CHEMISTRY REVISION GUIDE
for CIE IGCSE Coordinated Science (2012 Syllabus)

This revision guide is designed to help you study for the chemistry part of the IGCSE Coordinated Science course.

The guide contains everything that the syllabus says you need you need to know, and nothing extra.

The material that is in the supplementary part of the course (which can be ignored by core candidates) is highlighted in dashed boxes:

Whilst this guide is intended to help with your revision, it should not be your only revision. It is intended as a starting point but *only a starting point*. You should make sure that you also read your text books and use the internet to supplement your study in conjunction with your syllabus document.

Whilst this guide does contain the entire syllabus, it just has the bare minimum and is not in itself sufficient for those candidates aiming for the highest grades. If that is you, you should make sure you read around a range of sources to get a deeper knowledge and understanding.

Some very useful websites to help you further your understanding include:

- http://www.docbrown.info/ - whilst not the prettiest site this contains a lot of very useful and nicely explained information.
- http://www.bbc.co.uk/schools/gcsebitesize/science/ - well presented with many clear diagrams, animations and quizzes. Can occasionally lack depth.
- http://www.chemguide.co.uk/ - whilst mostly targeted at A-Levels this site contains very detailed information suitable for those looking to deepen their knowledge and hit the highest grades.

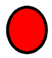
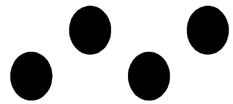
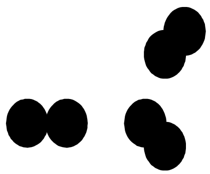
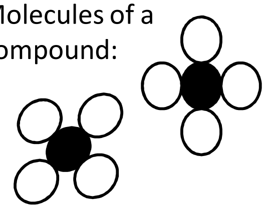
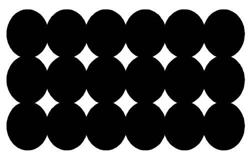
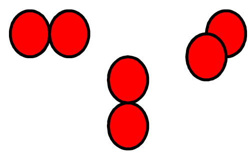
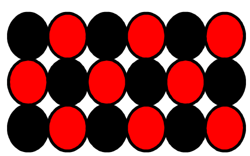
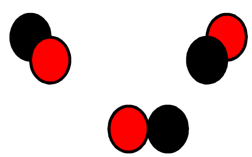
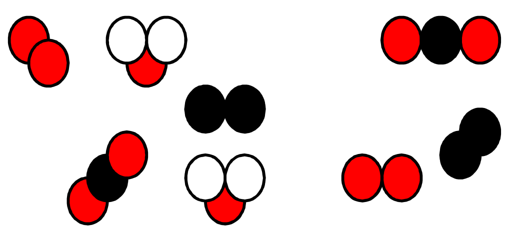
Finally, remember revision is not just reading but should be an active process and could involve:

- Making notes
- Condensing class notes
- Drawing Mind-maps
- Practicing past exam questions
- Making flashcards

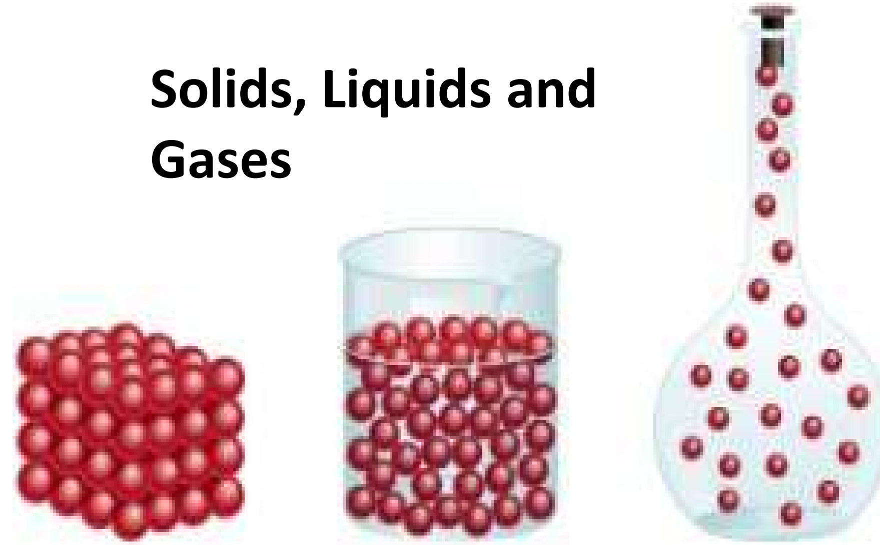
The golden rule is that **what makes you think makes you learn**.

Happy studying, Mr Field.

C1: THE PARTICULATE NATURE OF MATTER

Atom: The smallest particle of matter	An atom: 	Some atoms: 
Molecule: A small particle made from more than one atom bonded together	Molecules of an element: 	Molecules of a compound: 
Element: A substance made of only one type of atom	A solid element: 	A gaseous element: 
Compound: A substance made from two or more different elements bonded together	A solid compound 	A gaseous compound: 
Mixture: A substance made from two or more elements or compounds mixed but not joined	A mixture of compounds and elements: 	

Solids, Liquids and Gases



SOLIDS LIQUIDS AND GASES

The particles in solids, liquids and gases are held near to each other by **forces of attraction**. The strength of these forces determines a substance's melting and boiling points.

In a solid, the forces of attraction are strongest, holding the particles tightly in position. As the solid is heated, and the particles vibrate faster, these forces are partially overcome allowing the particles to move freely as a liquid – this is called **melting**. As the liquid is heated more, the particles gain so much energy that the forces of attraction break completely allowing particles to 'fly around' as a gas – this is called **boiling**. The reverse of the these processes are **condensing** and **freezing**. Under specific conditions, some solids can turn straight to gases – a process called subliming (the reverse is called **desubliming**).

PROPERTIES

Solids
- Have a fixed shape
- Can't be compressed
- Particles close together in a regular pattern
- Particles vibrate around a fixed point

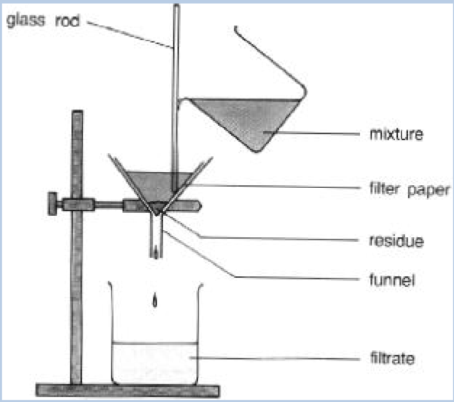
Liquids
- Take the shape of their container
- Can't be compressed
- Particles close together but disordered
- Particles move freely

Gases
- Take the shape of their container
- Can be compressed
- Particles widely spaced in random order
- Particles moving very fast.

C2: EXPERIMENTAL TECHNIQUES

FILTRATION

Used to separate solids from liquids. The mixture is poured through a filter paper in a funnel. The liquid can pass through the small holes in the filter paper (to become the filtrate) and the solid gets left behind (called the residue).



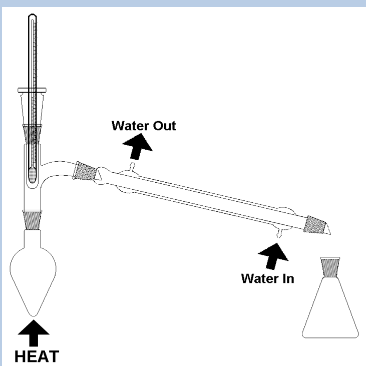
CRYSTALLISATION

Crystallisation is used to separate mixtures of solid dissolved in liquid and relies on the fact that solids are more soluble at higher temperatures. A solution containing a solid is cooled down until crystals form in the solution, these can then be collected by filtration.

The related technique of **recrystallisation** can be used to separate a mixture of two soluble solids by taking advantage of the difference in their solubility. The mixture is dissolved in the smallest possible amount of hot solvent. As the solution cools, the less soluble compound forms crystals that can be collected by filtration whilst the more soluble compound stays dissolved.

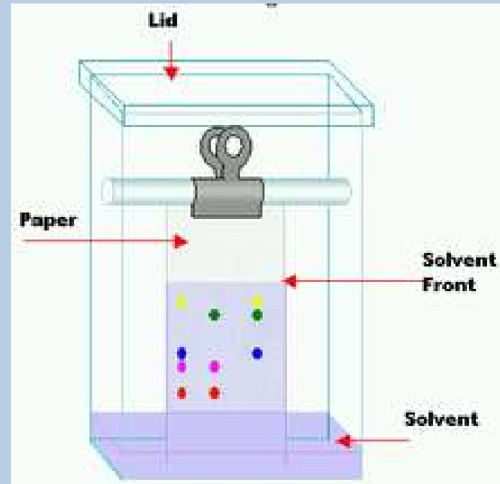
DISTILLATION

In distillation a mixture of liquids is separated using the differences in their boiling points. The mixture is heated until the liquid with the lowest boiling point boils, the vapours then condense on the cold surface of the **condenser** and the pure(er) liquid is collected.



PAPER CHROMATOGRAPHY

Paper chromatography is a technique that can be used to separate mixtures of dyes or pigments and is used to test the purity of a mixture or to see what it contains. Firstly a very strong solution of the mixture is prepared which is used to build up a small intense spot on a piece of absorbent paper. This is then placed in a jar of solvent (with a lid). As the solvent soaks up the paper, it dissolves the mixture-spot, causing it to move up the paper with the solvent. However since compounds have different levels of solubility, they move up the paper at different speeds causing the individual components to separate out. The solvent or combination of solvents can be changed to get the best possible separation of spots.



PURITY

It is important for chemists to be able to purify the compounds they make, this is because the impurities could be dangerous or just un-useful. This is especially true for chemists making compounds that are consumed by people such as drugs or food additives since the impurities may be toxic which would be very bad news!

WHICH TECHNIQUE?

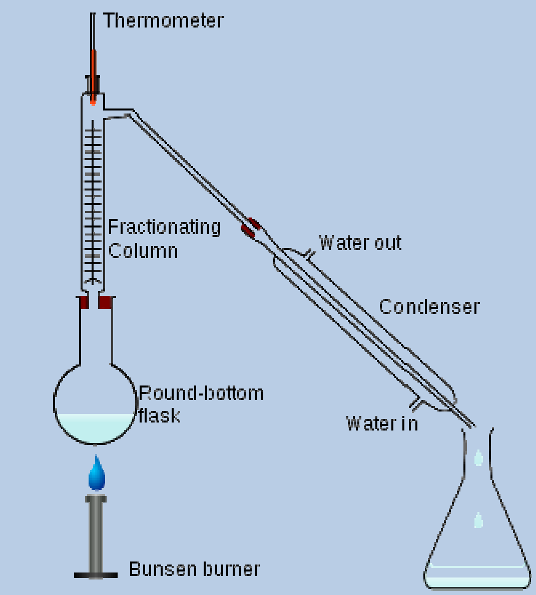
You need to be able to select appropriate methods to separate a given mixture. The key to this is look for differences in the properties of the components of the mixture such as their state, solubility, melting/boiling point and so on. Then pick the method that best takes advantage of this difference.

MELTING/BOILING POINTS

No two substances have the exact same melting and boiling points. We can take advantage of this to test the purity of a compound we have made. If we know what the melting or boiling point of the pure compound should be, we can then measure the melting or boiling point of a sample we have produced and the closer it is to the pure value, the more pure it is likely to be.

FRACTIONAL DISTILLATION

When the liquids being distilled have similar boiling points, normal distillation can't separate them completely but simply gives a purer mixture. In this case a fractionating column is used. This provides a large surface area for condensation meaning much purer 'fractions' are produced. The most important use of this is separating crude oil into it's useful components.

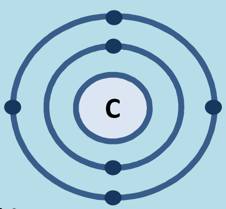


C3: ATOMS, ELEMENTS AND COMPOUNDS – Structures and Bonding

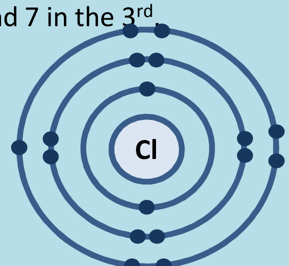
ELECTRON ARRANGEMENT/CONFIGURATION
Electrons are arranged around atoms in specific shells. The most important shell is the outer one as this controls an atom's chemistry. We call the electrons in the outer shell '**valence electrons**' because they are used for bonding. The number of electrons in the outer shell is the same an element's group number.

The number of electrons around an atom is given by the atom's proton number. They are arranged in shells as follows:
- **1st Shell** – Holds two electrons
- **2nd/3rd/4th Shells** – Hold 8 electrons

- **Example 1: Carbon**. Proton number is 6 which means there are 6 electrons: 2 in the 1st shell and 4 in the second

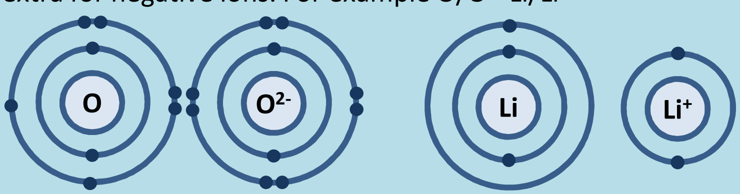


- **Example 2: Chlorine**. Proton number is 17 which means there are 17 electrons: 2 in the 1st shell, 8 in the second and 7 in the 3rd.



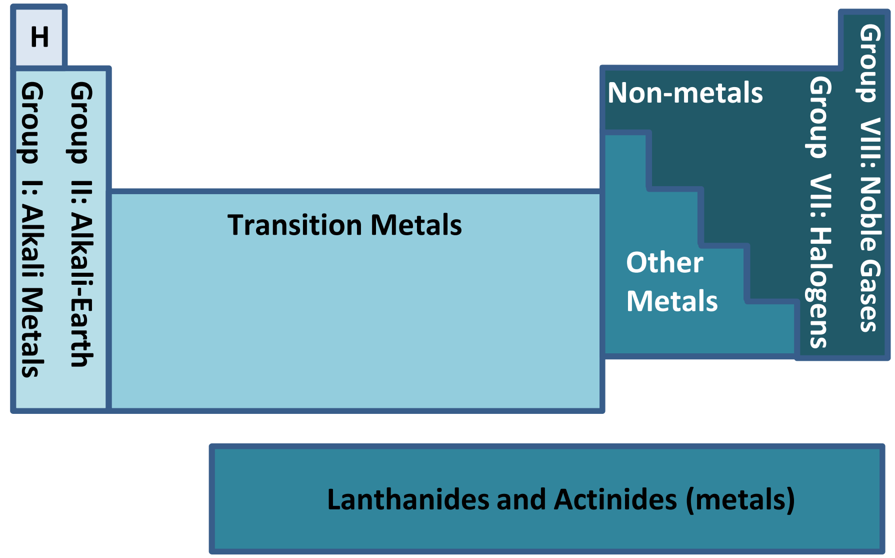
Checking Your Answer: To check you are right, the **period gives the number of shells** and the **group gives the number of electrons in the outer shell**. For example chlorine is in Period 3 and Group VII so it has 3 shells and 7 electrons in the outer shell.

Ions: The configuration of ions is the same as for atoms but you have to take electrons away from positive ions and add extra for negative ions. For example O/O^{2-} Li/Li^{+}



A NOBLE MATTER
The Noble Gases (He, Ne, Ar etc) have full outer shells containing either 2 or 8 electrons. This is **very stable** which is why the Noble gases are so unreactive.

Other elements tend to react in such a way as to achieve a full outer shell by gaining or losing electrons until they achieve this **Noble Gas configuration**.



CHEMICAL VS PHYSICAL CHANGES
Physical changes are reversible whereas **chemical changes** are not.

For example if you melted some solid sugar to a liquid and then left it to cool, it would freeze back to solid sugar – this is a **physical change**. If you took the same sugar and burned it to produce carbon dioxide and water, there would be no easy way to turn those back to sugar – this is a **chemical change** – new substances are made.

STRUCTURE OF THE PERIODIC TABLE (PT on last page!)
Elements arranged in order of increasing proton number.
Periods: The rows in the periodic table.
- For example Li, C and O are all in period 2.

Groups: The columns in the PT.
- Use roman numbers: I, II, III, IV, V, VI, VII, VIII
- Eg. F, Cl, Br, I are all in Group VII
- Elements in the same group have similar properties and react in similar ways: the halogens all react in the same way with sodium to form sodium fluoride (NaF), sodium chloride (NaCl), sodium bromide (NaBr) and sodium iodide (NaI)

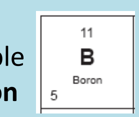
ISOTOPES
Isotopes are atoms with the **same proton number** but **different nucleon** number.

For example carbon has two main isotopes – C-12 and C-13. Carbon has a proton number of 6 so they both contain 6 protons and 6 electrons but C-12 has 6 neutrons and C-13 has 7.

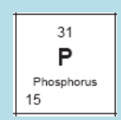
ATOMIC STRUCTURE
Atoms are made of:
Protons: mass = 1, charge = +1
Neutrons: mass = 1, charge = 0
Electrons: mass = 0, charge = -1

The numbers of each vary from element to element but it is the number of protons which decides what the element is.

In a square on the periodic table the smaller number, the **proton number,** gives the **number of protons or electrons** and the bigger number, the **nucleon number** the number of **protons and neutrons together**.



Eg 1: Boron has 5 protons, 6 neutrons (ie 11-5) and 5 electrons



Eg 2: Phosphorus has 15 protons, 16 neutrons (ie 31-16) and 15 electrons

C3: ATOMS, ELEMENTS AND COMPOUNDS – Bonding and Structure

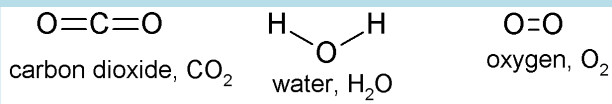
MOLECULES

A molecule is a small particle made from (usually) a few **non-metal** atoms bonded together.

The atoms in a molecule are joined by **strong covalent bonds**. In a solid each molecule is held close to its neighbour by weak **intermolecular forces**.

When a substance melts, it is these weak intermolecular forces that break **NOT** the strong covalent bonds.

Molecular compounds have low melting points and are volatile (evaporate easily) due to the weak intermolecular forces, and insulate electricity as all electrons are stuck in bonds and so unable to move.

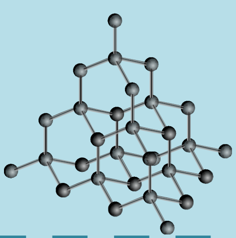


some molecules

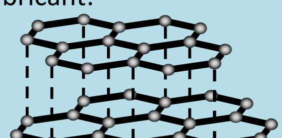
GIANT COVALENT LATTICES

A crystal made of a repeating pattern of atoms **joined with covalent bonds** that repeats millions of times in all directions.

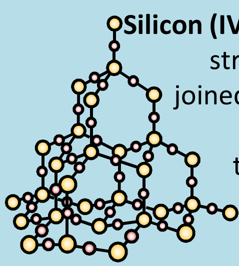
Diamond is made of carbon atoms arranged so that each C is bonded in a pyramid arrangement to 4 others. This makes it very hard, ideal for use in industrial drills:



Graphite: made of carbon atoms arranged in hexagonal sheets with long weak bonds between the sheets. This means the sheets can easily separate making graphite a good lubricant:



Silicon (IV) oxide (SiO_2) has a structure with each Si joined to 4 O and each O joined to 2 Si. It is the main ingredient in glass.

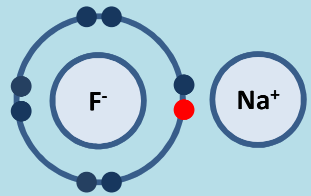


IONIC BONDING

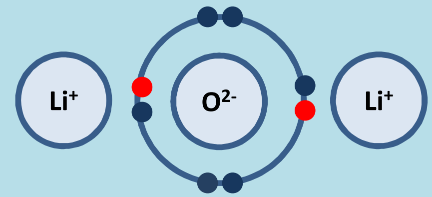
An ionic bond is the **attraction between two oppositely charged ions**. **Cations** (positive) are formed when atoms (usually metals) lose electrons. **Anions** (negative) are formed when atoms (usually non-metals) gain electrons.

Atoms will lose or gain electrons until they have a complete outer shell: elements in Groups I, II and III will lose 1, 2 and 3 electrons respectively to form 1+, 2+ and 3+ ions. Atoms in Groups V, VI and VII gain 3, 2 and 1 electrons to form 3-, 2- and 1- ions. In an ionic compound the number of positive and negative and charges must cancel out to neutral.

Example: NaF, sodium in Group I forms a 1+ ion and fluorine in group VII forms a 1- ion so one Na^+ is needed to balance out one F^-



Example: Li_2O, lithium in Group I forms a 1+ ion but oxygen in Group VI forms a 2- ion so two Li^+ are needed to balance out one O^{2-}

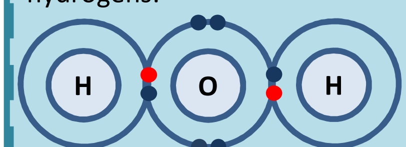


COVALENT BONDING

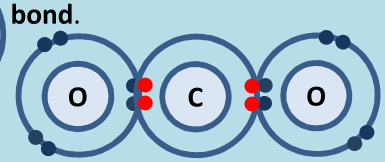
A **covalent bond** forms between two atoms and is the attraction of two atoms to a **shared pair of electrons**. Small groups of covalent bonded atoms can join together to form **molecules**.

The atoms share enough electrons to **complete their outer**

Example: H_2O*, hydrogen is has one valence electron and needs one more to complete the 1st shell, oxygen has six valence electrons electrons so needs two more. Thus one oxygen will react with two hydrogens:



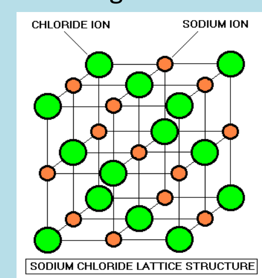
Example: CO_2*, carbon is has four valence electrons so needs four more to complete its outer shell, oxygen needs two more. Thus each carbon will react with two oxygens, sharing two electrons with each one. A bond involving two shared pairs is a **double bond**.



***Nb:** In these diagrams only draw the outer shell and use different shapes/colours to show where electrons have come from. You should be able to draw at least: H_2O, CH_4, Cl_2, HCl, H_2, N_2, O_2, CO_2, C_2H_4

GIANT IONIC LATTICES

The positive and negative ions in an ionic compound don't form molecules but form crystals made of a repeating pattern of positive and negative ions called a **giant ionic lattice**. Eg sodium chloride:



Properties of Ionic Compounds

When you melt or dissolve an ionic compound it conducts electricity because the ions are free to move towards the positive and negative electrodes. When solid the ions are stuck in position and there are no free electrons so they don't conduct.

C4: STOICHIOMETRY – Formulas and Equations

SYMBOL EQUATIONS

•Show the **reactants** you start with and the **products** you make using symbols not words

•Must contain an arrow (→) NOT an equals sign (=)

•Must be balanced – same number of atoms on each side.

•Balancing is done by placing numbers called **coefficients in front** of the formulas for the compounds/elements. For example, 'O_2' means there is one oxygen molecule involved in a reaction but '**2**O_2' would mean there are two.

Example:. $CH_{4(g)} + O_{2(g)} \rightarrow CO_{2)g)} + H_2O_{(g)}$*

This is unbalanced as there are 4 'H' on the left but only 2 'H' on the right. This must be corrected by placing a '2' in front of the 'H_2O' so there are now 2 waters:

$$CH_{4\,(g)} + O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

Now the 'H' balances but there 4 'O' on the right and only 2 on the left. This must be balanced by placing a '2' in front of the 'O_2' so that there are 2 oxygen molecules:

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

Now there is 1 'C', 4 'H' and 4 'O' on each side so it balances.

In **ionic equations**, we tend to look only at the ions that actually change. For example, when iron reacts with copper sulphate to form iron sulphate and copper the equation is:

$$Fe_{(s)} + Cu^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)} \rightarrow Fe^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)} + Cu_{(s)}$$

In this case, the sulphate ion ($SO_4{}^{2-}$) remains unchanged (we call it a spectator ion) so it can be left out of the equation to give:

$$Fe_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Fe^{2+}{}_{(aq)} + Cu_{(s)}$$

This allows us to see more clearly the actual chemical changes taking place.

Note: You **can't change the little numbers** (ie the $_2$ in H_2O) as this changes the compound to something completely different.

*The **state symbols** (s), (l), (g) and (aq) are used to indicate solid, liquid, gas and 'aqueous solution' (dissolved in water).

CHEMICAL FORMULAS

Formulas tell you the atoms that make up a compound

Eg 1. H_2O – two H, one O

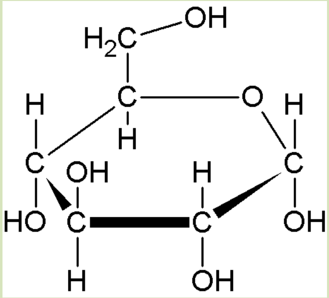
Eg 2. C_2H_6O – two C, six H, one O

Eg 3. $Mg(OH)_2$ – one Mg, two O, two H*

Eg 4. $CH_2(CH_3)_2$ – three C, 8 H*

*In this case everything in brackets is doubled

You may be asked to write a formula given a diagram of a molecule for example glucose. By counting you can see there are 6 carbons, 12 hydrogens and 6 oxygens so the formula is $C_6H_{12}O_6$



WORD EQUATIONS

•These tell you the names of the chemicals involved in reaction

•The left hand side shows you what you start with and is called the **reactants**

•The right hand side shows you what you make and is called the **products**

•The left and right are connected by an arrow (→ not '=') which means 'makes' or 'becomes'

•When you react a metal with oxygen to make a metal oxide, the equation might be:

Iron + oxygen → iron oxide

•Many fuels burn in oxygen to produce carbon dioxide and water for example:

Methane + oxygen → carbon dioxide + water

IONIC FORMULAS

You can deduce the formula of an ionic compound if you know the charges on the ions involved. **The total positive charge must balance out the total negative charge** so you must look for the lowest common multiple (LCM) of the charges.

Eg1. Calcium nitrate is made of Ca^{2+} ions and NO_3^- ions. The LCM of 2 and 1 is 2 which means you need 1 Ca^{2+} ion and 2 NO_3^- ions so the formula is $Ca(NO_3)_2$

Eg2. Aluminium oxide is made of Al^{3+} ions and O^{2-} ions. The LCM of 2 and 3 is 6 which means you need 2 Al^{3+} ions and 3 O^{2-} ions so the formula is Al_2O_3.

CHEMICAL MASSES

The **relative atomic mass (A_r)** of an element is the mass of one atom relative to 1/12th the mass of C-12. It is just a number that allows us to compare the mass of atoms of different elements. A_r can be found on the periodic table as the 'large' number in each square. For example A_r for carbon is 12.01 and for iron is 55.85. Ar has no units since it is only a relative number, allowing us to compare things.

The **relative formula mass (Mr)** is the combined A_r of all the elements in the formula for a substance. M_r also has no units for the same reason as above.

Example 1: Water, H_2O

The A_r for H and O are 1.01 and 16.00 so:

$M_r(H_2O) = 2 \times 1.01 + 1 \times 16.00 = 18.02$

Example 2: Magnesium Hydroxide, $Mg(OH)_2$

The A_r for Mg, O and H are 24.31, 16.00 and 1.01:

$M_r(Mg(OH)_2) = 1 \times 24.31 + 2 \times 16.00 + 2 \times 1.01$
$= 58.33$

Example 3: Decane, $CH_3(CH_2)_8CH_3$

The A_r for C and H are 12.01 and 1.01

$M_r(\text{decane}) = 10 \times 12.01 + 22 \times 1.01 = 142.34$

C4: STOICHIOMETRY – The Mole Concept

THE MOLE

A mole is 6.02×10^{23} of something. It is chosen so that a mole of something has the same mass in grams (molar mass, M_m) as its formula mass. For example the M_r of water is 18.02 so the M_m of water is 18.02g; the M_r of decane is 142.34 so the M_m of decane is 142.34g. Importantly this means that 18.02 g of water and 142.34g decane contains the **same number of molecules**.

EQUATIONS AND MOLE RATIOS

Equations can be used to help us calculate the numbers of moles of substances involved in a reaction. We can see this by studying the following reaction:

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

Q1: How many moles of CO_2 are produced by burning 1.0 mol of C_2H_6? We say that C_2H_6 is our '**known**' and CO_2 is our '**unknown**' so:

Moles CO_2 = moles known/knowns in eqn x unknowns in eqn
= 1.0 /2 x 4 = 1.0 x 2 = 2.0 mol

Q2: If 0.01 mol of CO_2 is produced, how much H_2O must also be produced? This time CO_2 is our known and H_2O is our unknown so:

Moles H_2O = moles known/knowns in eqn x unknowns in eqn
= 0.01/4 x 6 = 0.0025 x 6 = 0.015 mol

*You must make sure your **equation is balanced** or your mole ratio will be wrong.

CALCULATING REACTING QUANTITIES

Using what we know about calculating moles, we can now answer questions like: If I have 100g X, how much Y is made? **The key is to convert the known to moles 1st.**

Example: What volume of H_2 gas would be produced by reacting 12.15g magnesium with excess hydrochloric acid? First we need a balanced equation:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

Then calculate moles of Mg (our known) we start with:

Moles Mg = mass/molar mass = 12.15/24.30 = 0.50 mol

Next we work out how many moles of H_2 (our unknown) we expect to produce:

Moles H_2 = moles known/knowns in eqn x unknowns in eqn
= 0.50/1 x 1 = 0.50 mol

Finally we calculate the volume using our equations for a gas:

Volume H_2 = moles x 24.0 = 0.50 x 24.0 = 12.0 dm^3

LIMITING REACTANTS

This is the reactant that will run out first. It is important as this is the one you should then use for your calculations. You calculate it by dividing the number of moles of reactant by the number of times they appear in the equation. For example $2H_2 + O_2 \rightarrow 2H_2O$. How many moles of H_2O could you make from 3 mol of H_2 and 3 mol of O_2. H_2: 3/2 = 1.5, O_2: 3/1 = 3. This means there is enough O_2 to do the reaction 3 times but only enough H_2 for 1.5 times so H_2 is the limiting reactant. Thus, moles H_2O = 1.5 x (2/2) = 1.5 mol.

THE MOLES AND MASSES

If you know the mass in grams of substance, you can calculate the number of moles as follows:

Moles = Mass */ Molar mass

Eg 1. How many moles is 27.03 g of H_2O?

Moles (H_2O) = Mass / Molar mass = 27.03 / (2 x 1.01 + 16.00) = 1.50 mol

Eg 2. What is the mass of 0.05 mol of H_2O. This time the equation must be rearranged to give:

Mass (H_2O) = Moles x molar mass = 0.05 x (2 x 1.01 + 16.00) = 0.901g

*Mass must be given in grams – you may need to convert from kg: x1000

THE MOLES AND GASES

One mole of any gas has a volume of 24.0 dm^3 (remember dm^3 is the symbol for decimetres cubed, aka litres) at room temperature and pressure. So for a gas:

Moles = Volume / 24.0

Eg 1. How many moles of CO_2 are present in 60 dm^3?

Moles (CO_2) = Volume / 24.0 = 60/24.0 = 2.50 mol

Eg 2. What is the volume of 0.20 mol of H_2 gas?.This time the equation must be rearranged to give:

Volume (H_2) = Moles x 24.0 = 0.20 x 24.0 = 4.80 dm^3

*The volume must be in dm3 – to convert from cm^3 divide by 1000

THE MOLE AND SOLUTIONS

The concentration (strength) of a solution is measured in mol dm^{-3} (moles per decimetre cubed). A 1.0 mol dm^{-3} solution contains 1 mol of substance dissolved in each litre.

Moles = Concentration x Volume*

Eg 1. How many moles of NaOH are present in 2.5 dm^3 of a 1.5 mol dm^{-3} solution?

Moles (NaOH) = concentration x volume = 1.5 x 2.5 = 3.75 mol

Eg 2. 0.15 mol NaCl is dissolved in 250 cm^3 water. What concentration is this? This time you must rearrange the equation to:

Concentration = moles/volume = 0.15/(250/1000)* = 0.60 mol dm^{-3}

*The volume must be in dm^3 – to convert from cm^3 divide by 1000

C5: ELECTRICITY AND CHEMISTRY

	Molten Salt	Salt Solution
Cathode	Metal	Metal, except with reactive metals (K, Na, Li Ca, Mg) in which case H_2 gas is produced plus a solution of metal hydroxide
Anode	Non-metal	Non Metal, except sulphates in which case O_2

ELECTROLYSIS OF COPPER SULPHATE

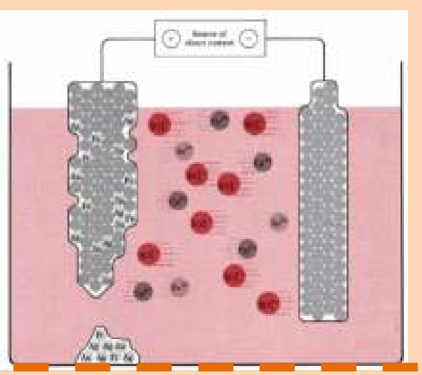
When copper sulphate is electrolysed using carbon electrodes, you produce O_2 gas at the anode and a layer of Cu metal at the cathode. This can be used to **electroplate** items by setting them as the cathode. However, when two copper electrodes are used, what ends up happening is a transfer of copper from the anode to the cathode, this is used to purify copper.

When copper is made it contains lots of impurities. The copper is purified by electrolysis. A large lump of **impure copper is used as the anode**, the electrolyte is copper sulphate solution and the cathode is made of pure copper.

At the anode, instead of anions losing electrons, neutral copper atoms lose electrons to become copper ions .

$$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$$

These then move through the electrolyte to the cathode where they become copper atoms again.



$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$$

The anode loses mass as copper atoms leave it and the cathode gains mass as copper atoms join it. The impurities sink to the bottom as a pile of sludge.

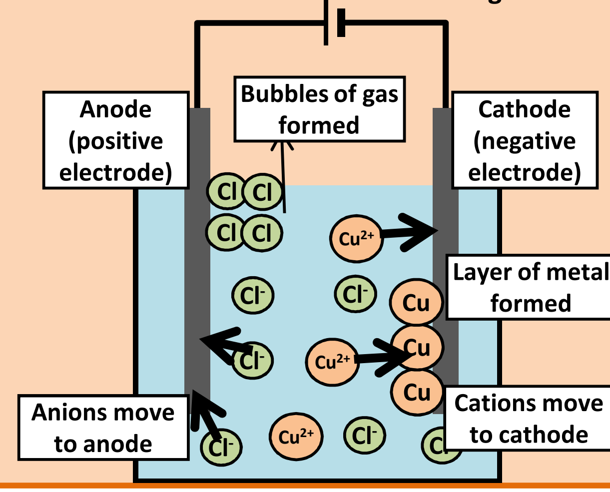
ELECTROLYSIS

Electrolysis is a process in which electricity is used to break compounds down into their elements. The mixture being electrolysed is called an **electrolyte** and must be liquid (either melted or dissolved) to allow the ions to move**.**

Cations (positive ions – remember they are 'puss-itive') ions move to the cathode (the negative electrode) where they gain electrons, usually forming a metal (or H).

Anions (negative ions) move to the anode (the positive electrode) where they lose electrons, usually forming a non-metal (other than H).

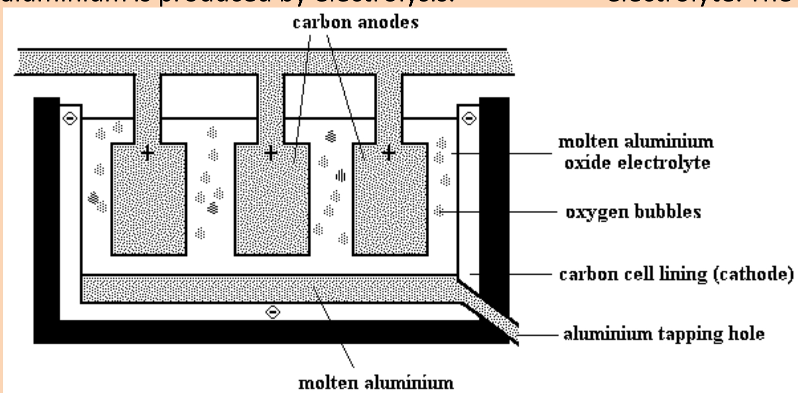
In the electrolysis of **copper chloride** ($CuCl_2$) (right) positive **copper ions** move to the **cathode** and form **copper metal**. Negative **chloride ions** more to the **anode** and form **chlorine gas**.



EXTRACTING ALUMINIUM

Aluminium can't be extracted by reduction of aluminium oxide (Al_2O_3) using carbon as carbon is less reactive than aluminium. Instead aluminium is produced by electrolysis.

Aluminium oxide (the electrolyte) is dissolved in molten 'cryolite' and placed in a large carbon lined vessel which acts as the cathode. A large anode made of carbon is lowered into the electrolyte. The processes that take place are:



ELECTROLYTIC PRODUCTION OF ALUMINIUM.

At the cathode:

Aluminium ions gain electrons to make liquid aluminium

$Al^{3+} + 3e^- \rightarrow Al_{(l)}$

At the anode:

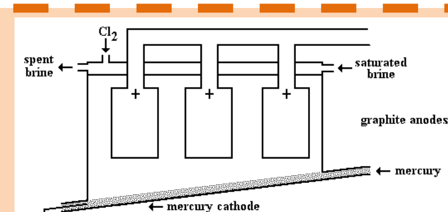
Oxide ions lose electrons to make oxygen gas

$O^{2-} \rightarrow \frac{1}{2} O_{2(g)} + 2e^-$

The oxygen reacts with the carbon anode so it has to be replaced regularly

ELECTROLYSIS OF BRINE

When sodium chloride solution (brine) is electrolysed , chlorine gas is produced at the anode and hydrogen gas at the cathode (because sodium is too reactive). A solution of sodium hydroxide is left behind.



INDUSTRIAL ELECTROLYSIS OF SODIUM CHLORIDE SOLUTION.

C6: ENERGY CHANGES IN CHEMICAL REACTIONS

EXOTHERMIC REACTIONS

Exothermic reactions get hotter – the temperature increases. The energy given out can be used to keep the reaction going so that once started, they don't stop until they have run out of reactants.

Important examples of exothermic reactions include:

- Combustion of fuels
- Acid-base neutralisations
- Displacement reactions
- Respiration in cells

ENDOTHERMIC REACTIONS

Endothermic reactions reactions get colder – the temperature decreases. Generally endothermic reactions need a constant energy supply to keep them going

Important examples of exothermic reactions include:

- Dissolving of many (but not all) salts
- Thermal decompositions
- Photosynthesis
- Cooking!!!

ENERGY CHANGES

In **exothermic** reactions, chemical energy stored in the reactants gets converted to heat energy. The products have less chemical energy than the reactants and the difference is the amount of heat released.

In **endothermic** reactions, heat energy gets converted to chemical energy. The products have more chemical energy than the reactants and the difference between the two is the energy that had to be supplied to make the reaction go.

QUANTIFYING ENERGY

Using the ideas you learn in physics about specific heat capacity, you may have to calculate the amount of energy released by one mole of a substance.

Example: When 0.250 mol of Metal X reacts fully with 500 cm^3 of 2.0 mol dm^{-3} HCl solution, the temperature increases by 15.4^{O}C. How much energy is released when 1.0 mol X reacts with HCl?

First calculate the heat evolved:

Heat evolved = m.c.ΔT = 500 x 4.2 x 15.4 = 32340 J*

Then calculate heat released per mole:

Heat per mole = heat evolved / moles = 32340/0.250 = 129360J = 129.4 kJ

***ΔT** is the temperature rise, **m** is the mass of the solution in grams which is assumed to equal its volume in cm^3, **c** is the specific heat capacity of water which is 4.2 J K^{-1} g^{-1}

Yes this unit really is this small – in fact you don't even really need the stuff about quantifying energy, I just put it in there as it often proves useful!!

C7: CHEMICAL REACTIONS

RATES OF REACTION

The ‘speed’ of a reaction is called its rate and is simply the amount of new product formed every second.

For a chemical reaction to happen, the **reacting particles need to collide with enough energy**. Anything that increases the number of collisions or their energy will increase the rate.

Temperature

Increasing temperature increases the rate of a reaction.

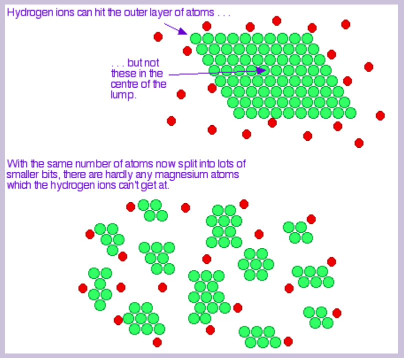
This is because particles are moving faster which means more collisions and **higher energy** collisions.

Concentration

Increasing the concentration of a solution increases the rate of a reaction.

This is because it means there are more particles available to react which leads to **more collisions**.

Surface Area/Particle size



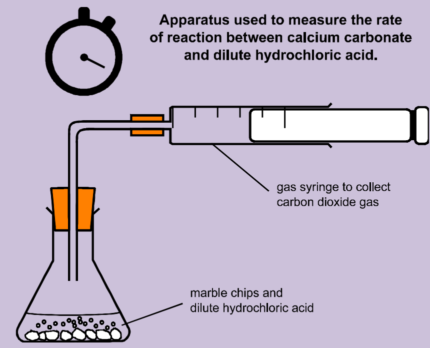
Increasing the total surface area of particles (by using finer powder) increases the rate of a reaction because it means **more particles at the surface are exposed to collisions**.

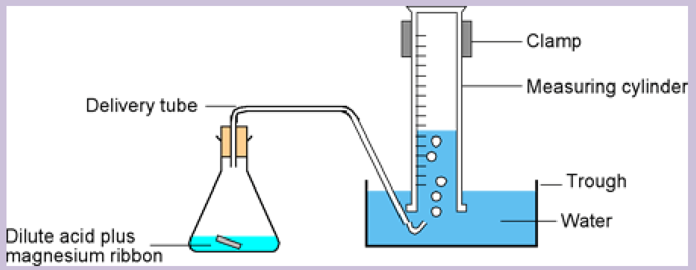
Catalysts

Catalysts are substances that speed up a reaction without getting used up. Whenever a catalyst is present, the rate of reaction increases.

MEASURING REACTION RATES

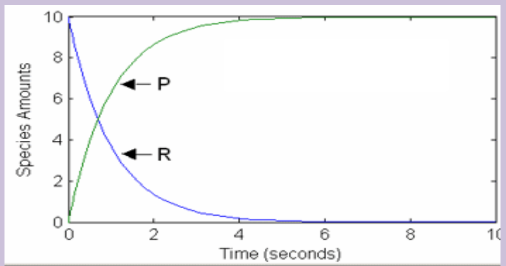
If a reaction produces gas, you can easily measure the reaction rate by collecting the gas (either in an upturned measuring cylinder full of water or a gas syringe) and recording how much has been collected each second.





MEASURING REACTION RATES

On a graph showing the change in concentration of reactants or products, the gradient of the line tells you the reaction rate: **steeper = faster, flat = stopped**



INVESTIGATING REACTION RATES

To investigate a factor influencing reaction rate, you must change it whilst keeping the others constant. For example, investigating the effect of concentration, you could carry out the reaction at 5 different concentrations whilst making sure the temperature, particle size and presence/absence of a catalyst remains the same.

REDOX REACTIONS

Reduction means a substances loses oxygen. **Oxidation** means a substance gains oxygen. For example:

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

Fe_2O_3 is reduced because it loses oxygen to become Fe. **C is oxidised** because it gains oxygen to become CO_2. C is called a reducing agent because it causes Fe_2O_3 to get reduced. Reactions like this are called **redox reactions** because an oxidation AND a reduction take place together.

Another way to look at this is to think of **oxidation as the loss of electrons** and **reduction as the gain of electrons (OILRIG)**. Eg: in the electrolysis of molten sodium bromide. At the anode:

$$2Br^- \rightarrow Br_2 + 2e^-$$

This is an oxidation because the bromide ions lose electrons. At the cathode:

$$Na^+ + e^- \rightarrow Na$$

This is a reduction because the sodium ions gain electrons.

DANGEROUS RATES

Factories that produce flammable powders (for example bread flour) have to be careful about sparks since the very fine powder particles burn with a VERY high reaction rate causing explosions.

Similar is true underground in coal mines where gas can build up. Gas can be thought of as the finest possible powder so they too react explosively fast.

C8: ACIDS, BASES AND SALTS – Reactions of Acids

REACTIONS OF ACIDS
You need to memorise these reactions, each one shows the general word equation then a specific example with symbols.

Acids and Metals

Acid + Metal → Salt + Hydrogen

- Hydrochloric acid + lithium → lithium chloride + hydrogen
- $2HCl(aq) + 2Li(s) \rightarrow 2LiCl_{(aq)} + H_{2(g)}$

Acids and Base (like alkali but not always soluble)

Acid + Base → Salt + Water

- Sulphuric acid + sodium hydroxide → sodium sulphate + water
- $H_2SO_4(aq) + 2NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

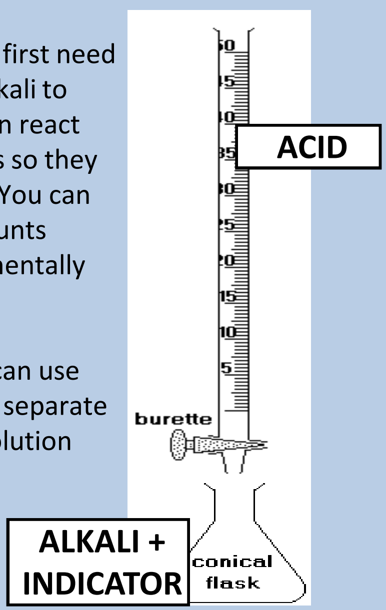
Acids and Carbonates

Acid + Carbonate → Salt + Water + Carbon Dioxide

- Nitric acid + calcium carbonate → calcium nitrate + water + carbon dioxide
- $2HNO_{3(aq)} + CaCO_{3(s)} \rightarrow Ca(NO_3)_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

PREPARING SALTS
To prepare any given salt, you first need to work out which acid and alkali to react together (see right). Then react them in appropriate quantities so they exactly neutralise each other. You can either **calculate** the right amounts (see Unit C4) or find it experimentally from a **titration**.

Once you have done this you can use the appropriate techniques to separate the salt from the rest of the solution (See Unit C2).

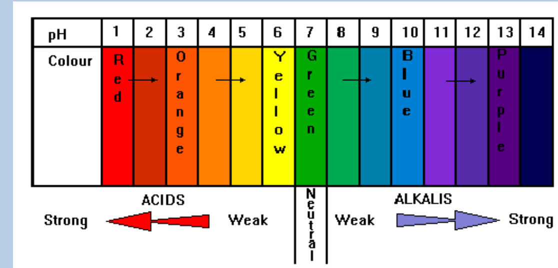


THE pH SCALE
Neutral substances have a pH=7
Acids have a pH of **less than 7**
Alkalis have a pH **greater than 7**

pH can be measured with colour changing **indicators** or **digital pH meters**

Litmus indicator turns red in acids and blue in alkalis.

Universal indicator has many colours (see chart).



WHAT IS THE SALT?
To work out which salt is formed during neutralisation reactions you need to know the ions formed by the acid or alkali when it dissolves.

Substance	Cation(s) Formed	Anion(s) Formed
Hydrochloric acid, HCl	1 H^+	Cl^- , chloride
Nitric acid, HNO_3	1 H^+	NO_3^- , nitrate
Sulphuric acid, H_2SO_4	2H^+	SO_4^{2-} , sulphate
Phosphoric acid, H_3PO_4	3 H^+	PO_4^{3-} , phosphate
Sodium hydroxide, $NaOH$	Na^+ , sodium	1 OH^-
Potassium hydroxide, KOH	K^+ , potassium	1 OH^-
Magnesium hydroxide, $Mg(OH)_2$	Mg^{2+} , magnesium	2 OH^-
Ammonium hydroxide, NH_4OH	NH_4^+ , ammonium	1 OH^-

Working out the name is easy, you just combine the name of the **cation from the alkali** with the **anion from the acid**.

For example potassium sulphate and sulphuric acid makes potassium sulphate.

Magnesium hydroxide and phosphoric acid makes magnesium phosphate

Working out the formula of the salt is a little more complicated, the key is to make sure the positive and negative charges on the cancel each other out to zero.

Eg 1. Potassium nitrate
K^+ has one plus charge
SO_4^{2-} has two minus charges

You need two K^+ to balance out one SO_4^{2-} so the formula is $\mathbf{K_2SO_4}$

Eg 2. Magnesium phosphate
Mg^{2+} has two plus charges
PO_4^{3-} has three minus charges

So you need three Mg^{2+} to balance out two PO_4^{3-} so the formula is $\mathbf{Mg_3(PO_4)_2}$

Finally, to write a balanced equation, you need to get the right number of waters, the simplest way is to remember that each 'H^+' from an acid makes one water.

Eg 1. Potassium hydroxide and sulphuric acid
As we have seen it makes K_2SO_4 which requires one H_2SO_4 and two KOH. Two H_2O are made since the one H_2SO_4 produces two H^+ ions

$$H_2SO_4 + 2KOH \rightarrow K_2SO_4 + 2H_2O$$

Eg 2. Magnesium phosphate
As we have seen it makes $Mg_3(PO_4)_2$ which requires two H_3PO_4 and three $Mg(OH)_2$. Six H_2O are made since each of the two H_3PO_4 produces three H^+ ions.

$$2H_3PO_4 + 3Mg(OH)_2 \rightarrow Mg_3(PO_4)_2 + 6H_20$$

C8: ACIDS, BASES AND SALTS – Chemical Testing

TESTING GASES

Hydrogen:
•A test tube of hydrogen produces a 'squeaky pop' with a lighted splint

Oxygen:
•A test tube of oxygen can re-light a glowing splint.

Chlorine:
•Bleaches the colour from damp litmus paper.

Ammonia:
•Turns damp red litmus paper blue.

Carbon dioxide:
•Turns limewater cloudy.

OXIDES

The oxides of most metals are **basic** (the opposite of acidic). For example sodium oxide (Na_2O) forms the alkali sodium hydroxide when it reacts with water.

Most oxides of non-metals are acidic. For example, sulphur trioxide (SO_3) forms sulphuric acid when it dissolves in water.

Some oxides can behave like acids or bases and are called **amphoteric**. For example aluminium oxide (Al_2O_3) can react with the alkali NaOH to from sodium aluminium hydroxide ($NaAl(OH)_4$) or with hydrochloric acid to form aluminium chloride ($AlCl_3$)

ACID ENVIRONMENTS

Acid soils grow poor crops so the acidity is reduced by neutralising it with lime (CaO, calcium oxide)

Acidic gases from factory chimneys (like sulphur dioxide) can dissolve in the water in clouds to form harmful acid rain.

TESTING FOR IONS: Most of these involve forming insoluble precipitates – they go cloudy.

Test for...	By....	Positive result	The reaction
Chloride ions, Cl^-	Add acidified silver nitrate	White precipitate	Forms insoluble silver chloride: $Cl^-_{(aq)} + AgNO_{3(aq)} \rightarrow AgCl_{(s)} + NO_3^-{}_{(aq)}$
Sulphate ions, SO_4^{2-}	Add acidified barium nitrate	White precipitate	Insoluble barium sulphate formed: $SO_4^{2-}{}_{(aq)} + Ba(NO_3)_{2(aq)} \rightarrow BaSO_{4(s)} + 2NO_3^-{}_{(aq)}$
Carbonate ions, CO_3^{2-}	Add acid and bubble the gas formed in limewater	Rapid gas formation which turns limewater cloudy	The acid reacts with carbonate to make carbon dioxide gas: $CO_3^{2-}{}_{(s)} + 2H^+{}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$ The CO_2 reacts with limewater to make insoluble calcium carbonate.
Nitrate ions, NO_3^-	Boil with sodium hydroxide and aluminium foil. Test the gas with damp red litmus paper.	Red litmus paper turns blue	The nitrate gets reduced by aluminium which is a strong reducing agent and forms ammonia. Ammonia is an alkali so can turn the red litmus paper blue.
Copper (II), Cu^{2+}	Add sodium hydroxide followed by ammonia solution	Blue precipitate that dissolves when ammonia added	Insoluble copper (II) hydroxide formed: $Cu^{2+}{}_{(aq)} + 2NaOH_{(aq)} \rightarrow Cu(OH)_{2(s)} + 2Na^+{}_{(aq)}$ When ammonia is added a soluble complex forms so the precipitate dissolves.
Iron (II), Fe^{2+}	Add sodium hydroxide followed by ammonia solution.	Green precipitate insoluble in ammonia	Insoluble iron (II) hydroxide formed: $Fe^{2+}{}_{(aq)} + 2NaOH_{(aq)} \rightarrow Fe(OH)_{2(s)} + 2Na^+{}_{(aq)}$ Ammonia does not react with the iron (II) hydroxide so it does not dissolve.
Iron (III), Fe^{3+}	Add sodium hydroxide followed by ammonia solution.	Brown precipitate insoluble in ammonia	Insoluble iron (III) hydroxide formed: $Fe^{3+}{}_{(aq)} + 3NaOH_{(aq)} \rightarrow Fe(OH)_{3(s)} + 3Na^+{}_{(aq)}$ Ammonia does not react with the iron (III) hydroxide so it does not dissolve.
Zinc, Zn^{2+}	Add sodium hydroxide followed by ammonia solution or more sodium hydroxide.	White precipitate soluble in both ammonia or more sodium hydroxide	Insoluble zinc hydroxide formed: $Zn^{2+}{}_{(aq)} + 2NaOH_{(aq)} \rightarrow Zn(OH)_{2(s)} + 2Na^+{}_{(aq)}$ Both ammonia and sodium hydroxide react with the zinc hydroxide to form a soluble complex.

C9: THE PERIODIC TABLE

THE PERIODIC TABLE

The periodic table is arranged in order of increasing proton number – starting at Hydrogen with a proton number of one and working along the rows.

Periods:
The rows in the periodic table are called periods. Going along a period, the elements change from metals to non-metals. Usually, one or two elements in the period are called **metalloids** – these have some properties of a metal and some properties of a non-metal.

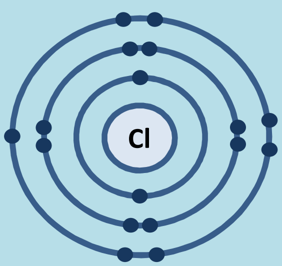
Groups:
These are the columns in the periodic table. Elements in the same group share similar properties.

Groups I and II are always metals. Groups VII and 0/VIII are always non-metals and elements in groups III, IV, V and VI can be metals, metalloids or non-metals depending on the period.

The Periodic Table and Atomic Structure:
The periodic table can be used to work out the arrangement of electrons:

- **Period number** = number of shells
- **Group number** = electrons in outer shell

For example: Chlorine is in Period 3 and Group VII so it has 3 electron shells and 7 electrons in the outer shell.



GROUP I (Li, Na, K....)

The metals of Group I (aka the **alkali metals**) are soft, silvery grey, reactive metals. **Down the group** they get:

- **Softer**
- **Lower melting point**
- **More reactive**

They all react with water as follows:

Metal + water → metal hydroxide + hydrogen

- Lithium + water → lithium hydroxide + hydrogen
- Li + H_2O → LiOH + H_2
- Lithium reacts the slowest, Na reacts faster, K reacts faster still and so on.

GROUP VII (F, Cl, Br, I...)

The elements of Group VII are better known as the **halogens**. As we go **down the group** they get:

- **Less reactive**
- **Higher melting point** (Cl_2 is gas, Br_2 is liquid, I_2 is solid)
- **Darker colour** (Cl_2 is pale green, Br_2 is reddy-brown, I_2 is dark brown)

They will react with ions of other halogens (halide ions) that are below them in the group. For example:

$$Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$$

Because Cl is more reactive than Br. However,

$$Br_2 + Cl^- \rightarrow \text{no reaction}$$

As Br is less reactive than Cl.

TRANSITION ELEMENTS

These are the metals in the long middle block of the periodic table.

Their important properties include:

- **High melting/boiling points**
- **High densities**
- Form strongly **coloured compounds**
- (Often) **Act as catalysts** – both as elements and when combined in compounds

GROUP 0/VIII (He, Ne, Ar, Kr....)

The gases of Group 0 are called the Noble Gases because they are very unreactive. This is because they have full outer shells of electrons which is very stable.

They exist as single atoms rather than molecules.

They are used whenever an inert (unreactive) atmosphere is needed. For example:

- Light Bulbs – Argon surrounds the coiled filament as even when white hot, it won't react.

Helium has a very low density (1/7th that of air) so is used to make airships and blimps float.

METALS (really belongs in C10 but didn't quite fit)

Most of the known elements are metals.

All metals:

Conduct electricity, conduct heat, are shiny

Most metals are also:

- Malleable – can be beaten into shape
- Strong
- High melting/boiling point
- Sonorous – 'ring' when hit
- Ductile – can be pulled into wires

Many metals react with:

- Acids – to form salt and hydrogen
- Oxygen – to form (basic) oxides
- Sulphur – to form sulphides

When metals bond to non metals they form ionic bonds.

C10: METALS

EXTRACTING METALS FROM THEIR ORES

Rocks that contain a significant amount of a metal are called **ores**. The metals are present as compounds – often oxides or sulphides of the metal. For example **lead** can be extracted from an ore called **galena** (PbS, lead sulphide).

Metals that are less reactive than carbon can be extracted by using carbon as a reducing agent (to steal the oxygen/ sulphur). More reactive metals are extracted by electrolysis.

Iron is less reactive than carbon so can be reduced by it. This is done in a **blast furnace**. Study the diagram then read the following:

- **Step 1: Carbon** (coke) reacts with **oxygen** (from the hot air blast)

  $C_{(s)} + O_{2(g)} \rightarrow CO_2(g)$

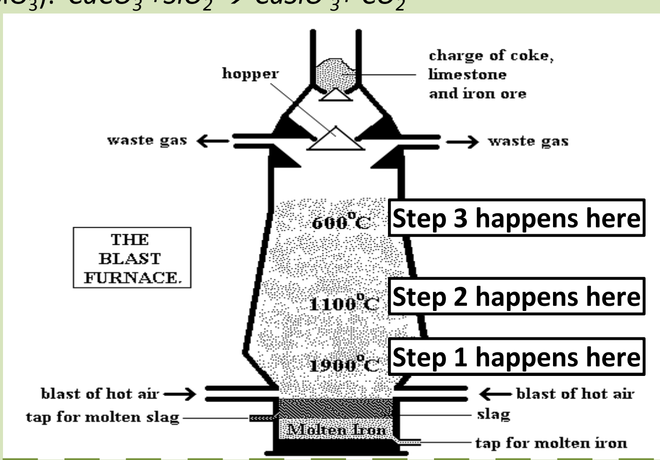
- **Step 2: Carbon dioxide** reacts with more **carbon** to make **carbon monoxide**

  $CO_{2(g)} + C_{(s)} \rightarrow 2CO_{(g)}$

- **Step 3: Carbon monoxide** reduces the **iron oxide** (iron ore) to make molten **liquid iron**.

  $Fe_2O_{3(s)} + CO_{(g)} \rightarrow Fe_{(l)} + CO_{2(g)}$

The **limestone ($CaCO_3$)** reacts with impurities such as silicon to form an easy-to-collect waste called **slag** (calcium silicate, $CaSiO_3$): $CaCO_3 + SiO_2 \rightarrow CaSiO_3 + CO_2$



REACTIVITY OF METALS

The reactivity of metals can be seen by the way they react with steam or with acid (see Unit C6 for the reactivity series).

Reaction with water (see Unit C2 for details of this reaction):
The most reactive metals (K-Ca) react with cold water, fairly reactive metals (Mg-Fe) will only react with steam whereas the least reactive metals (Sn-Pt) don't react at all.

Reaction with dilute acids (see Unit C9 for details)
The reaction of metals with acids shows a similar patter with the most reactive metals (K-Ca) reacting violently, the fairly reactive metals (Mg-Pb) reacting gradually more slowly and the least reactive metals (Cu-Pt) not reacting at all.

Displacement Reactions

The reactivity of metals relates to how easily they form ions, more reactive metals like K form K^+ ions much more easily than less reactive metals like Cu can form Cu^+ ions. A more reactive metal will reduce a less reactive metal:

Eg 1. Reaction with aqueous ions

Zinc + Copper sulphate → Zinc sulphate + copper

$Zn_{(s)} + Cu^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} + Cu^{2+}_{(aq)}$

This happens because Zn is more reactive than Cu so is able to reduce it.

Eg 2. Reaction with metal oxides

Iron oxide + aluminium → aluminium oxide + iron

This happens since Al is more reactive than Fe so is able to reduce it.

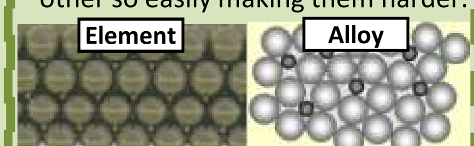
These are called displacement reactions because the more reactive metal takes the place of the less reactive metal.

ALLOYS

Alloys are '**mixtures of metals**' (although sometimes they can contain a non-metal) that are made by mixing molten metals.

Alloys often have very different properties to the metals they are made from and by varying their metals can be tailored to have specific desirable properties – this is called **metallurgy**.

Alloys are often harder than the metals they are made from. In pure metals atoms are neatly lined up meaning they can slip past each easily when hit. In alloys there are atoms of different sizes which don't line up neatly so can't slip past each other so easily making them harder.



REACTIVITY SERIES

MOST REACTIVE

Potassium, 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, Pt

LEAST REACTIVE

USES OF METALS

Metals have many uses including:

- **Aluminium** – and its alloys used for aircraft as they have low density and great strength
- **Aluminium** – used for food containers as the waterproof oxide layer on its surface prevents corrosion which could taint the food.
- **Zinc -** used to protect steel either by coating it (**galvanising**) or as **sacrificial protection** – i.e. on a ship's hull – a lump of zinc prevents rust as it is more reactive so corrodes instead of the steel hull.

C11: AIR AND WATER

CARBON DIOXIDE, CO_2

There are many ways to produce CO_2 including:

Burning carbon-containing fuels:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Thermal decomposition of carbonates e.g.:

$$CaCO_3 \rightarrow CaO + CO_2$$

As a by product of **respiration** in living cells:

$$C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$$

WATER, H_2O

Water is the most useful compound known to man. **In the home** it is used for cooking, cleaning and transporting waste. **In industry** it is used for cooling hot machinery, cleaning and as a solvent. Water is useful for cleaning as it is able to dissolve many types of 'dirt'.

A simple test for water is that it is able to turn cobalt chloride paper from blue to pink.

Drinking Water

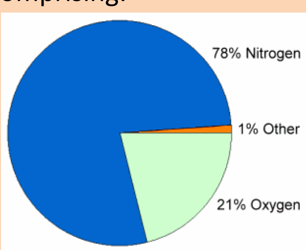
Water drawn from rivers can contain pollutants such as fertilizers, dissolved organic matter, harmful bacteria and industrial waste that make it unfit to drink. At treatment plants, two main processes are used to make water safe:

Filtration – the water is passed through a series of increasingly fine filters that trap suspended particles. Activated carbon is used to filter out dissolved pollutants.

Chlorination – chlorine is added to the water which destroys bacteria.

AIR

Air is a mixture of gases comprising:



The 'other' is mostly argon with CO_2, water vapour and many trace gases.

Although the proportion of carbon dioxide is very small (~0.04%) it is increasing due to man's activities such as burning fossil fuels and deforestation. This is a concern as CO_2 is able to absorb the infrared radiation (heat) radiated by the ground when the sun heats it up (the **greenhouse effect**). More CO_2 means more trapped heat leading to **global warming**.

Global warming is a major problem because temperatures are rising faster than nature's ability to adapt – this makes the future of both farming and of our ecosystems very uncertain.

AIR POLLUTION

Many of man's activities pollute the air. Pollutants include:

Carbon monoxide, CO

- Formed when fuels burn without enough O_2.
- CO prevents the blood from carrying oxygen leading to death by suffocation

Sulphur dioxide, SO_2

- Formed by burning fossil fuels containing sulphur impurities.
- Dissolves in water in clouds to form sulphurous acid which falls as acid rain
- Acid rain corrodes buildings and damages ecosystems
- Irritates the respiratory system when inhaled.

Nitrogen Oxides, NO_x

- Formed by burning fuels in engines and power stations.
- Dissolves in cloud water to form nitric acid thus acid rain.
- Irritates the respiratory system when inhaled.

NITROGEN AND AMMONIA

Ammonia (NH_3) is a smelly gas. One way to produce it is to react ammonium (NH_4^+) salts with an alkali (OH^-) eg:

$$NH_4Cl + NaOH \rightarrow NH_3 + H_2O + NaCl$$

Ammonia is vital to produce the nitrates used in fertilisers and explosives. It is produced by the **Haber process**:

$$N_{2(g)} + 3H_{2(g)} \leftrightarrows 2NH_3$$

The reaction is reversible which means much of the product turns back to reactants as soon as it is made, this means it takes a long time to make an economical amount of ammonia. To speed it up, the reaction is done at **high temperature** (~450^{o}C) with an **iron oxide catalyst**. **High pressure** (~200 times atmospheric pressure) is used to increase the proportion of NH_3 formed.

The nitrogen comes from the air and hydrogen comes from reacting methane (CH_4) gas with steam.

Nitrogen and **oxygen** can be separated from air by cooling it to a liquid and using **fractional distillation**.

RUSTING

Rust (hydrated iron (III) oxide) affects most structures made of iron (or steel) and causes huge damage:

Iron + oxygen + water → hydrated iron (III) oxide

Rust can be prevented by taking steps making sure either oxygen or water can't reach the iron. The main ways to do this involve covering the metal with: **paint** (bridges and other structures); **oil/grease** (moving machine parts) or another metal such as zinc (**galvanising**).

FERTILISERS

Fertilisers are chemicals applied to plants to improve their growth and increase the amounts of products such as fruits, nuts, leaves, roots and flowers that they produce for us. They work by supplying plants with the vital elements they need including **Nitrogen** - in the form or **nitrate** (**NO_3^-** containing) salts; **phosphorous** – in the form of **phosphate** (**PO_4^{3-}** containing) salts and **potassium** (**K^+** containing) salts.

CATALYTIC CONVERTERS

Fit to a car's exhaust and use a **platinum** or **palladium catalyst** to convert harmful gases to safer gases: for example nitrogen oxides are reduced back to nitrogen gas and oxygen gas.

C12: SULPHUR

SULPHURIC ACID, H_2SO_4

Sulphuric acid is a very important compound used in many industrial processes including:

- Fertiliser production
- Oil refining
- Paper making
- Steel making

It is also the acid found in car batteries.

Sulphuric acid is a strong acid which when diluted in water produces two protons and a sulphate ion:

$$H_2SO_{4(l)} \rightarrow 2H^+_{(aq)} + SO_4^{2-}{}_{(aq)}$$

It exhibits all the reactions typical of an acid as seen by its reactions with metals, alkalis, metal oxides and carbonates. (see Unit C8 for details).

THE CONTACT PROCESS

Sulphuric acid is produced by the **Contact Process**.

This involves are three chemical reactions. First sulphur is burnt in air to produce sulphur dioxide (SO_2):

$$S + O_2 \rightarrow SO_2$$

Secondly SO_2 is reacted with further oxygen to make sulphur trioxide (SO_3):

$$2SO_2 + O_2 \leftrightarrows SO_3$$

This reaction is reversible, so to maximise the amount of SO_3 made, they use a **high temperature** (425^{O}C), **medium-high pressure** (1-2 times atmospheric pressure) and a **catalyst** (**vanadium (V) oxide**, V_2O_5).

Finally, the sulphur trioxide is produced by first dissolving it in sulphuric acid to make oleum ($H_2S_2O_7$) which then makes more sulphuric acid on the addition of water:

$$SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$$

$$H_2S_2O_7 + H_2O \rightarrow 2H_2SO_4$$

Note: trying to dissolve SO_3 directly in water produces a very fine mist of sulphuric with limited uses.

This is another tiny unit with very little to learn.

C13: CARBONATES

CALCIUM CARBONATE, CaCO3

Calcium carbonate is a very common mineral and makes up the bulk of many common rocks including:

- Chalk
- Limestone
- Marble

Whilst solid limestone is often used in construction, powdered limestone has many industrial uses.

USES OF CALCIUM CARBONATE

Powdered calcium carbonate can be added directly to **soils to raise their pH** (reduce their acidity).

We can also make **calcium oxide** (CaO, aka 'quicklime') by heating powdered calcium carbonate to about 1000^{O}C, producing carbon dioxide as a by-product:

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$$

This is called a **thermal decomposition** as heat is used to break down or decompose the calcium carbonate. Calcium oxide is one of the key ingredients in **cement**.

Another useful product, **calcium hydroxide** ($Ca(OH)_2$, 'slaked lime') is made by adding water to calcium oxide:

$$CaO + H_2O \rightarrow Ca(OH)_2$$

Slaked lime has many uses including:

- **Raising soil pH quickly** (when powdered calcium carbonate might take too long)
- **Neutralising acidic industrial waste**
- **Sewage treatment** – it helps small particles of waste to clump together into easily removed lumps.

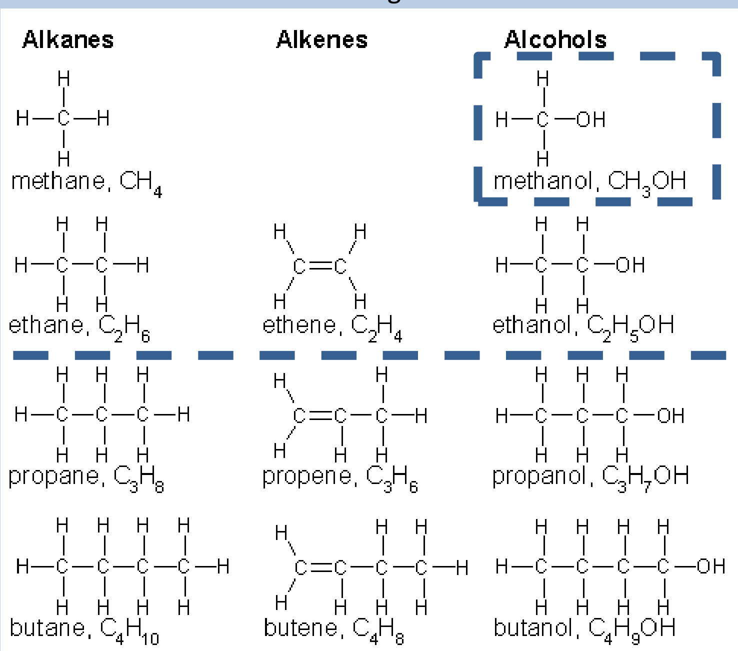
Another mini-unit with very little in it!

C14: ORGANIC CHEMISTRY - Oil

CHEMICAL FAMILIES

Organic chemistry is the chemistry of compounds containing carbon.

You need to know the structure of four organic compounds: **methane**, **ethane**, **ethene** and **ethanol** (check the diagram below). Methane and ethane are both members of the '**alkane**' family – you can tell this because their names end '**–ane**'. Ethene is an **alkene**, as shown by the '**–ene**' ending and ethanol is an alcohol which has the ending '**–ol**'.

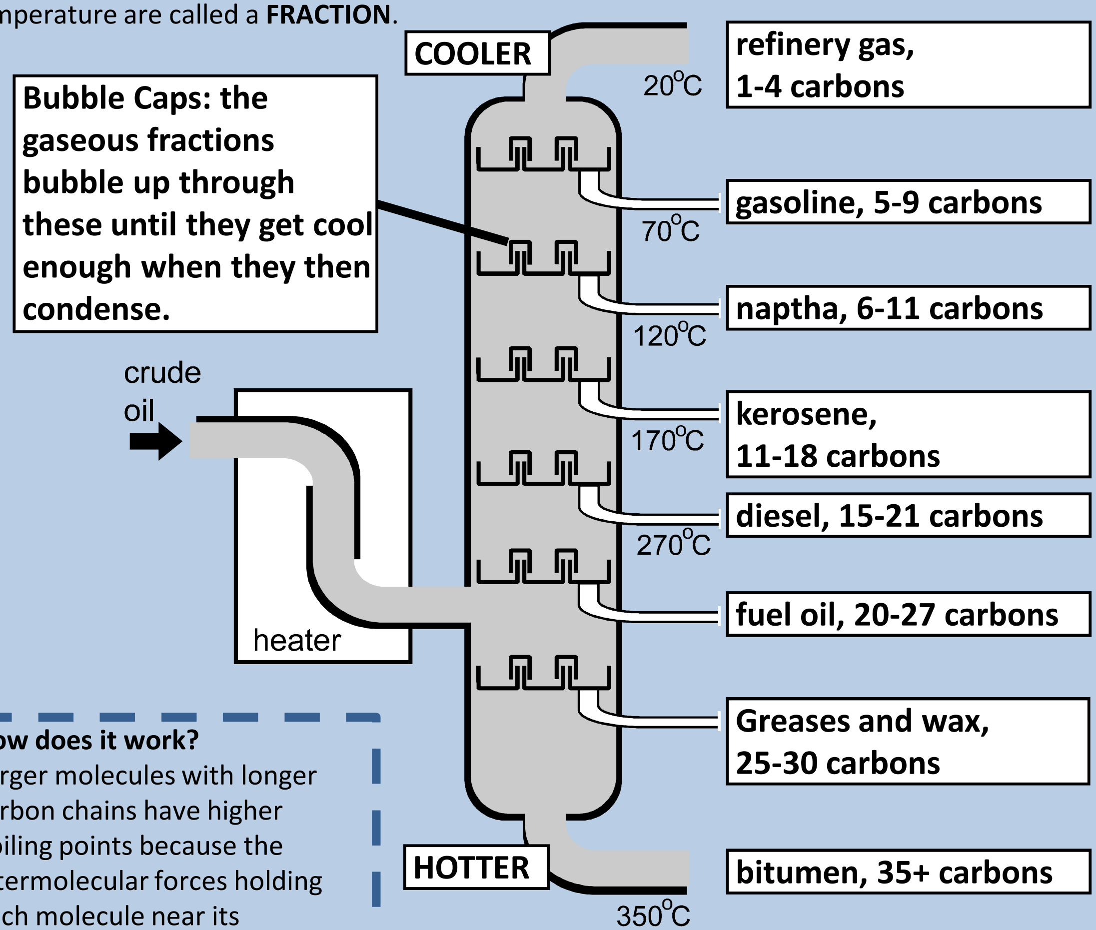


Homologous Series:

These are families of compounds that differ only in the length of their carbon chain. For example, looking at the diagram above you can see all alcohols contain an '-OH' group bonded to a carbon, all alkenes contain a 'C=C' double bond and all alkanes contain only single C-C and C-H bonds. The beginning of a name tells you the number of carbons in the chain: 'meth' means 1 C, 'eth' means 2, 'prop' is 3 and 'but' is 4 carbons.

OIL

Oil is a mixture of hundreds of **hydrocarbons** (compounds containing only H and C). This mixture must be separated into its useful components by **fractional distillation**. Very hot crude oil is pumped into the **fractionating column** where the hydrocarbons separate out by their boiling points, rising through the column until they get cold enough to condense. The compounds that condense at a particular temperature are called a **FRACTION**.



How does it work?

Larger molecules with longer carbon chains have higher boiling points because the intermolecular forces holding each molecule near its neighbour are stronger so take more energy to break.

Three important fractions:

Refinery gas: this is bottled and used for cooking and heating

Gasoline: the petrol used to fuel our cars

Diesel oil: used in diesel engines – particularly for large vehicles

Fossil Fuels:

Coal, oil and natural gas are all fossil fuels formed by the action of heat and pressure over millions of years on the remains of living organisms. All of them release carbon dioxide when burnt which contributes to global warming. Because coal is contains the most carbon, it also produces the most carbon dioxide so is not an environmentally sustainable fuels. Natural gas (made mostly of methane, CH_4) contains much less carbon and so is an environmentally better fuel.

C14: ORGANIC CHEMISTRY – Classes of Compounds

HYDROCARBONS

Hydrocarbons are compounds made of only hydrogen and carbon atoms.

Hydrocarbons – for example methane (CH_4) – burn very well producing only carbon dioxide (CO_2) and water (H_2O):

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Alkanes (see structure on previous page)

These are the simplest hydrocarbons. They are '**saturated**' which means they only contain single bonds. They are pretty unreactive but burn well making them good fuels.

Alkenes (see structure on previous page)

These are hydrocarbons containing a C=C double bond. The double bond makes them quite reactive and they are used as a starting material to make many other organic compounds.

Addition reaction of alkenes with bromine:

When an orange solution of bromine is added to alkenes in the presence of UV light, the bromine reacts with the double bond on the alkene to make a bromoalkane. The bromine water loses its colour so this makes it a good test for alkenes:

$$C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$$

Addition reaction of alkenes with steam:

Ethene reacts with **steam** in the presence of a **phosphoric acid catalyst** to make ethanol which can be used as a solvent or to make other useful compounds.

$$C_2H_{4(g)} + H_2O_{(g)} \rightarrow C_2H_5OH_{(g)}$$

Addition reaction of alkenes with hydrogen:

Alkenes reacts with **hydrogen** in the presence of a **nickel catalyst** to make alkanes.

$$C_2H_{4(g)} + H_2O_{(g)} \rightarrow C_2H_{6\ (g)}$$

Whilst not very useful in itself, this reaction applies to C=C double bonds in much more complex molecules to, and for example is one of the key steps in producing margarine.

ALCOHOLS

Alcohols such as ethanol are very important compound with many uses including as solvents and fuels. They can be made from alkenes (see left) by reacting them with steam.

Alcohols burn very cleanly producing very little soot and smoke:

$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$$

CRACKING

Because there is a greater need for hydrocarbons with shorter carbon chains we sometimes need to cut longer chains into shorter ones using the process of **cracking**.

A long alkane is **heated**, **vaporised** and passed over a **ceramic catalyst** produce a shorter alkane and an alkene.

Eg. 1: $C_8H_{18} \rightarrow C_4H_{10} + C_4H_8$

Eg. 2: $C_{10}H_{22} \rightarrow C_7H_{16} + C_3H_6$

Note:

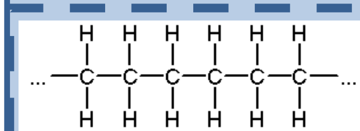
- The with the alkenes for each carbon there are 2 H (C_nH_{2n}); with the alkanes, for each C there are 2 H plus 2 extra (C_nH_{2n+2}).
- Any combination of alkene and alkane can be made, including straight and branched chains, so long as the numbers of atoms balance.

MACROMOLECULES

These are large molecules made from lots of smaller molecules – called monomers - joined together. Different monomers lead to different macromolecules.

Polythene

Poly(ethene) is a synthetic polymer (plastic) made from many ethene molecules joined together. It is formed by addition polymerisation whereby many individual monomers (in this case ethene) join together in one long chain.



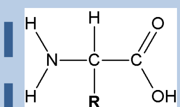
a 3-monomer length of poly(ethene)

The diagram shows a section of poly(ethene) made from three ethene monomers joined.

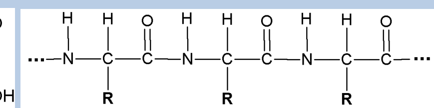
Natural Macromolecules

Proteins and starch are both examples of natural macromolecules. In a protein the monomer is various different amino acids:

They are condensation polymers where the 'acid' end of one amino acid joins to the 'amino' end of the next, forming an amide linkage and one molecule of water each time.



amino acid



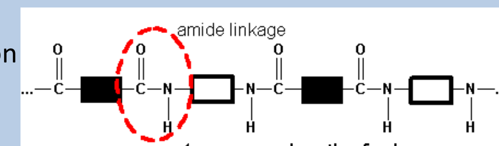
a 3-monomer length of a protein

Proteins can be broken back down to amino acids by strong acids or strong alkalis. This process is called hydrolysis.

Nb: The '**R**' on an amino acid means any small group of atoms as is different in each amino acid.

Condensation Polymers

In condensation polymerisation, each time two monomers join, one molecule of water is produced. In the case of nylon (pictured) there are two monomers – one with two acid ends (-COOH, black) and one with two amine ends (-NH2, white). They join with an 'amide' linkage, producing water.



a 4-monomer length of nylon

The Periodic Table of the Elements

Group

I	II											III	IV	V	VI	VII	0
							1 **H** Hydrogen 1										4 **He** Helium 2
7 **Li** Lithium 3	9 **Be** Beryllium 4											11 **B** Boron 5	12 **C** Carbon 6	14 **N** Nitrogen 7	16 **O** Oxygen 8	19 **F** Fluorine 9	20 **Ne** Neon 10
23 **Na** Sodium 11	24 **Mg** Magnesium 12											27 **Al** Aluminium 13	28 **Si** Silicon 14	31 **P** Phosphorus 15	32 **S** Sulphur 16	35.5 **Cl** Chlorine 17	40 **Ar** Argon 18
39 **K** Potassium 19	40 **Ca** Calcium 20	45 **Sc** Scandium 21	48 **Ti** Titanium 22	51 **V** Vanadium 23	52 **Cr** Chromium 24	55 **Mn** Manganese 25	56 **Fe** Iron 26	59 **Co** Cobalt 27	59 **Ni** Nickel 28	64 **Cu** Copper 29	65 **Zn** Zinc 30	70 **Ga** Gallium 31	73 **Ge** Germanium 32	75 **As** Arsenic 33	79 **Se** Selenium 34	80 **Br** Bromine 35	84 **Kr** Krypton 36
85 **Rb** Rubidium 37	88 **Sr** Strontium 38	89 **Y** Yttrium 39	91 **Zr** Zirconium 40	93 **Nb** Niobium 41	96 **Mo** Molybdenum 42	**Tc** Technetium 43	101 **Ru** Ruthenium 44	103 **Rh** Rhodium 45	106 **Pd** Palladium 46	108 **Ag** Silver 47	112 **Cd** Cadmium 48	115 **In** Indium 49	119 **Sn** Tin 50	122 **Sb** Antimony 51	128 **Te** Tellurium 52	127 **I** Iodine 53	131 **Xe** Xenon 54
133 **Cs** Caesium 55	137 **Ba** Barium 56	139 **La** Lanthanum 57 *	178 **Hf** Hafnium 72	181 **Ta** Tantalum 73	184 **W** Tungsten 74	186 **Re** Rhenium 75	190 **Os** Osmium 76	192 **Ir** Iridium 77	195 **Pt** Platinum 78	197 **Au** Gold 79	201 **Hg** Mercury 80	204 **Tl** Thallium 81	207 **Pb** Lead 82	209 **Bi** Bismuth 83	**Po** Polonium 84	**At** Astatine 85	**Rn** Radon 86
Fr Francium 87	226 **Ra** Radium 88	227 **Ac** Actinium 89 †															

*58-71 Lanthanoid series
†90-103 Actinoid series

140 **Ce** Cerium 58	141 **Pr** Praseodymium 59	144 **Nd** Neodymium 60	**Pm** Promethium 61	150 **Sm** Samarium 62	152 **Eu** Europium 63	157 **Gd** Gadolinium 64	159 **Tb** Terbium 65	162 **Dy** Dysprosium 66	165 **Ho** Holmium 67	167 **Er** Erbium 68	169 **Tm** Thulium 69	173 **Yb** Ytterbium 70	175 **Lu** Lutetium 71
232 **Th** Thorium 90	**Pa** Protactinium 91	238 **U** Uranium 92	**Np** Neptunium 93	**Pu** Plutonium 94	**Am** Americium 95	**Cm** Curium 96	**Bk** Berkelium 97	**Cf** Californium 98	**Es** Einsteinium 99	**Fm** Fermium 100	**Md** Mendelevium 101	**No** Nobelium 102	**Lr** Lawrencium 103

Key

a **X** b

a = relative atomic mass
X = atomic symbol
b = proton (atomic) number

The volume of one mole of any gas is 24 dm^3 at room temperature and pressure (r.t.p.).