0055 \text{ mol } (5.5 \times 10^{-3} \text{ mol})$$

Step 3: Calculate the unknown concentration

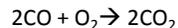
We can say that 0.0055 mol of acid will react with 0.0055 mol of alkali

$$\text{concentration of alkali} = \text{moles} \div \text{volume} = 0.0055 \div 0.025 = \mathbf{0.22 \text{ mol/dm}^3}$$

Moles of a Gas

We know that one mole of any gas occupies **24 dm³ at room temperature and pressure**. Room temperature and pressure is defined as 20 °C and 1 atm of pressure. Therefore if we know the volume that the gas occupies, we can divide this number by 24 and this will give us the number of moles. For example if we had 12 dm³ of argon gas at room temperature and pressure, then to find out the moles we would simply do $12/24 = 0.5$ moles.

We can also use balanced equations to work out volumes of gas, for example:



If we start with 24 dm³ of oxygen we will make 48 dm³ of carbon dioxide, as the ratio in the equation shows that one mole of oxygen will make 2 moles of carbon dioxide.

You may also be asked to calculate the volume a gas occupies after being given the mass in the equation. For Example: What volume would 2 grams of carbon dioxide occupy at room temperature and pressure?

Step 1: Calculate the moles of carbon dioxide = $2/44 = 0.05$ moles

Step 2: Multiply this number by 24 as we know 1 mole occupies 24 dm³ $0.05 \times 24 = 1.2 \text{ dm}^3$

Key Terms	Definitions
Decimetre	A unit of volume, often used in chemistry. It is the same as 1000 cm ³
Atmosphere	A unit of gas pressure. It is the same as 101325 Pa

Equation	Meanings of terms in equation
Volume = $\frac{\text{Mass of gas}}{\text{Mr of gas}} \times 24$	Mass of gas in grams Volume of gas will be in dm ³

Limiting Reagent

When a chemical reaction is carried out, one or more reagents are in excess and one reagent is the limiting reagent. The **limiting reagent** is the reagent which is used up first in a chemical reaction, if all of this reagent is used up the reaction can no longer continue, for example, if a tiny amount of sodium is dropped into a large bowl of water there are a lot more water particles than there are sodium atoms. We therefore say that the sodium is the **limiting reagent** and the water is in **excess**.

The amount of product formed is **directly proportional** to the amount of limiting reagent. Therefore if you double the amount of limiting reagent you will get double the amount of product.

Chemistry Knowledge Organiser

C4 - Chemical calculations

Concentration

Most chemical reactions are done in solution. The concentration can be measured in grams per dm^3

For example what is the concentration in grams/dm^3 of 2.4 grams of sodium chloride dissolved in 0.5 dm^3 of water?

Conc= Mass/Vol

Conc= $2.4/0.5$

Conc= $4.8 \text{ g}/\text{dm}^3$

In Chemistry we use dm^3 (decimetres cubed) to measure volume, a decimetre cubed is the same as a litre or 1000 cm^3 .

However it is far more common to calculate a concentration in moles per dm^3 . This is sometimes written as M.

For example, what is the concentration of 2.4 grams of sodium chloride dissolved in 0.5 dm^3 in mol/dm^3 ?

Moles of NaCl= $2.4/58.5 = 0.041$ moles

Conc=moles/vol

Conc= $0.041/0.5$

Conc= $0.082 \text{ mol}/\text{dm}^3$

It is also possible to convert between mol/dm^3 and g/dm^3 for example. If I had $0.5 \text{ mol}/\text{dm}^3$ HCl solution. We can work out the concentration in g/dm^3 :

Step 1: Work out the Mr of HCl: $35.5+1 = 36.5$

Step 2= Mass= Moles x Mr= $36.5 \times 0.5 = 18.25 \text{ g}/\text{dm}^3$

Titration

Titration is used to find out an **unknown concentration of a solution**, this is often used to find out the concentration of an acid or an alkali in a neutralisation reaction. To carry out a titration to find the concentration of an alkali you need to do the following:

1. A pipette is used to measure 25 cm^3 of alkali, this is then transferred to a conical flask.
2. 3-4 Drops of indicator is added (phenolphthalein).
3. An acid of known concentration is placed in the burette
4. The solution from the burette is allowed to slowly run into the conical flask. As the end point approaches the acid is added one drop at a time. When phenolphthalein is used as an indicator, the end point is where the solution turns from colorless to pink.
5. The volume of acid used from the burette is noted to calculate the concentration of the alkali in the conical flask. See the next page for how to carry out these calculations.

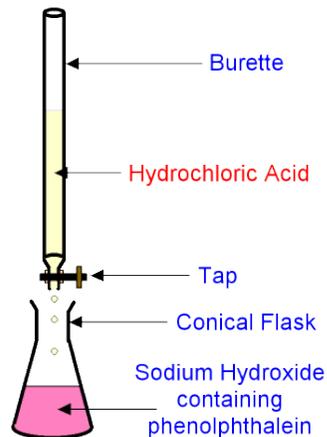
Key Terms	Definitions
Concentration	A measure of the number of moles or mass in a given volume.
Titration	An experimental techniques where unknown concentrations of solutions can be found.
Burette	A piece of apparatus used to accurately measure volumes of solution.

Equation	Meanings of terms in equation
$\text{conc} = \frac{\text{mass}}{\text{vol}}$	<i>Mass is the mass of the solute in grams</i> <i>Vol is the volume of the solvent in dm^3</i> <i>Conc is the concentration in grams/dm^3</i>
$\text{conc} = \frac{\text{Moles}}{\text{volume}}$	<i>Moles is the number of moles of the solute</i> <i>Vol is the volume of the solution in dm^3</i> <i>Conc is the concentration in grams/dm^3</i>

Indicators

For titrations universal indicator is not a suitable indicator to use. As the colour changes are too gradual. For a titration, a sharp colour change is required. Suitable indicators are listed below.

	In acid	In alkali
Litmus	Red	Blue
Methyl Orange	Red	Yellow
Phenolphthalein	Colourless	Pink



Chemistry Knowledge Organiser

C5 - Chemical changes

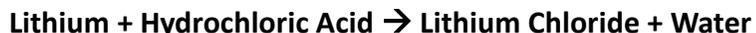
Reactions of Metals

When a metal reacts with water it produces a metal hydroxide and **hydrogen gas**.

The more reactive the metal is, the more vigorous the reaction. For example:



You see a similar pattern for the reaction between metals and acids however the products in these reactions are different, in this case you will make a salt and water, the salt will depend on the type of acid that you have used.



If sulphuric acid is used the salt made will be a sulphate, if nitric acid is used the salt will be a nitrate.

Metals also react with oxygen to form metal oxides; in this reaction the metal donates electrons to the oxygen. This means the metal is **oxidised as it has lost electrons**. The oxygen is reduced as it has gained electrons.

Extraction of Metals

A metal ore is a compound found in rock, dug out of the ground, that contains enough metal that it is **economical** to extract it. For example, magnesium oxide. In order for us to use the magnesium we need to **extract** it from the oxide.

Metals more reactive than carbon are extracted from their ore using **electrolysis**.

Metals which are less reactive than carbon are extracted from their ore using **reduction** (by adding carbon). Reduction is the removal of oxygen as seen in the example.



The least reactive metals such as gold and silver are found on their own—they do not form a compound. This means they do not need to be extracted from their ore.

Key Terms	Definitions
Oxidation	The loss of electrons from an atom OR when an atom gains an oxygen atom
Reduction	The opposite to oxidation, when an atom gains electrons OR when an atom loses an oxygen atom
REDOX Reaction	A reaction where one atom is oxidised and another atom is reduced

Other methods of extraction

The amount of some metals is running out, this means people are finding new ways to extract metals like copper.

Phytomining uses plants to absorb copper from the soil, the plants are then burnt and the copper extracted.

Bioleaching involves using bacteria to make a **leachate** that contains metal compounds.

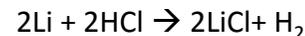
Scrap iron can also be used to **displace copper** from a solution.

Oxidation Reactions

When working out whether a reaction is oxidation or reduction: in terms of electrons, remember OILRIG. This stands for oxidation is loss and reduction is gain.

HT - Oxidation Reactions of Acids

When an acid reacts with a metal a salt and hydrogen are produced. For example the symbol equation for an acid reacting with lithium is:



In this reaction, lithium has been oxidised because it has lost an electron to form a **+1 ion** and hydrogen has been reduced from a +1 ion to a **hydrogen molecule**.

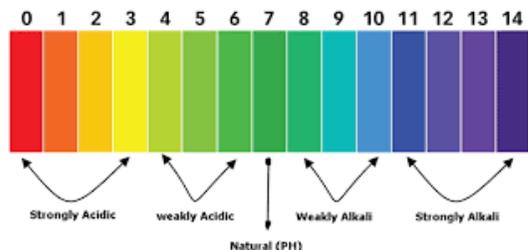
Chemistry Knowledge Organiser

C5 - Chemical changes

Acids and Alkalis

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

We measure the acidity of a substance using the **pH scale which runs from 0-14** between 0 and 6 the substances are acidic, 7 is neutral and between 8 and 14 is alkaline. The pH scale is a logarithmic scale: a decrease of 1 on the pH scale makes a substance **10 times more acidic**.



The pH scale is a measure of H^+ concentration: the lower the pH the higher the concentration of H^+ ions.

Key Terms	Definitions
Acid	A substance which forms H^+ ions in aqueous solution
Alkali	A substance which forms OH^- ions when dissolved: these are soluble bases
Neutralisation	A reaction between an acid and an alkali making a salt and water
Strong Acid	An acid which totally dissociates in water
Base	A substance that can neutralise an acid to make a salt and water

Neutralisation

To work out the names and formulae of salts you will need to know the names and formulae of the common acids

Acid	Name of salt	Ion that forms salt
Hydrochloric	Chloride	Cl^-
Sulphuric Acid	Sulphate	SO_4^{2-}
Nitric Acid	Nitrate	NO_3^{1-}

Neutralisation

When an acid reacts with an alkali a salt and water are produced. The ionic equation for the reaction of an **acid and an alkali is:**



HT - Strong and Weak Acids

Acids can be defined as either a **strong or weak acid** a strong acid is one which fully dissociates in water for example hydrochloric acid

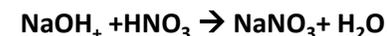


A weak acid is defined as one which only partially dissociates in water.

Strong acids are **not the same** as concentrated acids. Concentration is the number of particles in a given volume and not how much they dissociate.

Neutralisation

When an acid reacts with an alkali it will produce salt and water, below are the general equations for different types of neutralisation reaction:



Some of the reactants (for example copper oxide) are insoluble but these can still carry out a neutralisation reaction. We call these **bases** not **alkalis**.

Chemistry Knowledge Organiser

C6 - Electrolysis

Extracting Aluminium

Aluminium oxide is dissolved in molten cryolite .

Cryolite reduces the melting point of aluminium oxide meaning the process requires less energy.

Aluminium ions (Al³⁺) are attracted to the negative electrode.

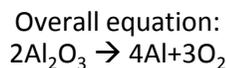
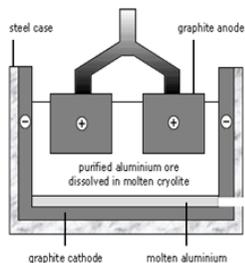
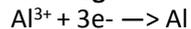
Aluminium atoms are formed at the negative electrode (gain 1 electron)

Oxide ions are attracted to the positive electrode

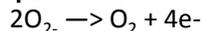
Oxygen is formed at the positive electrode (each ion loses 2 electrons)

Oxygen reacts with carbon to make carbon dioxide. This electrode needs to be replaced constantly.

At the negative electrode:



At the positive electrode



Electrolysis of Brine

Which elements form at which electrode depends on the **reactivity** of the elements involved.

For example, the electrolysis of brine, is the electrolysis of a solution of sodium chloride, however there are also H⁺ and OH⁻ ions from the water which is used as the solvent. This means there is more than one possible ion that can go to each electrode.

· **Positive ions:** sodium (Na⁺) and hydrogen (H⁺)

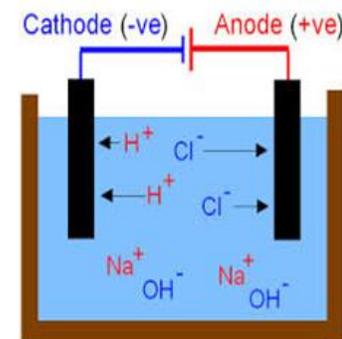
· **Negative ions:** chlorine (Cl⁻) and hydroxide (OH⁻)

When there is a mixture of ions, the products formed depend on the reactivity of the elements involved.

Hydrogen is less reactive than sodium, so hydrogen gas (H₂) is produced at the negative electrode.

Chlorine gas (Cl₂) is produced at the positive electrode.

Sodium hydroxide is produced from the ions that remain in solution.



Gas Tests

During electrolysis the products made are often gases. Below are the tests for three common gases you need to know:

Gas	Test	Result
Hydrogen	Place a lit splint into the gas	If a squeaky pop is heard hydrogen is present
Oxygen	Place glowing splint into gas	If splint is relighted then oxygen is present
Chlorine	Damp litmus paper placed in gas	If paper bleaches, chlorine is present
Carbon Dioxide	Bubble the gas through limewater	If the limewater goes cloudy carbon dioxide is present

Chemistry Knowledge Organiser

C6 - Electrolysis

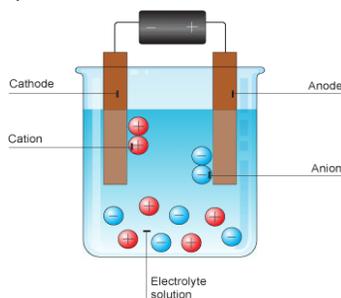
Electrolysis

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**.

If an electric current is passed through this solution the ions will move to the electrodes.

Remember-opposites attract. The positive ions (cations) will go to the negative electrode (cathode), the negative ions (anions) go to the positive electrode (anode).

For example in the electrolysis of lead bromide, Lead (Pb^{2+}) goes to the negative electrode and bromine (Br^{-1}) goes to the positive electrode.



Electrolysis of Copper Sulphate

Which elements form at which electrode depends on the **reactivity** of the elements involved. For example, in the electrolysis of aqueous copper sulphate is the electrolysis of copper sulphate, however there are also H^{+} and OH^{-1} ions from the water which is used as the solvent. This means there is more than one possible ion that can go to each electrode.

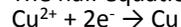
· **Positive ions:** sodium (Cu^{2+}) and hydrogen (H^{+})

· **Negative ions:** sulphate (SO_4^{2-}) and hydroxide (OH^{-})

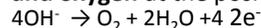
When there is a mixture of ions, the products formed depend on the reactivity of the ions involved.

Copper is **less reactive** than hydrogen, so copper (Cu) is produced at the negative electrode.

The half equation is:



The hydroxide ion is more reactive than the sulphate ion, therefore this **forms water (H_2O) and oxygen** at the positive electrode.



As a rule if a halide ion is present, this will form at the positive electrode, however if no halide is present then oxygen and water will form at the positive electrode.

Key Terms	Definitions
Electrolysis	The breaking down of a substance using electricity
Electrolyte	The solution which is being broken down during electrolysis
Oxidation	The loss of electrons
Reduction	The gain of electrons
Anode	The positive electrode
Cathode	The negative electrode
Half Equation	An equation that shows the reaction at each electrode

Oxidation and reduction

When a positive ion reaches the negative electrode, it gains electrons. This is a reduction reaction.

When the negative ion reaches the positive electrode, it loses electrons, this is an oxidation reaction.

We can represent these using half equations A half equation can represent the reaction at each electrode. Half equations show how electrons are transferred and an electron is represented in an equation by an e^{-} symbol

Half equations show electrons (e^{-}) and how ions become atoms.

For example $\text{Cu}^{2+} + 2\text{e}^{-} \rightarrow \text{Cu}$.

1. Write down the ion and atom: $\text{Cl}^{-} \rightarrow \text{Cl}_2$

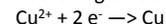
2. Adjust the number of ions (if needed) and add electrons to balance the charges if required $2\text{Cl}^{-} \rightarrow \text{Cl}_2 + 2\text{e}^{-}$

Remember that non metal ions will typically form diatomic molecules.

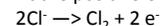
Ionic equations

Half equations can be combined to form an ionic equation, which shows the overall reaction. For example in the electrolysis of copper chloride the two half equations are:

At the negative electrode (cathode):

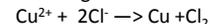


At the positive electrode (anode):



Combining these 2 equations gives us:

$\text{Cu}^{2+} + 2\text{e}^{-} + 2\text{Cl}^{-} \rightarrow \text{Cu} + \text{Cl}_2 + 2\text{e}^{-}$ The electrons either side of the equation cancel out, meaning the final ionic equation is:



In an ionic equation it is important to check both the atoms and the charges balance

Chemistry Knowledge Organiser

C7 - Energy changes

Energy in 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. In a chemical reaction, bond breaking and bond making occur. To break a chemical bond you need to overcome the force of attraction in the bond, this process requires energy therefore it is **endothermic**. The process of bond formation is **exothermic**, energy is released when bonds form. In a chemical reaction the difference between the energy required to break the bonds and the energy gained from making the bonds will decide whether a reaction is exothermic or endothermic.

Chemical reactions can therefore be divided into exothermic and endothermic chemical reactions.

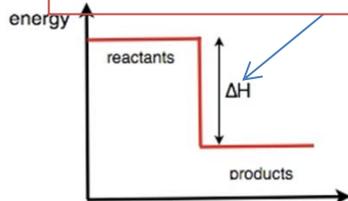
	What happens?	Why?	Example
Exothermic	Heat energy is transferred to the surroundings.	The energy required to break chemical bonds is less than the energy gained from making chemical bonds. Therefore the excess is given off as heat to the surroundings.	Combustion reaction, reactions used in hand warmers
Endothermic	Heat energy is taken in from the surroundings	The energy required to break chemical bonds is more than the energy gained from making chemical bonds. Therefore heat is taken in from the surroundings.	The reaction of citric acid and sodium hydrogencarbonate, the reactions used in ice packs

Reaction Profiles

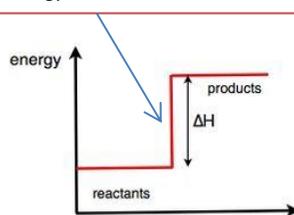
Reminder from topic 15: Chemical reactions can occur only when reacting particles collide with each other and with sufficient energy. The minimum amount of energy that particles must have to react is called the **activation energy**.

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.

This is the reaction profile of an **exothermic reaction**, the energy of the products is lower than that of the reactants. The difference in energy is released as heat to the surroundings.



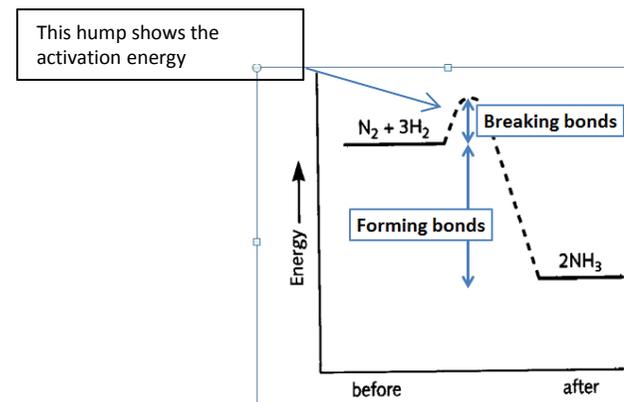
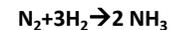
This is the reaction profile of an **endothermic reaction**, the energy of the products is higher than that of the reactants. The difference in energy is taken in from the surroundings.



Key Terms	Definitions
Reaction Profile	A graph which shows the energies of the products and reactants in a chemical reaction
Exothermic	A reaction that gives out heat to the surroundings
Endothermic	A reaction that takes heat in from the surroundings

Reaction Profiles- In more detail

The profile below shows the reaction which makes ammonia from nitrogen and hydrogen. The equation is given below:

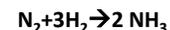


There are some key features to highlight on this graph, firstly the curved section represents the **activation energy** for this reaction, this hump shows how much energy is required to break the bonds in the reactants. To overcome the activation energy we often need to heat our reactants. The products are lower in energy than the reactants, this means it is an **exothermic reaction**. As the excess energy is given out to the surroundings, as **heat energy**.

Calculating bond energies -higher tier.

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.

For example consider the reaction:



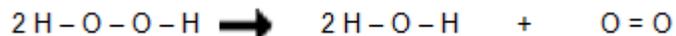
To work out the overall energy change you will need to subtract, the energy gained from forming the bonds in ammonia, from the energy required to break the nitrogen and hydrogen bonds. This will give you the overall energy change, if the value is negative then the reaction is exothermic, if the value is positive the reaction is endothermic

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C7 - Energy changes

Bond Energies continued- Higher Tier

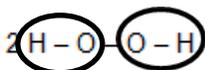
You can calculate the energy change in a reaction from bond energies given to you in a question. For example consider the reaction below:



This shows that hydrogen peroxide breaks down to make water and oxygen. We can use bond energies to work out the energy change in the reaction.

Bond	Bond energy in kJ per mole
H-O	464
O-O	146
O=O	498

The energy required to break the reactant bonds is:



2×464 (for the O-H bonds) = 928 + 146 (for the O-O bond) = 1074 however as there is a 2 in the equation this number needs to be doubled.

$$2 \times 1074 = 2148 \text{ kJ/mol}$$

The energy gained from making the product bonds is:



$2 \times 464 = 928$ but there is a 2 in the equation so this doubled to 1856 and we also need to add the 498 for the double bond in O_2

$$1856 + 498 = 2354 \text{ kJ/mol}$$

Therefore we do energy required to break reactant bonds - energy gained from making product bonds:

$$2148 - 2354 = -170 \text{ kJ/mole}$$

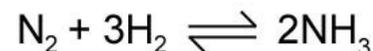
If the value is negative then the reaction is **exothermic**

If the value is positive the reaction is **endothermic**.

Key Terms	Definitions
Equilibrium	A reaction that is reversible
Le Chatelier's principle	A principle which states, "If a system is at equilibrium and a change is made to any of the conditions, then the system responds to counteract the change "
Dynamic Equilibrium	An equilibrium where the forward and backward reactions are happening at the same rate

Equilibrium

Some chemical reactions are reversible, this means they can happen in both the **forward and reverse directions**. The symbol we use to represent an equilibrium reaction is shown in the equation below:



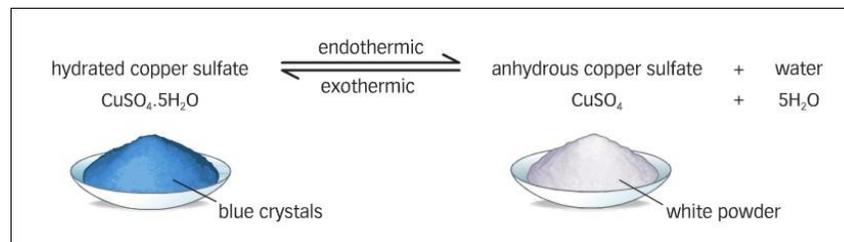
In a dynamic equilibrium reaction, the forward and reverse reactions are happening at the **same rate**.

A dynamic equilibrium has to occur in a **closed system**, where no reactants and products are allowed to escape.

If the equilibrium lies to the left, it means that there is a **greater concentration of reactants than products**

If the equilibrium lies to the right it means there is a **greater concentration of products than reactants**.

Most equilibrium reactions are endothermic in one direction and exothermic in another direction. A good example is the hydration and dehydration of copper sulphate. It is exothermic when water is added to the copper sulphate, it is endothermic when water is removed.



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C8 - Rates and equilibrium

Changing Conditions-Le Chatelier's principle- Higher Tier

The Haber process is a good example to explain Le Chatelier's principle, the equation for the Haber process is shown below. The reaction is carried out in the gaseous state. Remember this is one of many reactions but the principles always stay the same.

Endothermic in this direction $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$ Exothermic in this direction

Condition Change	Effect
Increase the temperature	Shifts the equilibrium to the left as this is the endothermic direction. The amount of reactants increases.
Decrease the temperature	Shifts the equilibrium to the right as this is the exothermic direction. The amount of product increases
Increase the concentration of reactants	Equilibrium shifts to the right to make more product, to reach equilibrium again
Increase the concentration of products	Equilibrium shifts to the left to reach equilibrium again
Increase the pressure in the gas	Equilibrium shifts to the right, where there are fewer molecules of gas, this will decrease the pressure.
Decrease the pressure in the gas	Shifts the equilibrium to the left as there are more gas molecules on that side of the equation.

Key Terms	Definitions
Equilibrium	A reaction that is reversible
Le Chatelier's principle	A principle which states, "If a system is at equilibrium and a change is made to any of the conditions, then the system responds to counteract the change "
Dynamic Equilibrium	An equilibrium where the forward and backward reactions are happening at the same rate

Equilibrium- Changing Conditions-Higher tier

The amounts of all the reactants and products at equilibrium depend on the conditions of the reaction. For example if we change things like **temperature, concentration of a reactant or product and pressure in gases.**

The French scientist Le Chatelier devised a principle to explain how equilibrium reactions, respond to a change in conditions, it states that:

"If a system is at equilibrium and a change is made to any of the conditions, then the system responds to counteract the change"

For example if the temperature is raised the equilibrium will shift to try to cool the surroundings down.



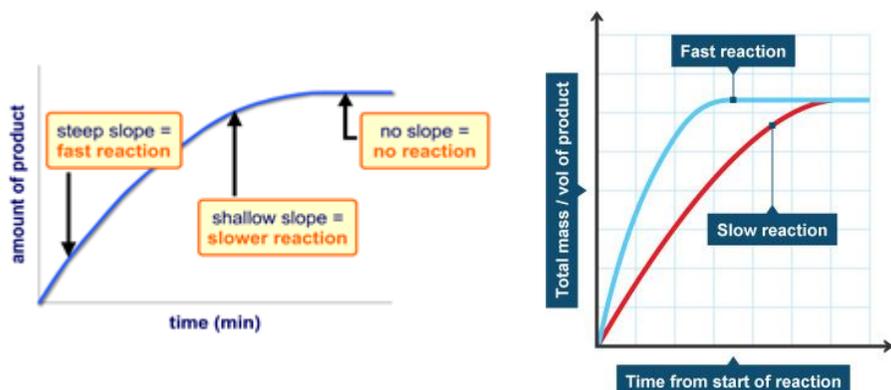
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C8 - Rates and equilibrium

Interpreting Rate of Reaction Graphs

The results from rate of reaction experiments can be plotted on a line graph. For example how the mass changes against time or how much gas is made against time. Different lines can be plotted for different conditions, the **steeper the gradient, the faster the reaction**.

It is important to remember that the graphs flatten off (plateau) at the same point as the same amount of reactant is being used.



Collision Theory

Collision Theory: reactions occur when particles **collide** with a certain amount of **energy**.

The minimum amount of energy needed for the particles to react is called the **activation energy**, which is different for each reaction.

The rate of a reaction depends on two things:

- the **frequency** of collisions between particles. The more often particles collide, the more likely they are to react.
- the **energy** with which particles collide. If particles collide with less energy than the activation energy, they will not react.

Key Terms	Definitions
Activation Energy	The minimum energy required for a chemical reaction to take place
Collision Theory	The theory that states for a chemical reaction to happen, particles must collide with sufficient energy
Gradient	The measurement of how steep a line is on a graph
Frequency	The amount of times something happens in one second
Concentration	The number of particles in a given volume

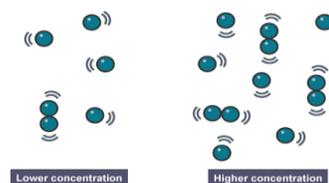
Factors which affect Rate of Reaction

Being able to slow down and speed up chemical reactions is important in everyday life and in industry. We can change the rate of a reaction by:

- Changing temperature
- Changing pressure
- Changing the concentration of a solution
- Changing the surface area
- Adding a catalyst

Collision Theory- in more detail Concentration

If the concentration of a solution is increased then there are more particles in a given volume, therefore collisions are **more frequent** and the chemical reaction is faster. Concentration is **directly proportional** to rate of reaction (if you double the concentration you double the rate



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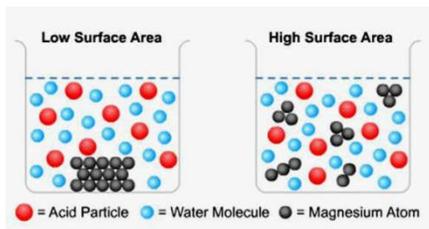
C8 - Rates and equilibrium

Collision Theory in more detail Temperature

When you increase the temperature of something the particles will move around faster, this increases the **frequency of the collisions**. As well as that, as the particles are moving faster the particles collide with more energy making it more likely that collisions exceed the **activation energy**.

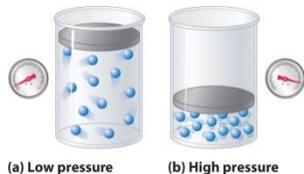
Collision Theory in more detail Surface Area

When you increase the surface area of a solid (you cannot increase the surface area of a liquid or gas). You increase the number of particles that are available for collision, therefore increasing the frequency of collisions therefore increase the rate of reaction.



Collision Theory- in more detail gas pressure

If the reaction is carried out in the gaseous state, then increasing the pressure will increase the rate of reaction. If there are more particles in a given volume of gas, then collisions will be more frequent and therefore the reaction will be faster.



Enzymes

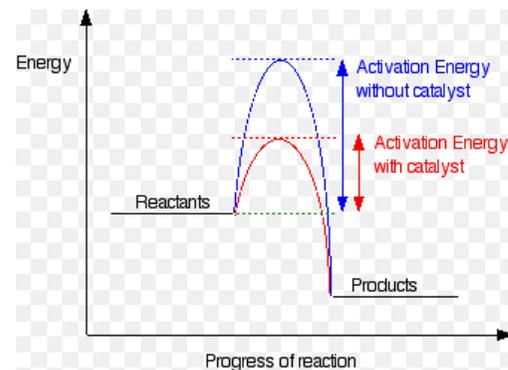
Enzymes **are biological catalysts**, they speed up chemical reactions in biological systems for example in digestion in animals. Unlike catalysts enzymes have an optimum temperature where they work best, this is usually around 37

Key Terms	Definitions
Enzymes	A biological catalyst
Reaction Profile	A graph which show the energies of the reactants and products at different stages of the chemical reaction

Collision Theory in more detail Catalysts

A catalyst is a substance which speeds up a chemical reaction without being used up. It speeds up a reaction because it lowers the activation energy by providing an alternative pathway and this means that there are more **successful collisions and a faster reaction**.

The effect of a catalyst is shown on the reaction profile below:



Catalysts are not included in a chemical equation as they are not used up in a chemical reaction.

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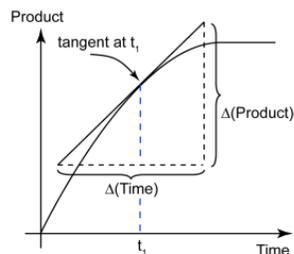
C8 - Rates and equilibrium

Rate of Reaction

The rate of reaction is the speed at which a chemical reaction is happening. This can vary hugely from reaction to reaction.

The rate of reaction can be calculated either by measuring the quantity of **reactant used** or the quantity of **product made in a certain length of time**. The quantity can either be a volume measured in cm^3 or a mass measure in grams (g).

Measuring Rate of Reaction-Higher Tier

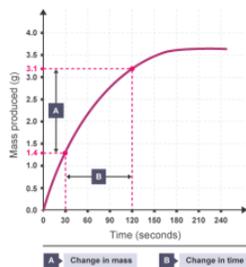


The gradient of a volume or mass/time graph will give you the rate of reaction at a given point. However when the line is a curve you need to draw a **tangent** to measure the gradient. To draw a tangent follow the following steps

1. Line your ruler up across your graph, so that it touches the line on the point that you want to find out the gradient
2. Adjust the ruler until the space between the ruler and the curve is equal on both sides
3. Draw the line and pick two easy points that will allow you to calculate the gradient of the line.

Calculating the Mean Rate of Reaction -Higher Tier

To calculate the mean rate of reaction from a graph you need to pick two y values on the graph and two x values, subtract the largest from the smallest and then divide the value on the y axis by the value on the x axis.



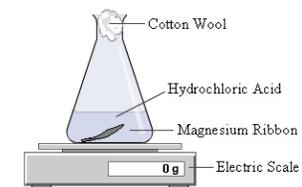
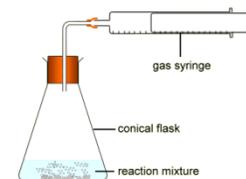
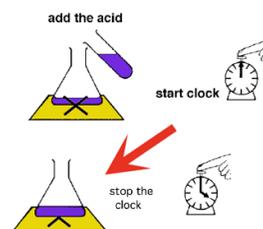
Key Terms	Definitions
Rate of Reaction	The rate at which reactants are being turned into products
Reactant	What is used in a chemical reaction
Product	What is made in a chemical reaction
Catalyst	A substance which speeds up a chemical reaction without being used up
Tangent	A straight line that touches a curve at a point

Equation	Meanings of terms in equation
Rate of Reaction = $\frac{\text{Reactant used}}{\text{time}}$	Reactant used can either be measured in grams or cm^3
Rate of Reaction = $\frac{\text{Product Made}}{\text{time}}$	Product made can either be measured in grams or cm^3

Measuring the Rate of Reaction

There are several experiments that can be used to measure the rate of a chemical reaction.

1. Measuring the mass lost in a chemical reaction (marble chips and acid is a good example)
2. Measuring the volume of gas produced (decomposition of hydrogen peroxide is a good example)
3. Time taken to make an X disappear (sodium thiosulphate and acid is a good example)



Chemistry Knowledge Organiser

C9 - Crude oil and fuels

Neutralisation Reaction

When a salt is made in a neutralisation reaction, it will either be **soluble** or **insoluble**. For example, sulphuric acid can be neutralised with copper oxide to make copper sulphate and water. The **copper sulphate is soluble in water**.

The steps outlined below can be used to make copper sulphate:

1. Add several spatulas of copper oxide to sulphuric acid in a **conical flask**
2. Stir until all the sulphuric acid has reacted
3. Filter off any excess copper oxide
4. Place solution in evaporating basin
5. Allow water to evaporate and blue crystals of copper oxide should be left

Crude Oil

Crude oils is a mixture of chemicals called hydrocarbons. These are chemicals that contain **hydrogen and carbon only**. It made from **ancient biomass, mainly plankton**. Crude oil straight out of the ground is not much use, as there are too many substances in it, all with **different boiling points**.

Before we can use crude oil we have to separate it into its different substances. We do this by fractional distillation.

How does fractional distillation work?

- Crude oil is heated and vaporises/boils.
- Vapours rise up the column, gradually cooling and condensing.
- Hydrocarbons with different size molecules condense at different levels/temperatures
- The crude oil is separated into a series of fractions with similar numbers of carbon atoms and boiling points. These are called fractions.

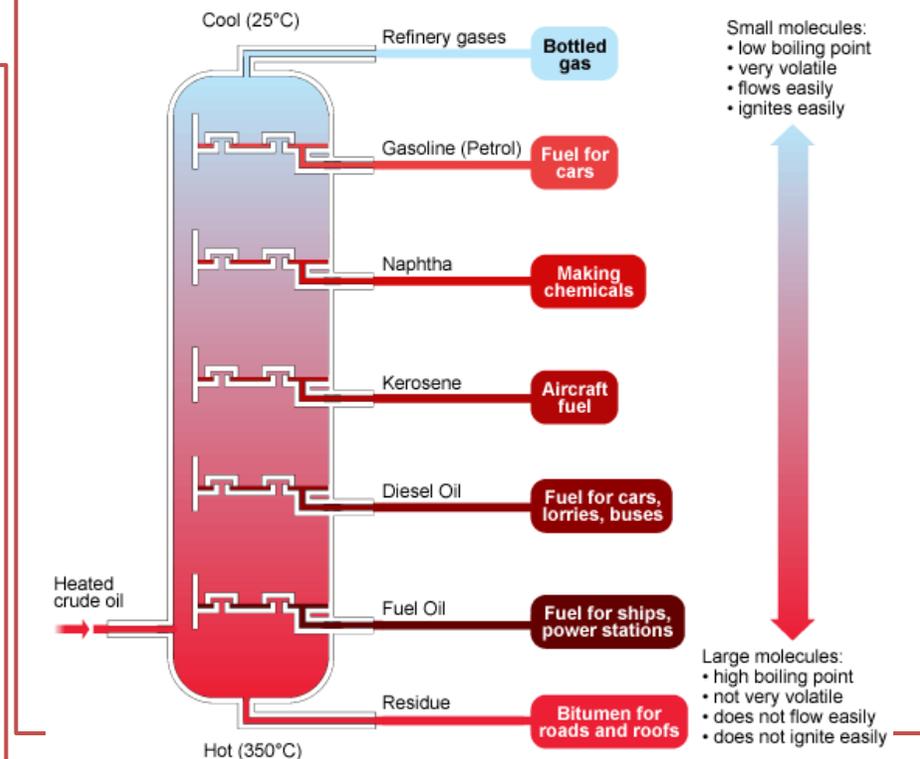
As the number of carbon atoms increases:

- Molecules become larger and heavier
- Boiling point increases
- Flammability decreases (catches fire less easily)
- Viscosity increases (liquid becomes thicker)

Key Terms	Definitions
Hydrocarbon	A compound which contains only hydrogen and carbon (covalently bonded)
Fractional Distillation	The process where crude oil is separated into different compounds through evaporation
Viscosity	The ability of a liquid to flow

Fractional Distillation Column

Below is a diagram of a fractionating column; you need to know the uses but not the names of each fraction:



Chemistry Knowledge Organiser

C9 - Crude oil and fuels

Alkanes

Crude oil is largely made up of a family of hydrocarbons called alkanes; these contain only a single (covalent) carbon to carbon bond.

You can either represent alkanes with a **molecular formula**, e.g.:



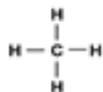
Methane

Ethane

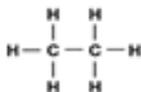
Propane

Butane

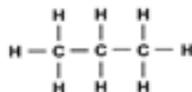
Or a **displayed formula**:



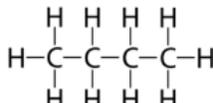
Methane



Ethane



Propane



Butane

[H = Hydrogen, C = Carbon, - indicates a chemical bond between atoms]

Cracking

Smaller hydrocarbons make better fuels as they are easier to ignite. However, crude oil contains a lot of longer chain hydrocarbons. To break a longer chain hydrocarbon down into a smaller one we use a process known as **cracking**.

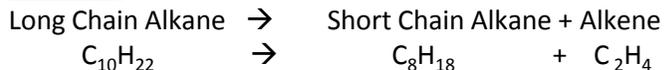
Cracking

So large/long alkanes get **CRACKED**, which means they get broken in two.

- They are **heated**, turned into a vapour and passed over a hot catalyst
- Cracking produces two molecules:

1. One shorter (useful as a fuel) alkane
2. One alkene (used to make polymers).

Summary



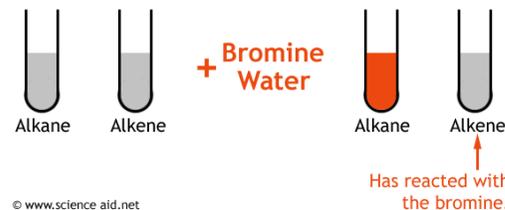
Key Terms	Definitions
Alkane	A hydrocarbon that contains only carbon to carbon single bonds
Cracking	A process where longer chain hydrocarbons are broken down into smaller more useful ones.
Alkene	A hydrocarbon that contains at least one carbon to carbon double bond.

Alkenes

These hydrocarbons have at least one double bonds between the carbon atom. The general formula for alkenes is C_nH_{2n}

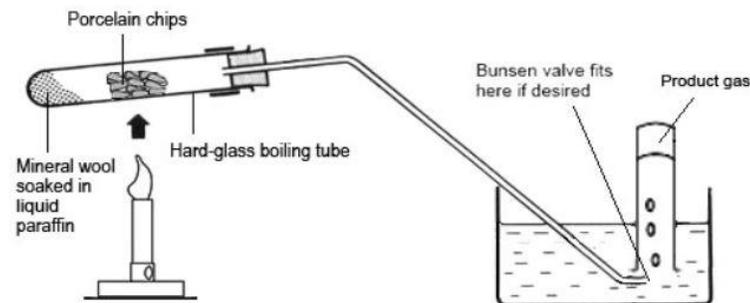
Alkenes are **more reactive** than alkanes. They react with bromine water and make it go from orange to colourless.

Alkanes do not have a double bond so the bromine water stays orange.



Cracking

Experimental set up for cracking:



Chemistry Knowledge Organiser

C9 - Crude oil and fuels

Alkenes

A second family of hydrocarbons is alkenes; these contain at least one double (covalent) carbon to carbon bond. The general formula for alkenes is C_nH_{2n} . Alkenes are **unsaturated** as there is room for 2 more hydrogens around some of the carbons. You need to know the names and structures of the first 4 alkenes. You can either represent alkenes with a **molecular formula**, e.g.:



Ethene

Propene

Butene

Propene

Or a **displayed (structural) formula**:

Name	Molecular formula	Full structural formula
Ethene	C_2H_4	<pre> H H C = C H H</pre>
Propene	C_3H_6	<pre> H H H H - C - C = C H H</pre>
Butene	C_4H_8	<pre> H H H H H - C - C - C = C H H H</pre>
Pentene	C_5H_{10}	<pre> H H H H H H - C - C - C - C = C H H H H</pre>

Key Terms

Definitions

Alkene

A hydrocarbon that contains at least one carbon to carbon double bond.

Unsaturated

A compound that contains at least one carbon to carbon double bond. An alkene is an example of something that is unsaturated.

Addition Reaction

A chemical reaction where an element or compound is added across a double bond.

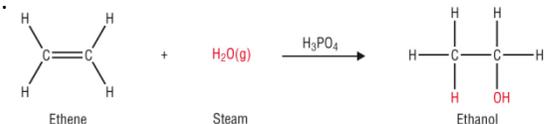
Alkenes

Alkenes undergo **addition reactions**, this is where another element or compound is added across the double bond.

Below is an example of bromine being added across a double bond:



Bromine could be replaced in this equation with another halogen, hydrogen or water. The same type of reaction would take place, however the products formed would be different. For example, the reaction of ethene with water.



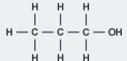
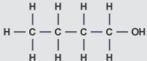
Reagent	Conditions	Product
Hydrogen	Nickel catalyst, 60°C.	Alkane
Water	Steam, high temperature, high pressure. Phosphoric acid catalyst	Alcohol
Halogen	Halogens in solution for example bromine water	Haloalkane

Chemistry Knowledge Organiser

C10 - Organic reactions

Alcohols

Another family of chemicals are alcohols. They are similar in structure to alkanes except that one of the C-H bonds is replaced with a C-OH. To name an alcohol, you use the same system as naming an alkane except the **-ane** section is replaced by **-ol** for example methanol.

Name	Molecular formula	Full structural formula
Methanol	CH ₃ OH	
Ethanol	C ₂ H ₅ OH	
Propan-1-ol	C ₃ H ₇ OH	
Butan-1-ol	C ₄ H ₉ OH	

Making ethanol

There are two techniques for making ethanol:

1. Fermentation
2. Hydration of ethane (see page on alkenes)

Process	Description	Advantages	Disadvantages
Fermentation	Sugar and yeast (anaerobic respiration) produces ethanol	Uses plants so is renewable, carbon neutral. Needs low temp 37°C	Slow, relies on crops
Hydration of ethene	Ethene reacted with steam at a high temp and pressure	Fast, produces ethanol in large quantities	Ethene from crude oil is non renewable. Lots of energy required

Key Terms	Definitions
Alcohol	A family of chemicals containing a C-OH functional group
Carboxylic acid	A family of chemicals containing a  functional group.
Fermentation	When bacteria or yeast is used to break down a chemical.

Reactions and uses of Alcohols

The first four alcohols are flammable and this means that they will undergo **complete combustion** in air making carbon dioxide and water. This means they make very good fuels.

Unlike carboxylic acids alcohols **will not form** acids when dissolved in water.

Alcohols will react with metal producing **hydrogen gas**.

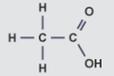
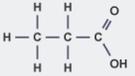
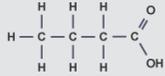
Alcohols are also **excellent solvents**, which means they are useful in the chemical industry.

Alcohols can be **oxidised** to carboxylic acids

They can react with carboxylic acids to form **esters**.

Carboxylic Acids

Another family of chemicals are carboxylic acids. They contain a COOH functional group. To name carboxylic acids you **add -oic acid** to the end of the name. For example ethanoic acid. Carboxylic acids can be made from the oxidation of an alcohol.

Name	Molecular formula	Full structural formula
Methanoic acid	HCOOH	
Ethanoic acid	CH ₃ COOH	
Propanoic acid	C ₂ H ₃ COOH	
Butanoic acid	C ₃ H ₇ COOH	

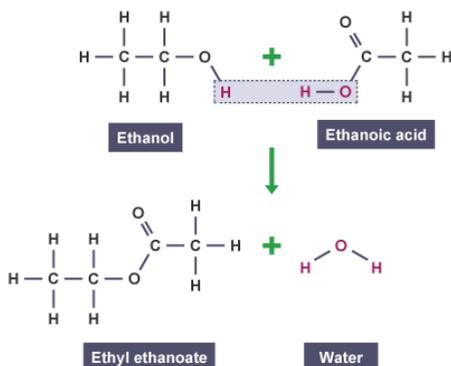
Chemistry Knowledge Organiser

C10 - Organic reactions

Reactions and properties of carboxylic acids

When dissolved in water carboxylic acids will **form a weak acid**. This means they will partially dissociate. This means that they will undergo similar reactions to other acid, for example, they will react with a metal to form hydrogen gas, they will also react with a **metal carbonate to form carbon dioxide**.

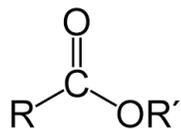
An alcohol and a carboxylic acid will react together to form **an ester**. This reaction needs to be done with a **strong acid catalyst** present, for example concentrated sulphuric acid.



This is known as a **condensation** reaction as two molecules have reacted to make a larger molecule and a small molecule. This small molecule is usually water but can be other small molecules.

Esters

Esters are chemicals with the following functional group:



Esters have a pleasant smell this means they are used in many artificial scents and flavours.

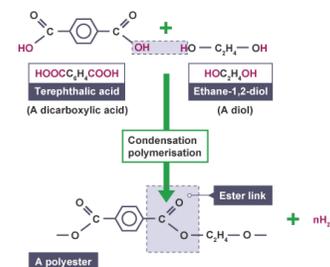
Key Terms	Definitions
Esters	A family of chemicals containing the  functional group
Condensation Reaction	A reaction when two molecules make a larger molecule and a smaller molecule, usually water.
Diol	An alcohol which contains 2 C-OH groups
Condensation Polymer	A polymer that has been formed through a condensation reaction

Condensation Polymers

There is a second type of polymer which is known as a condensation polymer. These are formed from a condensation reaction. For condensation polymerisation to occur, you need to have reagents that have the correct functional groups at **both ends of the molecule**. For example to make a polyester, you will need, a diol and a diacid.



The R in the diagram above just represents the rest of the molecule, to make different sorts of polyester, you simply change the R group. Below is an example showing the formation of one type of polyester.

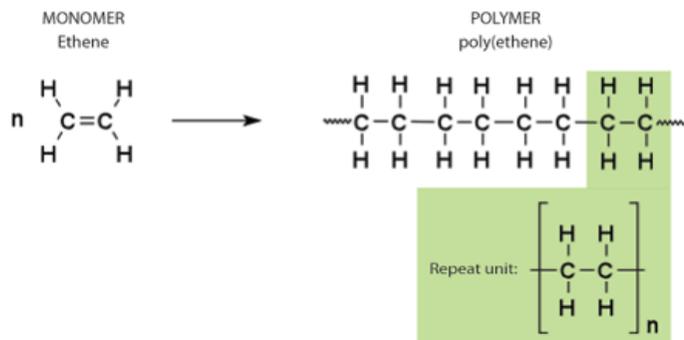


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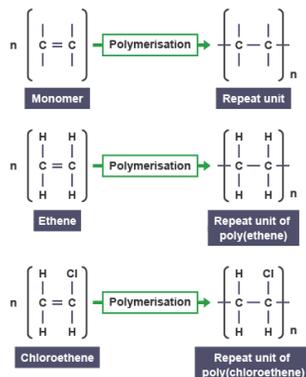
C11 - Polymers

Addition Polymers

Plastics are materials made from **polymers**, which are long chain molecules containing covalent bonds. Polymers are made by joining **monomers** together. Monomers have to have a carbon to carbon double bond. This happens when one of the bonds in a double bond is broken and the monomer joins to the next one, making a long chain. The name of a polymer comes from putting poly- in front of the name of the monomer. This type of polymerisation is known as **addition polymerisation**. Different monomers will give polymers with different properties.



To represent polymers we use a **repeating unit**. As the polymers are typically many thousands of carbon atoms long we use an n to represent a large number.



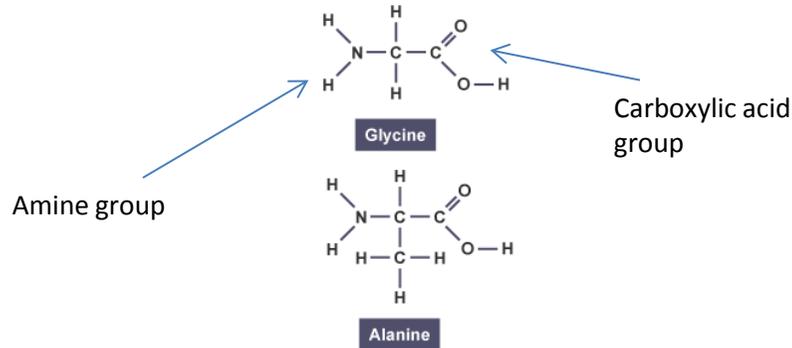
Key Terms	Definitions
Addition Polymer	An addition polymer is a polymer which is formed by an addition reaction , where many monomers bond together with no loss of atoms or molecules.
Monomer	A molecule that can be bonded to other molecules to form a polymer. An alkene is an example.
Repeating Unit	A repeating unit is a part of a polymer whose repetition would produce the complete polymer chain.

Chemistry Knowledge Organiser

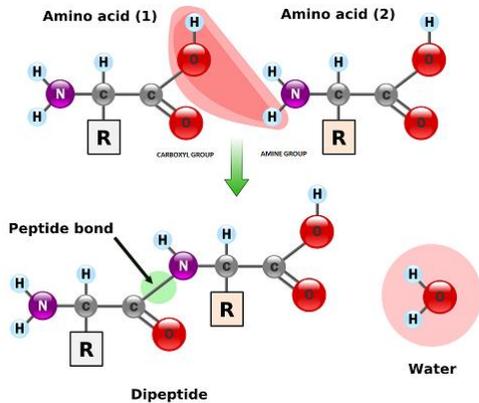
C11 - Polymers

Condensation Polymers

Other examples of condensation polymers include proteins. These are polymers of amino acids. Amino acids contain two functional groups, **a carboxyl group and an amino group**. Below is the diagram of two amino acids, alanine and glycine:



The carboxyl acid can react with the amino group on another amino acid molecule. This will form a peptide bond and as this reaction continues a large **polypeptide or protein** will form.

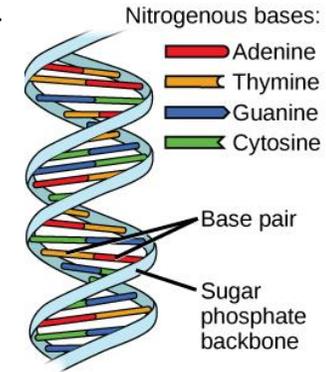


Proteins form molecules like enzymes, haemoglobin and a wide variety of body tissues.

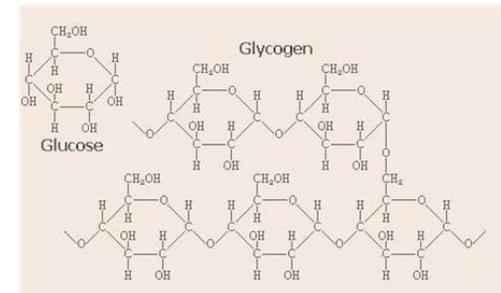
Key Terms	Definitions
Amino Acid	A molecule which contains a carboxyl and amino group.
Peptide bond	A bond formed between an amino and carboxyl group.

DNA

DNA is made up of a nucleotide strands with bases to form the double helix structure. The nucleotide strands (sugar phosphate backbones) are condensation polymers.



Sugars can also form condensation polymers. For example glucose can be stored as **glycogen**, a **polymer**. Plants also make a polymer called **cellulose** to make their cell walls.



Chemistry Knowledge Organiser

C13 - The Earth's atmosphere

The Atmosphere

For 200 million years, the amount of different gases in the atmosphere have been much the same as they are today:

- 78% nitrogen
- 21% oxygen
- The atmosphere also contains small proportions of various other gases, including carbon dioxide, water vapour and noble gases.

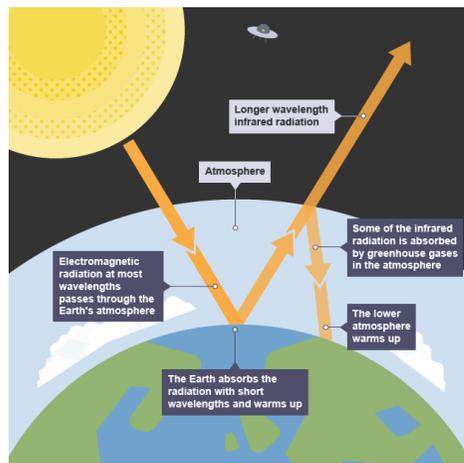
The Greenhouse Effect

The Earth has a layer of gases called the **Greenhouse layer**. These gases, which include carbon dioxide, methane and water vapour, maintain the temperature on Earth high enough to support life.

The greenhouse layer allows the short wave infrared radiation emitted by the Sun to pass through it but absorbs the long wave infra red radiation which is emitted by the Earth. This is how it insulates the Earth.

Some human activities increase the amounts of greenhouse gases in the atmosphere. These include:

- combustion of fossil fuels
- deforestation
- methane release from farming
- more animal farming (digestion, waste decomposition)



Key Terms

Definitions

Greenhouse Layer

The layer of gases which absorb infra red radiation emitted from the Earth

The Evolution of the Atmosphere

Scientists are not sure about the gases in the early atmosphere, as it was so long ago (4.6 billion years) and the lack of evidence. Many scientists believe the early atmosphere was made up of mainly carbon dioxide, water vapour and small amounts of methane, ammonia and nitrogen, released by **volcanoes**. **There was little or no oxygen around at this time**. The early Earth was very hot, but as it cooled the water vapour in the atmosphere condensed and **formed the oceans**.

As the oceans formed, carbon dioxide dissolved in the ocean. The carbon dioxide formed carbonates and precipitated out (formed solids). This process reduced the amount of carbon dioxide in the atmosphere.

Approximately 2.7 billion years ago, plants and algae evolved. This decreased the amount of carbon dioxide in the atmosphere and increased the amount of oxygen in the atmosphere.

When sea animals evolved they used the carbon dioxide in the ocean to form their shells and bones (which are made of carbonates). When these sea creatures died their shells and bones became limestone (calcium carbonate), which is a sedimentary rock.

Once enough oxygen was in the atmosphere, it could support animals, which carry out respiration. These processes have caused the levels of gases in the atmosphere to be where they are today.

Changes in the atmosphere

Recent activity by humans has changed the composition of the atmosphere. Combustion of fossil fuels has increased the amount of carbon dioxide in the atmosphere as well as other harmful gases such as nitrous oxides, which are made by nitrogen reacting with oxygen in the air.

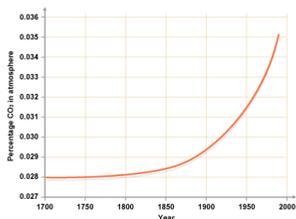
Sulphur is also present in many fuels, this has increased the amount of sulphur dioxide which causes acid rain. Carbon particles can also be released as carbon monoxide from incomplete combustion.

Chemistry Knowledge Organiser

C13 - The Earth's atmosphere

The Enhanced Greenhouse Effect

In the last 100 years humans have added to the greenhouse layer through combustion of fossil fuels, increased farming and deforestation. Many scientists believe this has led to a **rise in global temperature**.



However, this is such a complex system that misunderstandings of it can lead to **inaccurate or biased** opinions being reported in the media.

Consequences of Climate Change

An increase in average global temperature is a major cause of **climate change**.

The potential effects of global climate change include:

- sea level rise, which may cause flooding and increased coastal erosion
- more frequent and severe storms
- changes in the amount, timing and distribution of rainfall
- water shortages for humans and wildlife
- changes in the food producing capacity of some regions
- changes to the distribution of wildlife species.

Students should be able to discuss the scale, risk and environmental implications of global climate change.

Waste water and Sewage

Water from houses and farming needs to **be treated** before it can be released into rivers and lakes. It is firstly **filtered** to remove large particles and is then left so that the sediment drops to the bottom. The "sludge," this is the name given to the sediment at the bottom, is then anaerobically digested (broken down by bacteria) to make methane gas. Any remaining **effluent** is broken down by aerobic respiration. The water is then released back into the rivers and lakes.

Key Terms	Definitions
Carbon Footprint	The carbon footprint is the total amount of carbon dioxide and other greenhouse gases released over the life of a product
Carbon Neutral	There is no net increase in carbon dioxide in the atmosphere

Carbon Footprint

The **carbon footprint** is the total amount of carbon dioxide and other greenhouse gases released over the life of a product. Many people or businesses look to reduce their carbon footprint by:

- increased use of alternative energy supplies
- energy conservation
- carbon capture and storage
- carbon taxes and licences

People also try to **offset** their carbon by planting trees.

If something is carbon neutral, this means that there is no net increase in **carbon dioxide in the atmosphere** when it is used.

Water

Water of appropriate quality is **essential for life**. For humans, drinking water should have low levels of dissolved **salts and microbes**. Water that is safe to drink is called **potable water**.

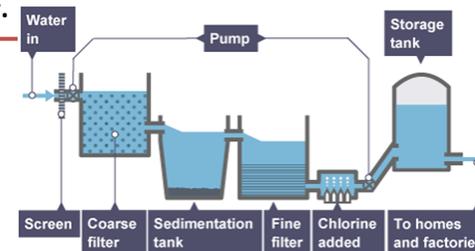
The methods used to produce potable water depend on available supplies of water and local conditions.

In the United Kingdom (UK), rain provides water with low levels of dissolved substances (fresh water) that collects in the ground and in lakes and rivers, and most potable water is produced by:

- passing the water through filter beds to remove any solids
- sterilising to kill microbes, using chlorine or UV light

In some parts of the world there is not enough fresh water so the salt has to be removed from water. This process is called **desalination**.

Desalination can be done by distillation or reverse osmosis. This requires a **large amount of energy**.

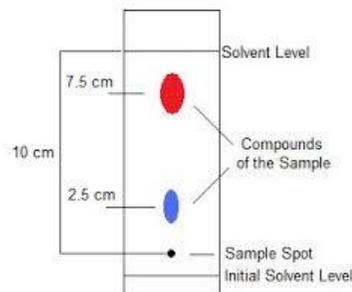


Chemistry Knowledge Organiser

C12 - Chemical Analysis

Chromatography and Rf values

- When carrying out chromatography we can calculate an Rf (retention factor) value/
- The retention factor is a ratio between the distance travelled by the solvent and the distance travelled by a compound.
- Chromatography has two phases- a stationary phase where particles can't move (the filter paper in most cases), a mobile phase where particles can move (a solvent for example water).
- Different compounds will have different Rf values in different solvents, this allow us to see whether a substance is pure or impure.
- To calculate Rf value you need to divide the distance moved by the solvent by the distance moved by the spot.
- For example to work out the Rf for the spot further up the paper:
- $Rf = \frac{B}{A}$ $Rf = \frac{7.5}{10} = 0.75$
- There are no units as the answer is a ratio
- The higher the Rf the further the spot has moved up the paper, compared to the solvent.



Transition Metals

- The central block (between group 2 and 3) of the Periodic Table is known as the transition metals.
- Compared to group 1 elements, transition metals have different physical properties. For example transition metals have a higher melting point and are more dense.
- The exception is mercury which is a liquid at room temperature.
- Transition metals also have different physical properties to group 1. They are much less reactive and do not react vigorously with oxygen or water.

Key Terms	Definitions
Retention Factor	The ratio between the distance travelled by the substance and the distance travelled by the solvent.

Equation	Meanings of terms in equation and units
$Rf = \frac{B}{A}$	<i>Rf = Retention Factor (no units)</i> <i>B = Distance travelled by substance (cm)</i> <i>A = Distance travelled by solvent (cm)</i>

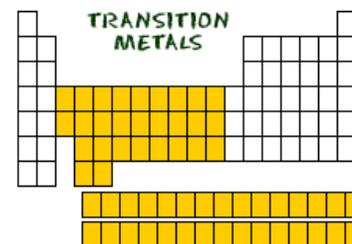
Melting Point and Boiling point

- A chemically pure substance will melt or boil at a very specific temperature.
- If a substance is chemically impure it will melt or boil at a lower temperature and across a broader range.
- The closer the substance is to the melting point the purer the substance.

Formulations

- Formulations are mixtures made using a precise amount of each substance, so they can serve a particular purpose.
- For example in paints or in pills.

Transition Metals



Chemistry Knowledge Organiser

C12 - Chemical analysis

Testing for positive ions

Positive ions (metal ions) can be identified by flame tests:

Metal and ion	Colour of flame test
Sodium Na ⁺	Yellow
Lithium Li ⁺	Crimson
Potassium K ⁺	Purple
Copper Cu ²⁺	Green
Calcium Ca ²⁺	Red/Orange

To carry out a flame test you need to do the following:

1. Dip platinum loop in dilute HCl, hold in Bunsen burner flame (blue flame), until no colour is seen.
2. Dip the loop into the sample you are testing
3. Place this into the flame and observe the colour

Testing for negative ions

Ion	Test	Equation
Carbonate (CO ₃ ²⁻)	Add metal carbonate to dilute acid in a boiling tube. Connect the boiling tube to a test tube containing limewater. If the limewater turns cloudy then a carbonate ion is present	$K_2CO_3 + 2HCl \rightarrow 2KCl + CO_2 + H_2O$
Sulphate (SO ₄ ²⁻)	Add 5 drops of dilute HCl, followed by 5 drops of barium chloride. If sulphate ions are present then a white precipitate will be formed.	$Ba^{2+} + SO_4^{2-} \rightarrow BaSO_4$ This is the ionic equation for the reaction.
Halides (Cl ⁻ , Br ⁻ , I ⁻)	Add 5 drops of dilute nitric acid and 5 drops of silver nitrate, the colour of the silver halide precipitate formed will vary depend on the halogen Cl ⁻ – White Br ⁻ – Cream I ⁻ – Yellow	$Ag^+ + Cl^- \rightarrow AgCl$ This is the ionic equation for the reaction.

More tests for metal ions

Some metal hydroxides are insoluble. Therefore if some drops of sodium hydroxide are added to a solution of the metal hydroxide a precipitate may form. Transition metal hydroxides are usually coloured. Where as main group elements normally form a white precipitate.

Gas	Colour of precipitate	Ionic Equation
Magnesium Mg ²⁺	White	$Mg^{2+} + 2OH^- \rightarrow Mg(OH)_2$
Calcium Ca ²⁺	White	$Ca^{2+} + 2OH^- \rightarrow Ca(OH)_2$
Iron(II) Fe ²⁺	Green	$Fe^{2+} + 2OH^- \rightarrow Fe(OH)_2$
Iron(III) Fe ³⁺	Brown	$Fe^{3+} + 3OH^- \rightarrow Fe(OH)_3$
Copper Cu ²⁺	Blue	$Cu^{2+} + 2OH^- \rightarrow Cu(OH)_2$
Aluminium Al ³⁺	White initially. In excess NaOH it dissolves to form a colourless solution.	$Al^{3+} + 3OH^- \rightarrow Al(OH)_3$

Chemistry Knowledge Organiser

C14 - The Earth's resources

LCA's

Life cycle assessments (LCAs) are carried out to assess the environmental impact of products in each of these stages of a products life:

1. extracting and processing raw materials
2. manufacturing and packaging
3. use and operation during its lifetime
4. disposal at the end of its useful life, including transport and distribution at each stage.

Some things for example the energy required to make the product are easy to measure. However some things like how much pollution it releases are hard to measure and therefore difficult to give a value to.

Example of an LCA

	Plastic Bag	Paper Bag
Raw Material	Crude Oil	Timber
Manufacturing and Packaging	Made from crude oil by fractional distillation, then cracking and polymerisation, high energy process. Little waste as other fractions are used for other things	Made by pulping timber. Lots of waste, high energy process
Use of product	Has multiple uses, can be reused.	Usually only used once.
Disposal	Can be recycled but are not biodegradable	Can be recycled and are biodegradable

Key Terms

Definitions

LCA

An evaluation of the environmental impact a product had over its lifetime

Recycling

Many of the Earth's resources are finite: for example, metals and crude oil. It is therefore vital we recycle resources. The processes for extracting these materials are often high energy and damaging to the environment.

Metals can be recycled by melting and **recasting or reforming** into different products.

Some products, such as glass bottles, can be reused. Glass bottles can be **crushed and melted** to make different glass products. Other products cannot be reused and so are recycled for a different use.

Required Practical Chemistry – Making salts

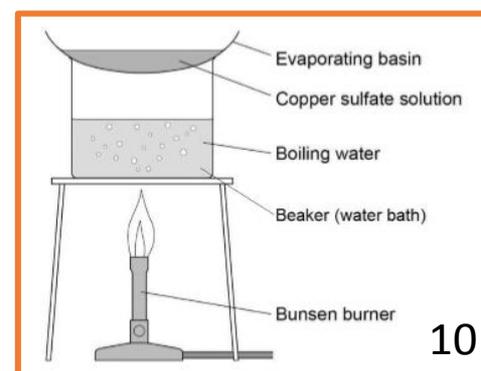
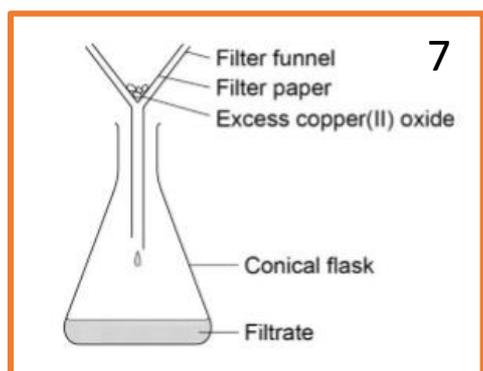
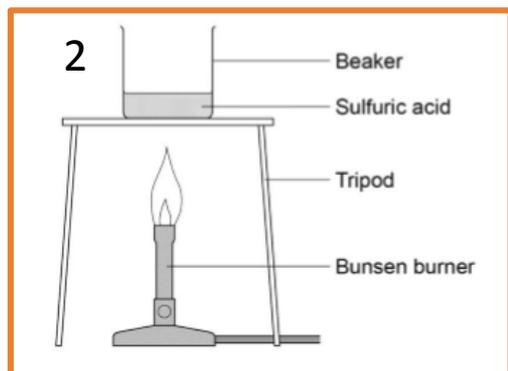
Objective: Making soluble salts: preparation of pure dry copper sulfate crystals.

In this practical you will:

- react sulfuric acid with insoluble copper (II) oxide to prepare an aqueous solution of the salt copper sulfate.
- separate out unreacted copper (II) oxide by filtration.
- prepare pure dry crystals of copper sulfate from the solution.

Method

1. Measure 40cm³ sulfuric acid and put it into the 100cm³ 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 (II) oxide powder to the hot acid. Stir with the glass rod. The copper (II) oxide will disappear and the solution will turn clear blue.
4. Add some more copper (II) oxide and stir again.
5. Keep adding the copper (II) 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 250cm³ 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.
14. Remove the crystals from the concentrated solution with a spatula. Gently pat the crystals dry between two pieces of filter paper.



Required Practical Chemistry – Temperature changes

Objective: Investigate the variables that affect temperature change in chemical reactions eg acid plus alkali.

In this practical you will:

- react sodium hydroxide solution with hydrochloric acid.
- measure the temperature changes during the reaction.
- plot a graph of your results and record the temperature change.

Method

1. Measure 30cm³ 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 like this.
4. Measure 5cm³ 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 5cm³ amounts of sodium hydroxide to the cup each time, recording your temperature reading in the results table.
9. Repeat until a maximum of 40cm³ 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			
5			
10			

Apparatus

- dilute hydrochloric acid
- dilute sodium hydroxide solution
- an expanded polystyrene cup and lid
- 250 cm³ beaker
- 10 cm³ measuring cylinder
- 50 cm³ measuring cylinder
- a thermometer

Required Practical
Chemistry – Temperature changes

Objective: Investigate the variables that affect temperature change in chemical reactions eg acid plus alkali.

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	

Analysis and conclusion

- Calculate the mean maximum temperature reached for each volume of sodium hydroxide. Record these means in your table.
- Plot a graph from your results and draw two straight lines of best fit.
- From the graph read off the maximum temperature change.
- This is an example of an exothermic reaction, when heat is given out. Can you explain why the results you recorded show that the temperature started to fall after a certain volume of sodium hydroxide had been added?
- Can you explain what is happening to the chemical bonds in the reactants and the products when an exothermic reaction is taking place?

Health and safety

- Eye protection
- sulfuric acid: IRRITANT
- hydrochloric acid: IRRITANT

Required Practical
Chemistry – Concentration and rate of reaction.
Activity 1

Objective: Investigate how changes in concentration affect the rates of reactions by measuring the volume of a gas produced.

In this practical you will:

- react magnesium ribbon with different concentrations of hydrochloric acid.
- measure the volume of gas produced for each concentration.
- use your results to work out how the rate of reaction is affected by the concentration of the acid.

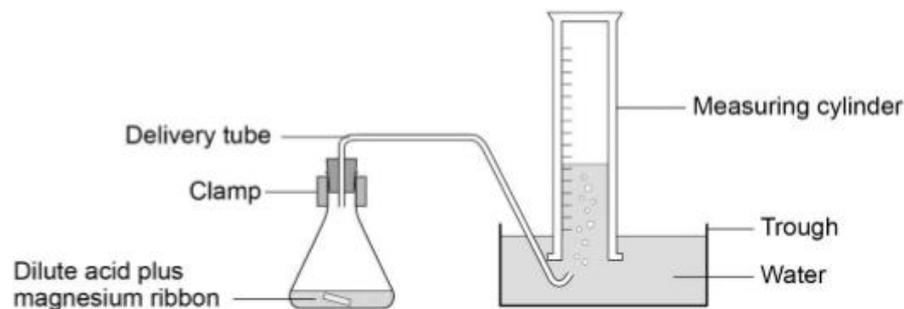
Method

1. Measure 50cm³ of 1.0m hydrochloric acid using one of the measuring cylinders. Pour the acid into the 100cm³ conical flask.
2. Fit the bung and delivery tube to the top of the flask.
3. Half fill the trough or bowl with water.
4. Fill the other measuring cylinder with water. Make sure it stays filled with water when you invert it into the water trough and that the delivery tube is positioned correctly.
5. Add a single 3cm strip of magnesium ribbon to the flask, put the bung back into the flask as quickly as you can, and start the stop clock.
6. Record the volume of hydrogen gas given off at suitable intervals (eg 10 seconds) in a table.
7. Repeat steps 1–6 using 1.5m hydrochloric acid.

Analysis, conclusion and evaluation

- a. Plot a graph of your results.
- b. Draw a line of best fit. The results should generate a curve not a straight line.
- c. Plot the curve for both 1.0M and 1.5M hydrochloric acid on the same graph.
- d. Use your graph to compare the rates of reaction with different concentrations of hydrochloric acid with magnesium.
- e. Use kinetic theory to explain your findings.

Time in seconds	Volume of gas produced cm ³	
	1.0 M	1.5 M
10		
20		
30		
40		
50		
60		
70		
80		
90		
100		



Required Practical
Chemistry – Concentration and rate of reaction.
Activity 1

Objective: Investigate how changes in concentration affect the rates of reactions by measuring the volume of a gas produced.

In this practical you will:

- react magnesium ribbon with different concentrations of hydrochloric acid.
- measure the volume of gas produced for each concentration.
- use your results to work out how the rate of reaction is affected by the concentration of the acid.

Apparatus

- magnesium ribbon cut into 3cm lengths
- dilute hydrochloric acid, 1.0M and 1.5M
- safety goggles
- conical flask (100cm³)
- single-holed rubber bung and delivery tube to fit conical flask
- water trough
- two measuring cylinders (100cm³)
- clamp stand, boss and clamp
- stop clock

Health and safety

- 2.0 mol/dm³ aqueous hydrochloric acid: IRRITANT
- eye protection.

Time in seconds	Volume of gas produced (cm ³)	
	1.0 M	1.5 M
10	4.0	12.0
20	9.0	25.0
30	16.0	36.0
40	23.0	49.0
50	31.0	56.0
60	39.0	56.0
70	44.0	56.0
80	50.0	56.0
90	50.0	56.0
100	50.0	56.0

Required Practical
Chemistry – Concentration and rate of reaction.
Activity 2

Objective: Investigate how changes in concentration affect the rates of reactions by monitoring a change in colour or turbidity.

In this practical you will:

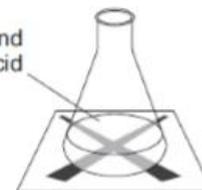
- react different concentrations of sodium thiosulfate with hydrochloric acid.
- use a stop clock to time how long it takes for the mixture to become cloudy for each concentration.
- use your results to work out how the rate of reaction changes as the concentration of the sodium thiosulfate changes.

Method

1. Measure 10cm³ sodium thiosulfate solution and put it into the conical flask.
2. Measure 40cm³ of water. Add the water to the conical flask.
3. This dilutes the sodium thiosulfate solution to a concentration of 8 g/dm³.
4. Put the conical flask on the black cross.
5. Measure 10cm³ of dilute hydrochloric acid.
6. Put this acid into the flask. At the same time swirl the flask gently and start the stop clock.
7. Look down through the top of the flask. Stop the clock when you can no longer see the cross.
8. Record the time it takes for the cross to disappear in the table below. Record the time in seconds.
9. Repeat steps 1–7 changing the concentration of sodium thiosulphate each time as below
 1. 20cm³ sodium thiosulfate + 30cm³ water (concentration 16g/dm³)
 2. 30cm³ sodium thiosulfate + 20cm³ water (concentration 24g/dm³)
 3. 40cm³ sodium thiosulfate + 10cm³ water (concentration 32g/dm³)
 4. 50cm³ sodium thiosulfate + no water (concentration 40g/dm³)



Sodium thiosulfate and dilute hydrochloric acid



Analysis, conclusion and evaluation

- a. Share results with two other groups. Record these results in the second and third blank columns of your table.
- b. Calculate the mean time for each of the sodium thiosulfate concentrations. Leave out anomalous values from your calculations.
- c. Plot a graph and draw a smooth curved line of best fit.
- d. Describe the relationship between the independent variable and the dependent variable? What were your control variables?
- e. Compare your results with those of others in the class. Is there evidence that this investigation is reproducible?
- f. Evaluate the two methods that you have used to investigate the effect of concentration on rate of reaction.

Concentration of sodium thiosulfate in g/dm ³	Time taken for cross to disappear in seconds			
	First trial	Second trial	Third trial	Mean
8				
16				
24				
32				
40				

Required Practical
Chemistry – Concentration and rate of reaction.
Activity 2

Objective: Investigate how changes in concentration affect the rates of reactions by monitoring a change in colour or turbidity.

In this practical you will:

- react different concentrations of sodium thiosulfate with hydrochloric acid.
- use a stop clock to time how long it takes for the mixture to become cloudy for each concentration.
- use your results to work out how the rate of reaction changes as the concentration of the sodium thiosulfate changes.

Apparatus

- 40g/dm³ sodium thiosulfate solution
- 1.0M dilute hydrochloric acid
- a conical flask (100cm³)
- a printed black paper cross
- a stop clock

Health and safety

- 2.0 mol/dm³ aqueous hydrochloric acid: IRRITANT.
- eye protection.
- Take care to avoid breathing in any sulfur dioxide fumes.

Concentration of sodium thiosulfate in g/dm ³	Time taken for cross to disappear in seconds			
	First trial	Second trial	Third trial	Mean
8	137	138	156	
16	76	75	77	
24	46	47	46	
32	37	40	37	
40	27	25	27	

Required Practical Chemistry – Chromatography

Objective: Investigate how paper chromatography can be used to separate and identify a mixture of food colourings.

Method

1. Use a ruler to draw a horizontal pencil line 2cm from the bottom short edge of the chromatography paper. This is your origin line.
2. Mark five pencil spots at equal intervals across the origin line. Make sure you keep at least 0.5cm away from each edge of the paper.
3. Use a glass capillary tube to put a small spot of each colouring A, B, C and D on four of the pencil spots. Use a different tube for each colouring. Use the fifth tube to put a small spot of the unknown mixture U on the fifth pencil spot. Try to make sure each spot is no more than 2-3mm in diameter. Label each spot in pencil.
4. Pour water into the beaker to a depth of no more than 1cm.
5. Clip the top short edge of the chromatography paper to the wooden spill. The top end is the end furthest from the spots.
6. Carefully rest the wooden spill on the top edge of the beaker. The bottom edge of the paper should dip into the water solvent.
7. Wait for the water solvent to travel at least three quarters of the way up the paper. Do not disturb the beaker during this time.
8. Carefully remove the paper from the beaker. Draw another pencil line on the dry part of the paper as close to the wet edge as possible. This is called the solvent front line.
9. Hang the paper up to dry thoroughly.
10. Measure the distance in mm between the two pencil lines. This is the distance travelled by the water solvent.
11. For each of food colour A, B, C and D measure the distance in mm from the start line to the middle of the spot.
12. Record your measurements in the table on the next page.

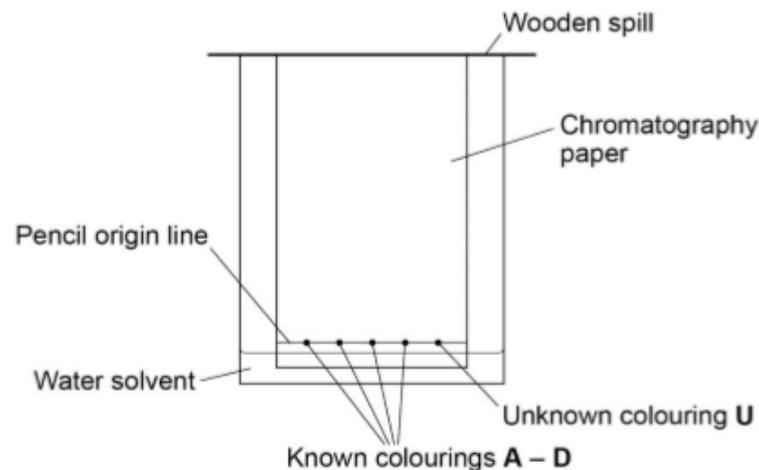
In this practical you will:

- use paper chromatography to separate a mixture of food colourings
- calculate the R_f value for each pure food colouring in the solvent water
- use your R_f values to identify which colourings are in the unknown mixture, U

Make sure that:

- the pencil start line is above the water surface
- the sides of the paper do not touch the beaker wall.

Your apparatus should look like this:



Required Practical Chemistry – Chromatography

Objective: Investigate how paper chromatography can be used to separate and identify a mixture of food colourings.

In this practical you will:

- use paper chromatography to separate a mixture of food colourings.
- calculate the R_f value for each pure food colouring in the solvent water.
- use your R_f values to identify which colourings are in the unknown mixture, U.

Apparatus

- a 250cm³ beaker
- a wooden spill
- a rectangle of chromatography paper
- four known food colourings labelled A–D
- an unknown mixture of food colourings labelled U
- five glass capillary tubes
- a paper clip
- a ruler
- a pencil

Food colouring	Distance travelled in mm		R_f value
	Solvent	Spot	
A			
B			
C			
D			

Analysis, conclusion and evaluation

- Calculate the R_f value for each of the known colours. Use the equation: $R_f = \frac{\text{distance moved by solute}}{\text{distance moved by solvent}}$
- Observe the spots made by food colouring mixture U.
- What conclusions can you draw from your results?
- Are there any other colourings in mixture U which do not match A–D?
- How do you know that a chemical is pure from chromatography results? Are there any other ways that chemists use to ensure that a substance is a pure substance?

Required Practical Chemistry – Water purification

Objective: Analysing and purifying a sample of water and making it safe to drink.

In this practical you will:

- analyse a water sample
- purify a water sample by distillation

Method - Activity 1: Analysing a sample of water

1. Use the universal indicator paper to measure the pH of the water sample.
2. Accurately weigh an empty evaporating basin and record to two decimal places.
3. Pour 10cm^3 of water sample 1 into the evaporating basin.
4. Heat the evaporating basin on a tripod and gauze using a Bunsen burner until the solids start to form and the majority of water has evaporated.
5. Weigh the cooled evaporating basin again and calculate the mass of the solids that were dissolved in the water.
6. Record your results in a table:

	pH	Mass of solids dissolved in 10 cm^3
Water sample 1		
Water sample 2		
Water sample 3		

Apparatus

- 10cm^3 of each of the water samples to be tested
- universal indicator paper or solution.

Task –after completing activity 1 and 2

Use the following key words to write a paragraph to explain how you have collected a sample of pure water:

- evaporation
- condensation
- steam
- gas
- liquid
- boiling
- temperature
- cool surface

What is the difference between pure water and potable water?

Required Practical Chemistry – Water purification

Objective: Analysing and purifying a sample of water and making it safe to drink.

In this practical you will:

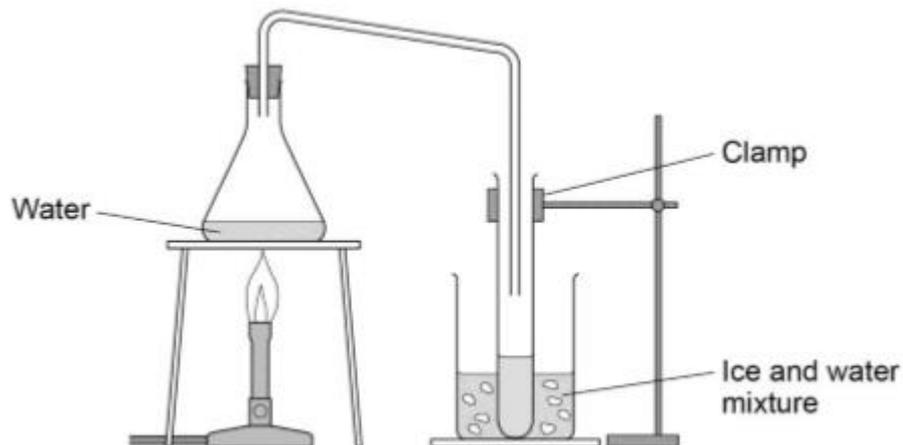
- analyse a water sample
- purify a water sample by distillation

Apparatus

- 10cm³ of water sample 1
- a Bunsen burner
- a tripod
- gauze
- a heatproof mat
- clamp and clamp stand
- conical flask with delivery tube and bung
- a boiling tube
- ice bath
- cobalt chloride paper

Method - Activity 2: Purifying a sample of water by distillation

1. Place the water sample in the conical flask. Set up the apparatus for distillation as shown in the diagram.
2. Heat the water using the Bunsen burner until it boils. Then reduce the heat so that the water boils gently.
3. The distilled water will collect in the cooled test tube. Collect about 1cm depth of water in this way, then stop heating.
4. Analyse the water you have distilled with cobalt chloride paper.



Required Practical Chemistry – Electrolysis

Objective: Investigate what happens when aqueous solutions are electrolysed using inert electrodes.

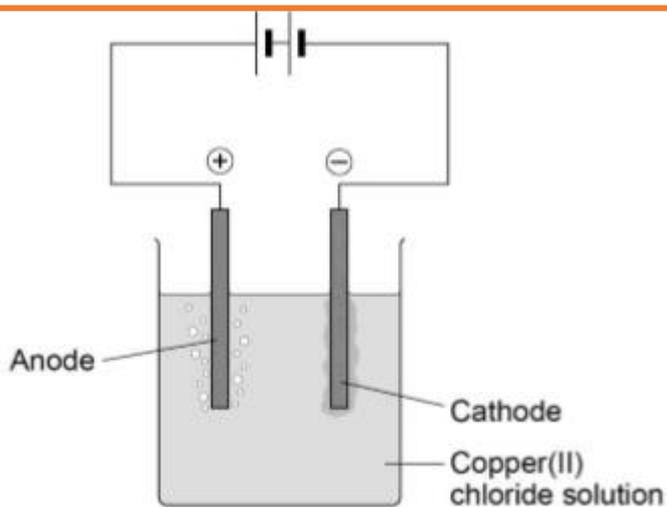
In this practical you will:

- use a low voltage power supply and carbon rod electrodes to pass a current through two different salt solutions.
- identify the element formed at the positive and negative electrodes for each solution.
- add extra detail to the basic electrochemical diagram provided.

Method

1. Pour approximately 50cm³ copper (II) 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 4V on the power supply and switch on.
5. Look at both electrodes and record your initial observations in the table below.
6. Use forceps to hold a piece of blue litmus paper in the solution next to the anode (positive electrode) and identify the element? Write all your observations in a table like the one below.
7. Rinse the electrochemical cell apparatus and collect a new set of electrodes.
8. Repeat steps 1–7 using the other solution sodium chloride and complete the following tasks to show your understanding of the chemistry of electrolysis.

Care - switch off the power supply after 5 minutes.



Solution	Positive electrode (anode)			Negative electrode (cathode)		
	Observations	Element formed	State	Observations	Element formed	State
Copper (II) chloride						
Sodium chloride						

Required Practical Chemistry – Electrolysis

Objective: Investigate what happens when aqueous solutions are electrolysed using inert electrodes.

In this practical you will:

- use a low voltage power supply and carbon rod electrodes to pass a current through two different salt solutions.
- identify the element formed at the positive and negative electrodes for each solution.
- add extra detail to the basic electrochemical diagram provided.

Apparatus

- copper(II) chloride solution
- sodium chloride solution
- 100cm³ beaker
- petri dish lid
- two carbon rod electrodes
- two crocodile/4mm plug leads
- low voltage power supply
- blue litmus paper
- forceps

Health and safety

- Copper(II) chloride solution – IRRITANT
- Oxygen gas – OXIDISING
- Hydrogen gas – EXTREMELY FLAMMABLE
- Chlorine gas – TOXIC
- The electrolysis of brine produces a solution of sodium hydroxide, which is corrosive.
- Wear chemical splash proof eye protection

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