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:

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

- [http://www.docbrown.info/](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/](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/](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.

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.

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

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

## 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 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**).

### SOLIDS
- **Properties**
  - Have a fixed shape
  - Can’t be compressed
  - Particles close together in a regular pattern
  - Particles vibrate around a fixed point

### LIQUIDS
- **Properties**
  - Take the shape of their container
  - Can’t be compressed
  - Particles close together but disordered
  - Particles move freely

### GASES
- **Properties**
  - Take the shape of their container
  - Can be compressed
  - Particles widely spaced in random order
  - Particles moving very fast
C2: EXPERIMENTAL TECHNIQUES

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.

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.

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.

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.

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).

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**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 as the 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/O²⁻ Li/Li⁺

**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)

**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.**

**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.

**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.**

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**Examining Atoms**
The number of electrons is the same as the proton number. The number of neutrons is different for each isotope.

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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.

**Some molecules**
- Carbon dioxide, CO₂
- Water, H₂O
- Oxygen, O₂

**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:

Silicon (IV) oxide (SiO₂) has a structure with each Si joined to 4 O and each O joined to 2 Si. It is the main ingredient in glass.

**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.

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:

Example: H₂O*, hydrogen is joined to 2 O and each O joined to 2 H. It is the main ingredient in glass.

*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: H₂O, CH₄, Cl₂, HCl, H₂, N₂, O₂, CO₂, C₂H₂.

**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:

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.

**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 ions 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: Li₂O, 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 O²⁻.
**C4: STOICHIOMETRY – Formulas and Equations**

**CHEMICAL FORMULAS**
Formulas tell you the atoms that make up a compound

Eg 1. H₂O – two H, one O
Eg 2. C₂H₆O₂ – two C, six H, one O
Eg 3. Mg(OH)₂ – one Mg, two O, two H*
Eg 4. CH₂(CH₃)₂ – three C, 8 H*

*In this case everything in brackets is doubled

**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 (→) 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.

Eg 1. Calcium nitrate is made of Ca²⁺ ions and NO₃⁻ ions. The LCM of 2 and 1 is 2 which means you need 1 Ca²⁺ ion and 2 NO₃⁻ ions so the formula is Ca(NO₃)₂
Eg 2. Aluminium oxide is made of Al³⁺ ions and O²⁻ ions. The LCM of 2 and 3 is 6 which means you need 2 Al³⁺ ions and 3 O²⁻ ions so the formula is Al₂O₃.

**CHEMICAL MASSES**
The relative atomic mass (Aₐ) 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. Aₐ can be found on the periodic table as the ‘large’ number in each square. For example Aₐ 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.

Example 1: Water, H₂O
The Aₐ for H and O are 1.01 and 16.00 so:
Mᵣ(H₂O) = 2 x 1.01 + 1 x 16.00 = 18.02

Example 2: Magnesium Hydroxide, Mg(OH)₂
The Aₐ for Mg, O and H are 24.31, 16.00 and 1.01:
Mᵣ(Mg(OH)₂) = 1 x 24.31 + 2 x 16.00 + 2 x 1.01 = 58.33

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

Example 3: Decane, CH₃(CH₂)₁₆CH₃
The Aₐ for C and H are 12.01 and 1.01
Mᵣ(decane) = 10 x 12.01 + 22 x 1.01 = 142.34

**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, ‘O₂’ means there is one oxygen molecule involved in a reaction but ‘2O₂’ would mean there are two.

Example: CH₄(g) + O₂(g) → CO₂(g) + H₂O(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 ‘H₂O’ so there are now 2 waters:
CH₄(g) + 2O₂(g) → CO₂(g) + 2H₂O(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 ‘O₂’ so that there are 2 oxygen molecules:
CH₄(g) + 2O₂(g) → CO₂(g) + 2H₂O(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) + Cu²⁺(aq) + SO₄²⁻(aq) → Fe²⁺(aq) + SO₄²⁻(aq) + Cu(s)
In this case, the sulphate ion (SO₄²⁻) remains unchanged (we call it a spectator ion) so it can be left out of the equation to give:
Fe(s) + Cu²⁺(aq) → Fe²⁺(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 H₂O) 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).
C4: STOICHIOMETRY – The Mole Concept

THE MOLE
A mole is $6.02 \times 10^{23}$ 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 $M_m$ of water is 18.02 so the $M_m$ of water is 18.02 g; the $M_m$ of decane is 142.34 so the $M_m$ of decane is 142.34 g. Importantly this means that 18.02 g of water and 142.34 g decane contain 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$_2$H$_6$? We say that C$_2$H$_6$ is our ‘known’ and CO$_2$ is our ‘unknown’ so:

$$\text{Moles CO}_2 = \text{moles known/knowns in eqn} \times \text{unknowns in eqn}$$

$$= 1.0/2 \times 4 = 1.0 \times 2 = 2.0 \text{ mol}$$

Q2: If 0.01 mol of CO$_2$ is produced, how much H$_2$O must also be produced? This time CO$_2$ is our known and H$_2$O is our unknown so:

$$\text{Moles H}_2O = \text{moles known/knowns in eqn} \times \text{unknowns in eqn}$$

$$= 0.01/4 \times 6 = 0.0025 \times 6 = 0.015 \text{ mol}$$

*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:

$$\text{Moles Mg} = \text{mass/molar mass} = 12.15/24.30 = 0.50 \text{ mol}$$

Next we work out how many moles of H$_2$ (our unknown) we expect to produce:

$$\text{Moles H}_2 = \text{moles known/knowns in eqn} \times \text{unknowns in eqn}$$

$$= 0.50/1 \times 1 = 0.50 \text{ mol}$$

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

$$\text{Volume H}_2 = \text{moles} \times 24.0 = 0.50 \times 24.0 = 12.0 \text{ dm}^3$$

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 H$_2$ by 1.5 times so H$_2$ is the limiting reactant. Thus, moles H$_2$O = 1.5 x (2/2) = 1.5 mol.

THE MOLES AND MASSES
If you know the mass in grams of a substance, you can calculate the number of moles as follows:

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar mass}}$$

Eg 1. How many moles is 27.03 g of H$_2$O?

$$\text{Moles (H}_2\text{O)} = \frac{\text{Mass}}{\text{Molar mass}} = \frac{27.03}{2 \times 1.01 + 16.00} = 1.50 \text{ mol}$$

Eg 2. What is the mass of 0.05 mol of H$_2$O. This time the equation must be rearranged to give:

$$\text{Mass (H}_2\text{O)} = \text{Moles} \times \text{molar mass} = 0.05 \times (2 \times 1.01 + 16.00) = 0.901 \text{ g}$$

*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:

$$\text{Moles} = \frac{\text{Volume}}{24.0}$$

Eg 1. How many moles of CO$_2$ are present in 60 dm$^3$ gas?. This time the equation must be rearranged to give:

$$\text{Volume (CO}_2\text{)} = \text{Moles} \times 24.0 = \frac{0.20}{1} \times 24.0 = 4.80 \text{ dm}^3$$

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

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

$$\text{Moles} = \text{Concentration} \times \text{Volume}$$

Eg 1. How many moles of NaOH are present in 2.5 dm$^3$?

$$\text{Moles (NaOH)} = \text{Concentration} \times \text{Volume} = 1.5 \times 2.5 = 3.75 \text{ mol}$$

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:

$$\text{Concentration} = \frac{\text{moles}}{\text{volume}} = \frac{0.15/(250/1000)}{1} = 0.60 \text{ mol dm}^{-3}$$

*The volume must be in dm$^3$ – to convert from cm$^3$ divide by 1000
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 (CuCl₂)

EXTRACTING ALUMINIUM

Aluminium can’t be extracted by reduction of aluminium oxide (Al₂O₃) using carbon as carbon is less reactive than aluminium. Instead aluminium is produced by electrolysis.

When copper sulphate is electrolysed using carbon electrodes, you produce O₂ 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.

\[ \text{Cu}_\text{(s)} \rightarrow \text{Cu}^{2+} \text{(aq)} + 2e^- \]

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

\[ \text{Cu}^{2+} \text{(aq)} + 2e^- \rightarrow \text{Cu}_\text{(s)} \]

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.

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.

ELECTROLYSIS OF COPPER SULPHATE

When copper sulphate is electrolysed using carbon electrodes, you produce O₂ 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.

\[ \text{Cu}_\text{(s)} \rightarrow \text{Cu}^{2+} \text{(aq)} + 2e^- \]

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

\[ \text{Cu}^{2+} \text{(aq)} + 2e^- \rightarrow \text{Cu}_\text{(s)} \]

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.
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 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$^\circ$C. How much energy is released when 1.0 mol X reacts with HCl?

First calculate the heat evolved:

\[
\text{Heat evolved} = m \cdot c \cdot \Delta T = 500 \times 4.2 \times 15.4 = 32340 \text{ J}^* 
\]

Then calculate heat released per mole:

\[
\text{Heat per mole} = \frac{\text{heat evolved}}{\text{moles}} = \frac{32340}{0.250} = 129360 \text{ J} = 129.4 \text{ kJ}
\]

\*\(\Delta T\) 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.

In 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.

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 substance loses oxygen. Oxidation means a substance gains oxygen. For example:

\[ 2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2 \]

Fe₂O₃ is reduced because it loses oxygen to become Fe. C is oxidised because it gains oxygen to become CO₂. C is called a reducing agent because it causes Fe₂O₃ 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**

---

**REATIONS 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) → 2LiCl(aq) + H₂(g)

**Acids and Base (like alkali but not always soluble)**

- Acid + Base → Salt + Water
- Sulphuric acid + sodium hydroxide → sodium sulphate + water
  - H₂SO₄(aq) + 2NaOH(aq) → Na₂SO₄(aq) + 2H₂O(l)

**Acids and Carbonates**

- Acid + Carbonate → Salt + Water + Carbon Dioxide
- Nitric acid + calcium carbonate → calcium nitrate + water + carbon dioxide
  - 2HNO₃(aq) + CaCO₃(s) → Ca(NO₃)₂(aq) + H₂O(l) + CO₂(g)

---

**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.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Cation(s) Formed</th>
<th>Anion(s) Formed</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrochloric acid, HCl</td>
<td>1 H⁺</td>
<td>Cl⁻, chloride</td>
</tr>
<tr>
<td>Nitric acid, HNO₃</td>
<td>1 H⁺</td>
<td>NO₃⁻, nitrate</td>
</tr>
<tr>
<td>Sulphuric acid, H₂SO₄</td>
<td>2H⁺</td>
<td>SO₄²⁻, sulphate</td>
</tr>
<tr>
<td>Phosphoric acid, H₃PO₄</td>
<td>3 H⁺</td>
<td>PO₄³⁻, phosphate</td>
</tr>
<tr>
<td>Sodium hydroxide, NaOH</td>
<td>Na⁺, sodium</td>
<td>1 OH⁻</td>
</tr>
<tr>
<td>Potassium hydroxide, KOH</td>
<td>K⁺, potassium</td>
<td>1 OH⁻</td>
</tr>
<tr>
<td>Magnesium hydroxide, Mg(OH)₂</td>
<td>Mg²⁺, magnesium</td>
<td>2 OH⁻</td>
</tr>
<tr>
<td>Ammonium hydroxide, NH₄OH</td>
<td>NH₄⁺, ammonium</td>
<td>1 OH⁻</td>
</tr>
</tbody>
</table>

---

**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).

---

### Acids and Metals

- Hydrochloric acid + lithium → lithium chloride + hydrogen
  - 2HCl(aq) + 2Li(s) → 2LiCl(aq) + H₂(g)

### Acids and Base

- Sulphuric acid + sodium hydroxide → sodium sulphate + water
  - H₂SO₄(aq) + 2NaOH(aq) → Na₂SO₄(aq) + 2H₂O(l)

### Acids and Carbonates

- Nitric acid + calcium carbonate → calcium nitrate + water + carbon dioxide
  - 2HNO₃(aq) + CaCO₃(s) → Ca(NO₃)₂(aq) + H₂O(l) + CO₂(g)

### 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.

- Hydrochloric acid, HCl → 1 H⁺, Cl⁻
- Nitric acid, HNO₃ → 1 H⁺, NO₃⁻
- Sulphuric acid, H₂SO₄ → 2H⁺, SO₄²⁻
- Phosphoric acid, H₃PO₄ → 3 H⁺, PO₄³⁻
- Sodium hydroxide, NaOH → Na⁺, 1 OH⁻
- Potassium hydroxide, KOH → K⁺, 1 OH⁻
- Magnesium hydroxide, Mg(OH)₂ → Mg²⁺, 2 OH⁻
- Ammonium hydroxide, NH₄OH → NH₄⁺, 1 OH⁻

### Working out the name

- potassium nitrate: K⁺ has one plus charge
  - SO₄²⁻ has two minus charges
  - You need two K⁺ to balance out one SO₄²⁻ so the formula is K₂SO₄

### Working out the formula of the salt

- 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 nitrate**

- K⁺ has one plus charge
- SO₄²⁻ has two minus charges
- You need two K⁺ to balance out one SO₄²⁻ so the formula is K₂SO₄

**Eg 2. Magnesium phosphate**

- Mg²⁺ has two plus charges
- PO₄³⁻ has three minus charges
- So you need three Mg²⁺ to balance out two PO₄³⁻ so the formula is Mg₃(PO₄)₂

---

**Eg 1. Potassium hydroxide and sulphuric acid**

As we have seen it makes K₂SO₄ which requires one H₂SO₄ and two KOH. Two H₂O are made since the one H₂SO₄ produces two H⁺ ions

- H₂SO₄ + 2KOH → K₂SO₄ + 2H₂O

**Eg 2. Magnesium phosphate**

As we have seen it makes Mg₃(PO₄)₂ which requires two H₃PO₄ and three Mg(OH)₂. Six H₂O are made since each of the two H₃PO₄ produces three H⁺ ions.

- 2H₃PO₄ + 3Mg(OH)₂ → Mg₃(PO₄)₂ + 6H₂O
### TESTING FOR IONS: Most of these involve forming insoluble precipitates – they go cloudy.

<table>
<thead>
<tr>
<th>Test for...</th>
<th>By...</th>
<th>Positive result</th>
<th>The reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Chloride ions, Cl(^-)</td>
<td>Add acidified silver nitrate</td>
<td>White precipitate</td>
<td>Forms insoluble silver chloride: (\text{Cl}^-<em>{(aq)} + \text{AgNO}<em>3</em>{(aq)} \rightarrow \text{AgCl}</em>(s) + \text{NO}<em>3^-</em>{(aq)})</td>
</tr>
<tr>
<td>Sulphate ions, (\text{SO}_4^{2-})</td>
<td>Add acidified barium nitrate</td>
<td>White precipitate</td>
<td>Insoluble barium sulphate formed: (\text{SO}<em>4^{2-}</em>{(aq)} + \text{Ba(NO}<em>3)</em>{2(aq)} \rightarrow \text{BaSO}_4(s) + 2\text{NO}<em>3^-</em>{(aq)})</td>
</tr>
</tbody>
</table>
| Carbonate ions, \(\text{CO}_3^{2-}\) | 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: \(\text{CO}_3^{2-}_{(s)} + 2\text{H}^+_{(aq)} \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l)\)
| Nitrate ions, \(\text{NO}_3^-\) | 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. |

### OXIDES
The oxides of most metals are basic (the opposite of acidic). For example sodium oxide (Na\(_2\)O) 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\(_2\)O\(_3\)) can react with the alkali NaOH to from sodium aluminium hydroxide (NaAl(OH)\(_3\)) 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.

<table>
<thead>
<tr>
<th>Test for...</th>
<th>By...</th>
<th>Positive result</th>
<th>The reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Copper (II), Cu(^{2+})</td>
<td>Add sodium hydroxide followed by ammonia solution</td>
<td>Blue precipitate that dissolves when ammonia added</td>
<td>Insoluble copper (II) hydroxide formed: (\text{Cu}^{2+}<em>{(aq)} + 2\text{NaOH}</em>{(aq)} \rightarrow \text{Cu(OH)}<em>{2(s)} + 2\text{Na}^+</em>{(aq)}) When ammonia is added a soluble complex forms so the precipitate dissolves.</td>
</tr>
<tr>
<td>Iron (II), Fe(^{2+})</td>
<td>Add sodium hydroxide followed by ammonia solution.</td>
<td>Green precipitate insoluble in ammonia</td>
<td>Insoluble iron (II) hydroxide formed: (\text{Fe}^{2+}<em>{(aq)} + 2\text{NaOH}</em>{(aq)} \rightarrow \text{Fe(OH)}<em>{2(s)} + 2\text{Na}^+</em>{(aq)}) Ammonia does not react with the iron (II) hydroxide so it does not dissolve.</td>
</tr>
<tr>
<td>Iron (III), Fe(^{3+})</td>
<td>Add sodium hydroxide followed by ammonia solution.</td>
<td>Brown precipitate insoluble in ammonia</td>
<td>Insoluble iron (III) hydroxide formed: (\text{Fe}^{3+}<em>{(aq)} + 3\text{NaOH}</em>{(aq)} \rightarrow \text{Fe(OH)}<em>{3(s)} + 3\text{Na}^+</em>{(aq)}) Ammonia does not react with the iron (III) hydroxide so it does not dissolve.</td>
</tr>
<tr>
<td>Zinc, Zn(^{2+})</td>
<td>Add sodium hydroxide followed by ammonia solution or more sodium hydroxide</td>
<td>White precipitate soluble in both ammonia or more sodium hydroxide</td>
<td>Insoluble zinc hydroxide formed: (\text{Zn}^{2+}<em>{(aq)} + 2\text{NaOH}</em>{(aq)} \rightarrow \text{Zn(OH)}<em>{2(s)} + 2\text{Na}^+</em>{(aq)}) Both ammonia and sodium hydroxide react with the zinc hydroxide to form a soluble complex.</td>
</tr>
</tbody>
</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.

---

### 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
- **Some metals are also:**
  - Malleable – can be beaten into shape
  - Strong

**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.

---

### 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 \( \rightarrow \) metal hydroxide + hydrogen
- \( Li \ + \ H_2O \rightarrow LiOH + H_2 \)

Lithium reacts the slowest, Na reacts faster, K reacts fastest 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,
- \( Br_2 + Cl^- \rightarrow \) no reaction

As \( Br \) is less reactive than \( Cl \).

### 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 (1/7\(^{th}\) that of air) so is used to make airships and blimps float.

---

### 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

---

### GROUPS

#### 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.

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- **Period number** = number of shells
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For example: Chlorine is in Period 3 and Group VII so it has 3 electron shells and 7 electrons in the outer shell.
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 pattern 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**
Zn + Copper sulphate \(\rightarrow\) Zinc sulphate + copper
\[\text{Zn} + \text{Cu}^{2+} + \text{SO}_4^{2-} \rightarrow \text{Zn}^{2+} + \text{SO}_4^{2-} + \text{Cu}^{2+}\]
This happens because Zn is more reactive than Cu so is able to reduce it.

**Eg 2. Reaction with metal oxides**
Iron oxide + aluminium \(\rightarrow\) aluminium oxide + iron
This happens since Al is more reactive than Fe so is able to reduce it.

**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 they 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.

**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

**WATER, H₂O**
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:

<table>
<thead>
<tr>
<th>Component</th>
<th>Percentage</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nitrogen</td>
<td>79%</td>
</tr>
<tr>
<td>Oxygen</td>
<td>21%</td>
</tr>
<tr>
<td>Other</td>
<td>1%</td>
</tr>
</tbody>
</table>

The ‘other’ is mostly argon with CO₂, 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 CO₂ is able to absorb the infrared radiation (heat) radiated by the ground when the sun heats it up (the greenhouse effect). More CO₂ 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.

**NITROGEN AND AMMONIA**
Ammonia (NH₃) is a smelly gas. One way to produce it is to react ammonium (NH₄⁺) salts with an alkali (OH⁻) eg:
\[ \text{NH}_4\text{Cl} + \text{NaOH} \rightarrow \text{NH}_3 + \text{H}_2\text{O} + \text{NaCl} \]

Ammonia is vital to produce the nitrates used in fertilisers and explosives. It is produced by the Haber process:
\[ \text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3 \]

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°C) with an iron oxide catalyst.

High pressure (~200 times atmospheric pressure) is used to increase the proportion of NH₃ formed.

The nitrogen comes from the air and hydrogen comes from reacting methane (CH₄) gas with steam.

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

**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 (NO₃⁻ containing) salts; phosphorous – in the form of phosphate (PO₄³⁻ 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.

**CARBON DIOXIDE, CO₂**
There are many ways to produce CO₂ including:

- **Burning carbon-containing fuels:**
  \[ \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \]

- **As a by product of respiration** in living cells:
  \[ \text{C}_6\text{H}_{12}\text{O}_6 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \]

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

Rusting

- Many of man’s activities pollute the air. Pollutants include:
  - **Carbon monoxide, CO**
    - Formed when fuels burn without enough O₂.
    - CO prevents the blood from carrying oxygen leading to death by suffocation
  - **Sulphur dioxide, SO₂**
    - 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, NOₓ**
    - 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.
  - **Fertilisers**
    - Added to the water which destroys bacteria.
  - **Air and water**
    - 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.
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:

\[ H_2SO_4(\text{l}) \rightarrow 2H^+_{(\text{aq})} + SO_4^{2-}_{(\text{aq})} \]

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 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):

\[ 2SO_2 + O_2 \rightleftharpoons SO_3 \]

This reaction is reversible, so to maximise the amount of SO_3 made, they use a high temperature (425°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.
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°C, producing carbon dioxide as a by-product:

\[
\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})
\]

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:

\[
\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2
\]

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!
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.

Bubble Caps: the gaseous fractions bubble up through these until they get cool enough when they then condense.

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, CH₄) contains much less carbon and so is an environmentally better fuel.
### HYDROCARBONS

Hydrocarbons are compounds made of only hydrogen and carbon atoms. Hydrocarbons – for example methane (CH₄) – burn very well producing only carbon dioxide (CO₂) and water (H₂O):

\[
\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}
\]

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.

### MACROMOLECULES

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

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

**Note:**

• The with the alkenes for each carbon there are 2 H (CₙH₂n); with the alkanes, for each C there are 2 H plus 2 extra (CₙH₂n₊₂).

• Any combination of alkene and alkane can be made, including straight and branched chains, so long as the numbers of atoms balance.

### 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**, **vapourised** and passed over a **ceramic catalyst** produce a shorter alkane and an alkene.

**Eg. 1:** C₈H₁₈ → C₄H₁₀ + C₄H₈

**Eg. 2:** C₁₀H₂₂ → C₇H₁₆ + C₃H₆

### 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:

\[
\text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O}
\]

### ADDITION REACTIONS

- **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:
    \[
    \text{C}_2\text{H}_4 + \text{Br}_2 \rightarrow \text{C}_2\text{H}_4\text{Br}_2
    \]

- **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.
    \[
    \text{C}_2\text{H}_4(+g) + \text{H}_2\text{O}(g) \rightarrow \text{C}_2\text{H}_5\text{OH}(g)
    \]

- **Addition reaction of alkenes with hydrogen:**
  - Alkenes reacts with hydrogen in the presence of a nickel catalyst to make alkanes.
    \[
    \text{C}_2\text{H}_4(+g) + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g})
    \]

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.
The periodic table of the elements

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

*58-71 Lanthanoid series
190-103 Actinoid series

The volume of one mole of any gas is 24 dm³ at room temperature and pressure (r.t.p.).