AQA GCSE Atomic structure and periodic table part 1

The development of the model of the atom

- **Pre 1900**: Tiny solid spheres that could not be divided. Before the discovery of the electron, John Dalton said the solid sphere made up the different elements.
- **1897 ‘plum pudding’**: A ball of positive charge with negative electrons embedded in it. JJ Thompson’s experiments showed that an atom must contain small negative charges (discovery of electrons).
- **1909 nuclear model**: Positively charge nucleus at the centre surrounded negative electrons. Ernest Rutherford’s alpha particle scattering experiment showed that the mass was concentrated at the centre of the atom.
- **1913 Bohr model**: Electrons orbit the nucleus at specific distances. Niels Bohr proposed that electrons orbited in fixed shells; this was supported by experimental observations.

**Atoms, elements and compounds**

- **Atom**: The smallest part of an element that can exist. Have a radius of around 0.1 nanometres and have no charge (0).
- **Element**: Contains only one type of atom. Around 100 different elements each one is represented by a symbol e.g. O, Na, Br.
- **Compound**: Two or more elements chemically combined. Compounds can only be separated into elements by chemical reactions.

**Central nucleus**
- Contains protons and neutrons

**Electron shells**
- Contains electrons

<table>
<thead>
<tr>
<th>Name of Particle</th>
<th>Relative Charge</th>
<th>Relative Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>+1</td>
<td>1</td>
</tr>
<tr>
<td>Neutron</td>
<td>0</td>
<td>1</td>
</tr>
<tr>
<td>Electron</td>
<td>-1</td>
<td>Very small</td>
</tr>
</tbody>
</table>

**Electrons**

- **Electronic structures**

<table>
<thead>
<tr>
<th>Electronic shell</th>
<th>Max number of electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>2</td>
</tr>
<tr>
<td>2</td>
<td>8</td>
</tr>
<tr>
<td>3</td>
<td>8</td>
</tr>
<tr>
<td>4</td>
<td>2</td>
</tr>
</tbody>
</table>

**Mixtures**

- Two or more elements or compounds not chemically combined together. Can be separated by physical processes.

**Chemical equations**

- *Show chemical reactions - need reactant(s) and product(s) energy always involves and energy change*
- *Law of conservation of mass states the total mass of products = the total mass of reactants.*

<table>
<thead>
<tr>
<th>Method</th>
<th>Description</th>
<th>Example</th>
<th>Law of conservation of mass states the total mass of products = the total mass of reactants.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Filtration</td>
<td>Separating an insoluble solid from a liquid</td>
<td>To get sand from a mixture of sand, salt and water.</td>
<td></td>
</tr>
<tr>
<td>Crystallisation</td>
<td>To separate a solid from a solution</td>
<td>To obtain pure crystals of sodium chloride from salt water.</td>
<td></td>
</tr>
<tr>
<td>Simple distillation</td>
<td>To separate a solvent from a solution</td>
<td>To get pure water from salt water.</td>
<td></td>
</tr>
<tr>
<td>Fractional distillation</td>
<td>Separating a mixture of liquids each with different boiling points</td>
<td>To separate the different compounds in crude oil.</td>
<td></td>
</tr>
<tr>
<td>Chromatography</td>
<td>Separating substances that move at different rates through a medium</td>
<td>To separate out the dyes in food colouring.</td>
<td></td>
</tr>
</tbody>
</table>

**Word equations**

- Uses words to show reaction
  - **reactants → products**
  - **magnesium + oxygen → magnesium oxide**
- **Does not show what is happening to the atoms or the number of atoms.**

**Symbol equations**

- Uses symbols to show reaction
  - **reactants → products**
  - **2Mg + O₂ → 2MgO**
- Shows the number of atoms and molecules in the reaction, these need to be balanced.

**Isotopes**

- **Atoms of the same element with the same number of protons and different numbers of neutrons**
  - **35Cl (75%) and 37Cl (25%)**
    - Relative abundance = (% isotope 1 x mass isotope 1) + (% isotope 2 x mass isotope 2) / 100
    - e.g. (25 x 37) + (75 x 35) / 100 = 35.5
### AQA GCSE Atomic structure and periodic table part 2

**Metals and non-metals**

<table>
<thead>
<tr>
<th>Metals</th>
<th>To the left of the Periodic table</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Form positive ions. Conductors, high melting and boiling points, ductile, malleable.</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Non-metals</th>
<th>To the right of the Periodic table</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Form negative ions. Insulators, low melting and boiling points.</td>
</tr>
</tbody>
</table>

**Halogen**

- Have seven electrons in their outer shell. Form -1 ions.
- Melting and boiling points increase down the group (gas → liquid → solid)
- Reactivity decreases down the group

**With metals**

- Forms a metal halide
  - e.g. Sodium + chlorine → sodium chloride
  - Metal + halogen → metal halide

**With hydrogen**

- Forms a hydrogen halide
  - Hydrogen + halogen → hydrogen halide
e.g. Cl₂ + H₂ → 2HCl

**With aqueous solution of a halide salt**

- More reactive halogen will displace the less reactive halogen from the salt
  - e.g. Cl₂ + 2KBr → 2KCl + Br₂

**Boiling points increase down the group**

**Boiling points increase down the group**

**Unreactive, do not form molecules**

This is due to having full outer shells of electrons.

**Noble gases**

- Very reactive with oxygen, water and chlorine
- Reactivity increases down the group

**Elements arranged in order of atomic number**

- Early periodic tables were incomplete, some elements were placed in inappropriate groups if the strict order of atomic weights was followed.

**Early periodic tables**

- Before discovery of protons, neutrons and electrons
- Only have one electron in their outer shell. Form +1 ions.

**Development of the Periodic table**

- Mendeleev predicted elements that hadn’t been discovered yet
- Only have one electron in their outer shell. Form +1 ions.

**Elements in the same group have the same number of outer shell electrons and elements in the same period (row) have the same number of electron shells.**

**Elements arranged in order of atomic weight**

- Elements with similar properties are in columns called groups
- Elements with properties predicted by Mendeleev were discovered and filled in the gaps. Knowledge of isotopes explained why order based on atomic weights was not always correct.

**With oxygen**

- Forms a metal oxide
  - Metal + oxygen → metal oxide
  - e.g. 4Na + O₂ → 2Na₂O

**With water**

- Forms a metal hydroxide and hydrogen
  - Metal + water → metal hydroxide + hydrogen
  - e.g. 2Na + 2H₂O → 2NaOH + H₂

**With chlorine**

- Forms a metal chloride
  - Metal + chlorine → metal chloride
  - e.g. 2Na + Cl₂ → 2NaCl

**Typical properties**

- Compared to group 1:
  - Less reactive
  - Harder
  - Higher melting points

- Typical properties:
  - Many have different ion possibilities with different charges
  - Used as catalysts
  - Form coloured compounds

- 

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- Cu²⁺ is blue
- Ni²⁺ is pale green, used in the manufacture of margarine
- Fe²⁺ is green, used in the Haber process
- Fe³⁺ is reddish-brown
- Mn²⁺ is pale pink
AQA BONDING, STRUCTURE AND THE PROPERTIES OF MATTER 1

**Ionic bonding**
- Electrons are transferred so that all atoms have a noble gas configuration (full outer shells).
- Metal atoms lose electrons and become positively charged ions
- Non metals atoms gain electrons to become negatively charged ions

**Properties of ionic compounds**
- Melting and freezing happen at melting point, boiling and condensing happen at boiling point.
- The amount of energy needed for a state change depends on the strength of forces between particles in the substance.

**Properties of metallic elements and alloys**
- Pure metals can be bent and shaped
- Harder than pure metals because atoms of different sizes disrupt the layers so they cannot slide over each other.

**Chemical bonds**
- Ionic bonds
- Covalent bonds
- Metallic bonds

**Melting and boiling points**
- High melting and boiling points
- Large amounts of energy needed to break the bonds.
- Do not conduct electricity when solid
- Ions are held in a fixed position in the lattice and cannot move.
- Do conduct electricity when molten or dissolved
- Lattice breaks apart and the ions are free to move.

**Metallic bonding**
- Electronic configuration of outer shells
- -2 ions
- -1 ions
- Group 1 metals form +1 ions
- Group 2 metals form +2 ions
- Group 6 non metals form -2 ions
- Group 7 non metals form -1 ions

**Dot and cross diagram**
- Na⁺, Cl⁻
- (2, 8, 1)
- (2, 8, 7)
- (2, 8, 8)

**Giant structure**
- Na⁺, Cl⁻
- Liquid
- Gas

**Metals as conductors**
- Good conductors of electricity
- Delocalised electrons carry electrical charge through the metal.
- Energy is transferred by the delocalised electrons.

**Limitations of simple model**
- There are no forces in the model
- All particles are shown as spheres
- Spheres are solid

**Pure metal**
- Mixture of two or more elements at least one of which is a metal

**Alloy**
- Electrons in the outer shell of metal atoms are delocalised and free to move through the whole structure. This sharing of electrons leads to strong metallic bonds.
**Conservation of mass**

No atoms are lost or made during a chemical reaction.

Mass of the products equals the mass of the reactants.

Mass of the products equals the sum of the masses of the reactants.

The reactant that is completely used up is the limiting reactant.

Chemical measurements can determine whether the mean value falls within the range of uncertainty of the result.

Concentration of solutions is measured in mass per given volume of solution (g/dm$^3$).

Concentration of solutions can be calculated from the masses of reactants and products.

**Balanced symbol equations**

Chemical amounts are measured in moles (mol).

One mole of any substance will contain the same number of particles, atoms, molecules or ions.

Subscript numbers show the number of atoms of the element to its left.

Normal script numbers show the number of molecules.

Mass changes when a reactant or product is a gas.

Mass appears to increase during a reaction.

One of the reactants is a gas.

Mass appears to decrease during a reaction.

One of the products is a gas and has escaped.

One mole of H$_2$O = 18g (1 + 1 + 16)

One mole of Mg = 24g

Mg + 2HCl $\rightarrow$ MgCl$_2$ + H$_2$

One mole of magnesium reacts with two moles of hydrochloric acid to make one mole of magnesium chloride and one mole of hydrogen.

If you have a 60g of Mg, what mass of HCl do you need to convert it to MgCl$_2$?

$A_r$ : Mg = 24 so mass of 1 mole of Mg = 24g

$M_r$ : HCl (1 + 35.5) so mass of 1 mole of HCl = 36.5g

So 60g of Mg is 60/24 = 2.5 moles

Balanced symbol equation tells us that for every one mole of Mg, you need two moles of HCl to react with it.

So you need 2.5x2 = 5 moles of HCl

You will need 5 x 36.5g of HCl = 182.5g
A measure of the amount of starting materials that end up as useful products. Atom economy is important or sustainable development and economic reasons.

Calculate the atom economy for making hydrogen by reacting zinc with hydrochloric acid:

\[ \text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2 \]

- \( M_r \text{ of } \text{H}_2 = 1 + 1 = 2 \)
- \( M_r \text{ of } \text{Zn} + 2\text{HCl} = 65 + 1 + 1 + 35.5 + 35.5 = 138 \)

Atom economy = \( \frac{2}{138} \times 100 = \frac{2}{138} \times 100 = 1.45\% \)

This method is unlikely to be chosen as it has a low atom economy.

Concentration of a solution is the amount of solute per volume of solution:

\[ \text{Concentration} = \frac{\text{amount (mol)}}{\text{volume (dm}^3)} \]

If the volumes of two solutions that react completely are known and the concentrations of one solution is known, the concentration of the other solution can be calculated.

2NaOH(aq) + H₂SO₄(aq) → Na₂SO₄(aq) + 2H₂O(l)

It takes 12.20cm³ of sulfuric acid to neutralise 24.00cm³ of sodium hydroxide solution, which has a concentration of 0.50mol/dm³.

Calculate the concentration of the sulfuric acid in mol/dm³:

\[ 0.5 \text{ mol/dm}^3 \times \left( \frac{24}{1000} \right) \text{ dm}^3 = 0.012 \text{ mol of NaOH} \]

The equation shows that 2 mol of NaOH reacts with 1 mol of H₂SO₄, so the number of moles in 12.20cm³ of sulfuric acid is \( \left( \frac{0.012}{2} \right) = 0.006 \text{ mol of sulfuric acid} \)

Calculate the concentration of sulfuric acid in mol/ dm³

\[ 0.006 \text{ mol x (1000/12.2)} \text{ dm}^3 = 0.49 \text{mol/dm}^3 \]

Use of amount of substance in relation to volumes of gases (HT only, chemistry only)

\[ \text{Use of } \text{amount of substance } \rightarrow \frac{\text{vol of gas}}{\text{at RTP}} = 24 \text{ dm}^3 \]

A piece of sodium metal is heated in chlorine gas. A maximum theoretical mass of 10g for sodium chloride was calculated, but the actual yield was only 8g.

\[ \text{Calculate the percentage yield.} \]

Percentage yield = \( \frac{8}{10} \times 100 = 80\% \)

What is the volume of 11.6 g of butane (C₄H₁₀) gas at RTP?

\[ M_r : \text{(4 x 12) + (10 x 1)} = 58 \]

\[ 11.6/58 = 0.20 \text{ mol} \]

\[ \text{Volume} = 0.20 \times 24 = 4.8 \text{ dm}^3 \]

6g of a hydrocarbon gas had a volume of 4.8 dm³. Calculate its molecular mass.

\[ 1 \text{ mole} = 24 \text{ dm}^3, \text{ so } 4.8/24 = 0.2 \text{ mol} \]

\[ M_r = 6 / 0.2 = 30 \]

If 6g = 0.2 mol, 1 mol equals 30 g
Better hope – brighter future
The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved.

**At the negative electrode**
- Metal will be produced on the electrode if it is less reactive than hydrogen. Hydrogen will be produced if the metal is more reactive than hydrogen.

**At the positive electrode**
- Oxygen is formed at positive electrode. If you have a halide ion (\(\text{Cl}^-, \text{I}^-, \text{Br}^-\)) then you will get chlorine, bromine or iodine formed at that electrode.

**Electrolysis of aqueous solutions**

- **Strong acids**: Completely ionised in aqueous solutions e.g. hydrochloric, nitric and sulfuric acids.
- **Weak acids**: Only partially ionised in aqueous solutions e.g. ethanoic acid, citric acid.
- **Hydrogen ion concentration**: As the pH decreases by one unit (becoming a stronger acid), the hydrogen ion concentration increases by a factor of 10.

**Soluble salts**
- Soluble salts can be made from reactants of acids with solid insoluble substances (e.g. metals, metal oxides, hydroxides and carbonates).

**Production of soluble salts**
- Add the solid to the acid until no more dissolves. Filter off excess solid and then crystallise to produce solid salts.

**Electrolysis**
- When an ionic compound is melted or dissolved in water, the ions are free to move. These are then able to conduct electricity and are called electrolytes. Passing an electric current through electrolytes causes the ions to move to the electrodes.

**Electrode**
- **Anode**
  - The positive electrode is called the anode.
- **Cathode**
  - The negative electrode is called the cathode.

**Where do the ions go?**
- **Cations** (positive ions) and **anions** (negative ions) move to the electrodes.

**Calculating the chemical quantities in titrations involving concentrations in mol/dm³ and in g/dm³**

**HT ONLY**:  
\[2\text{NaOH(aq)} + \text{H}_2\text{SO}_4(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O(l)}\]

It takes 12.20 cm³ of sulfuric acid to neutralise 24.00 cm³ of sodium hydroxide solution, which has a concentration of 0.50 mol/dm³.

**Calculating the concentration of the sulfuric acid in g/dm³**

\[0.5 \text{ mol/dm}^3 \times (24/1000) \text{ dm}^3 = 0.012 \text{ mol of NaOH}\]

**Calculating the concentration of sulfuric acid in mol/dm³**

\[0.006 \text{ mol} \times (1000/12.2) \text{ dm}^3 = 0.49 \text{ mol/dm}^3\]

**Calculating the concentration of sulfuric acid in g/dm³**

\[\frac{2 \times (2x1) + 32 + (4x16)}{98} = 98g\]

\[0.49 \times 98g = 48.2g/dm^3\]
AQA GCSE Energy changes

**Endothermic**
- Energy is taken in from the surroundings so the temperature of the surroundings decreases
- Thermal decomposition
- Sports injury packs

**Exothermic**
- Energy is transferred to the surroundings so the temperature of the surroundings increases
- Combustion
- Hand warmers
- Neutralisation

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**Types of reaction**

- **Fuel cells (Chemistry only)**
- **Cells and batteries (Chemistry only)**

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**Reaction profiles**

- Show the overall energy change of a reaction

**Breaking bonds in reactants**
- **Endothermic process**

**Making bonds in products**
- **Exothermic process**

---

**The energy change of reactions (HT only)**

- **AQA GCSE Energy changes**

**Endothermic**
- Energy released making new bonds is greater than the energy taken in breaking existing bonds.

**Exothermic**
- Energy needed to break existing bonds is greater than the energy released making new bonds.

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**Overall energy change of a reaction**

- **Endothermic**
- **Exothermic**

---

**Bond energy calculation**

- Calculate the overall energy change for the forward reaction: \( \text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3 \)
- Bond energies (in kJ/mol): H-H 436, H-N 391, N≡N 945
- Bond breaking: 945 + (3 x 436) = 945 + 1308 = 2253 kJ/mol
- Bond making: 6 x 391 = 2346 kJ/mol
- Overall energy change = 2253 - 2346 = -93kJ/mol
- Therefore reaction is exothermic overall.

---

**Endothermic**
- Products are at a higher energy level than the reactants. As the reactants form products, energy is transferred from the surroundings to the reaction mixture. The temperature of the surroundings decreases because energy is taken in during the reaction.

**Exothermic**
- Products are at a lower energy level than the reactants. When the reactants form products, energy is transferred to the surroundings. The temperature of the surroundings increases because energy is released during the reaction.

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**Activation energy**

- Chemical reactions only happen when particles collide with sufficient energy

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**Simple cell**
- Make a simple cell by connecting two different metals in contact with an electrolyte
- Increase the voltage by increasing the reactivity difference between the two metals.

**Batteries**
- Consist of two or more cells connected together in series to provide a greater voltage.
- Non-rechargeable cells: Stop when one of the reactants has been used up
- Alkaline batteries
- Rechargeable cells: Can be recharged because the chemical reactions are reversed when an external electrical current is supplied
- Rechargeable batteries

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**Hydrogen fuel cells**
- Word equation: hydrogen + oxygen \( \rightarrow \) water
- Symbol equation: \( 2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \)
- Advantages:
  - No pollutants produced
  - Can be a range of sizes
- Disadvantages:
  - Hydrogen is highly flammable
  - Hydrogen is difficult to store

---

**Bond making**
- Chemical reactions only happen when particles collide with sufficient energy
**Factors affecting the rate of reaction**

<table>
<thead>
<tr>
<th>Quantity</th>
<th>Unit</th>
</tr>
</thead>
<tbody>
<tr>
<td>Temperature</td>
<td>The higher the temperature, the quicker the rate of reaction.</td>
</tr>
<tr>
<td>Concentration</td>
<td>The higher the concentration, the quicker the rate of reaction.</td>
</tr>
<tr>
<td>Surface area</td>
<td>The larger the surface area of a reactant solid, the quicker the rate of reaction.</td>
</tr>
<tr>
<td>Pressure (of gases)</td>
<td>When gases react, the higher the pressure upon them, the quicker the rate of reaction.</td>
</tr>
</tbody>
</table>

**Calculating rates of reactions**

Rate of reaction

\[
\text{Rate} = \frac{\text{quantity of reactant used}}{\text{time taken}}
\]

\[
\text{Rate} = \frac{\text{quantity of product formed}}{\text{time taken}}
\]

**Collision theory and activation energy**

**Activation energy**

This is the minimum amount of energy colliding particles in a reaction need in order to react.

Increasing the temperature increases the frequency of collisions and makes the collisions more energetic, therefore increasing the rate of reaction.

**Factors affecting rates**

**Collision theory and activation energy**

Catalysts

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

Enzymes

These are biological catalysts.

**How do they work?**

Catalysts provide a different reaction pathway where reactants do not require as much energy to react when they collide.

**Reversible reactions and dynamic equilibrium**

**Reversible reactions**

In some chemical reactions, the products can react again to re-form the reactants.

\[
A + B \rightleftharpoons C + D
\]

**Changing conditions and equilibrium (HT)**

The relative amounts of reactants and products at equilibrium depend on the conditions of the reaction.

**Equilibrium**

The direction of reversible reactions can be changed by changing conditions:

- Heat: \[A + B \rightleftharpoons C + D\]
- Cool: \[A + B \rightleftharpoons C + D\]

**Equilibrium in reversible reactions**

When a reversible reaction occurs in apparatus which prevents the escape of reactants and products, equilibrium is reached when the forward and reverse reactions occur exactly at the same rate.

**Changing concentration**

If the concentration of a reactant is increased, more products will be formed.

If the concentration of a product is decreased, more reactants will react.

**Changing temperature**

If the temperature of a system at equilibrium is increased:

- Exothermic reaction = products decrease
- Endothermic reaction = products increase

**Changing pressure (gaseous reactions)**

For a gaseous system at equilibrium:

- Pressure increase = equilibrium position shifts to side of equation with smaller number of molecules.
- Pressure decrease = equilibrium position shifts to side of equation with larger number of molecules.

**Le Chatelier’s Principles**

States that when a system experiences a disturbance (change in condition), it will respond to restore a new equilibrium state.

**Representing reversible reactions**

\[
A + B \rightleftharpoons C + D
\]

**The direction**

If one direction of a reversible reaction is exothermic, the opposite direction is endothermic. The same amount of energy is transferred in each case.

For example:

- **Hydrated copper sulfate** (endothermic) \[\rightleftharpoons\] **Anhydrous copper + Water** (exothermic)
Crude oil, hydrocarbons and alkanes

Carbon compounds as fuels and feedstock

Crude oil, a finite resource

Consisting mainly of plankton that was buried in the mud, crude oil is the remains of ancient biomass.

Hydrocarbons

These make up the majority of the compounds in crude oil

Most of these hydrocarbons are called alkanes.

General formula for alkanes

For example:

- \( \text{C}_2\text{H}_6 \)
- \( \text{C}_6\text{H}_{14} \)

Crude oil, hydrocarbons

Display formula for first four alkanes

- Methane \((\text{CH}_4)\)
- Ethane \((\text{C}_2\text{H}_6)\)
- Propane \((\text{C}_3\text{H}_8)\)
- Butane \((\text{C}_4\text{H}_{10})\)

AQA GCSE Organic chemistry 1

Carbon compounds as fuels and feedstock

Cracking and alkenes

Hydrocarbon chains in crude oil come in lots of different lengths.

The boiling point of the chain depends on its length. During fractional distillation, they boil and separate at different temperatures due to this.

Using fractions

Fractions can be processed to produce fuels and feedstock for petrochemical industry

We depend on many of these fuels; petrol, diesel and kerosene.

Many useful materials are made by the petrochemical industry; solvents, lubricants and polymers.

Alkanes to alkenes

Long chain alkanes are cracked into short chain alkenes.

Alkenes

Alkenes are hydrocarbons with a double bond (some are formed during the cracking process).

Properties of alkenes

Alkenes are more reactive that alkanes and react with bromine water. Bromine water changes from orange to colourless in the presence of alkenes.

Cracking

The breaking down of long chain hydrocarbons into smaller chains

The smaller chains are more useful. Cracking can be done by various methods including catalytic cracking and steam cracking.

Catalytic cracking

The heavy fraction is heated until vapourised

After vaporisation, the vapour is passed over a hot catalyst forming smaller, more useful hydrocarbons.

Steam cracking

The heavy fraction is heated until vapourised

After vaporisation, the vapour is mixed with steam and heated to a very high temperature forming smaller, more useful hydrocarbons.

Combustion

During the complete combustion of hydrocarbons, the carbon and hydrogen in the fuels are oxidised, releasing carbon dioxide, water and energy.

Decane \( \rightarrow \) pentane + propene + ethane

\( \text{C}_{10}\text{H}_{22} \rightarrow \text{C}_5\text{H}_{12} + \text{C}_3\text{H}_6 + \text{C}_2\text{H}_4 \)

Complete combustion of methane:

\( \text{CH}_4 (g) + \text{O}_2 (g) \rightarrow \text{CO}_2 (g) + 2 \text{H}_2\text{O} (l) \)

Properties of hydrocarbons

- Boiling points
- Hydrocarbon chains in oil

Alkenes and uses as polymers

Used to produce polymers. They are also used as the starting materials of many other chemicals, such as alcohol, plastics and detergents.

Why do we crack long chains?

Without cracking, many of the long hydrocarbons would be wasted as there is not much demand for these as for the shorter chains.

Boiling point

(temperature at which liquid boils)

As the hydrocarbon chain length increases, boiling point increases.

Viscosity

(how easily it flows)

As the hydrocarbon chain length increases, viscosity increases.

Flammability

(how easily it burns)

As the hydrocarbon chain length increases, flammability decreases.
AQA GCSE Organic chemistry 2 (CHEMISTRY ONLY)

**Carboxylic acids**

**Functional group**

-COOH

For example: CH₃COOH

**Carboxylic acid reactions**

- Carboxylic acids and carbonates: These acids are neutralised by carbonates
- Carboxylic acids and water: These acids dissolve in water.
- Carboxylic acids and alcohols: The acids react with alcohols to form esters.

**Strength** (HT only)

Carboxylic acids are weak acids

Carboxylic acids only partially ionise in water. An aqueous solution of a weak acid with a high pH (but still below 7).

**Synthetic and naturally occurring polymers**

**Addition polymerisation**

- Alkenes are used to make polymers by addition polymerisation.
- Many small molecules join together to form polymers (very large molecules).
- In addition polymers, the repeating unit has the same atoms as the monomer.

**Condensation polymerisation (HT only)**

- Amino acids have two functional groups in a molecule. They react by condensation polymerisation to produce peptides.
- Deoxyribonucleic acid is a large molecule essential for life. DNA gives the genetic instructions to ensure development and functioning of living organisms and viruses.
- Most DNA molecules are two polymer chains made from four different monomers, called nucleotides. They are in the double helix formation.
- Other naturally occurring polymers include proteins, starch and cellulose and are all important for life.

**DNA and naturally occurring polymers**

**Amino acids**

- Glycine

**Dyes**

**Condensation polymerisation involves monomers with two functional groups**

- When these types of monomers react they join together and usually lose small molecules, such as water. This is why they are called condensation reactions.

**Reactions of alkenes and alcohols**

- Alkenes are hydrocarbons in the functional group C=C.
- Alkenes react with oxygen in the same way as other hydrocarbons, just with a smoky flame due to incomplete combustion.
- Alkenes also react with hydrogen, water and the halogens. The C=C bond allows for the addition of other atoms.

**Functional group**

-OH

Methanol, ethanol, propanol and butanol are the first four of the homologous series.

**Alcohol reactions**

- Alcohol and sodium: bubbling, hydrogen gas given off and salt formed.
- Alcohol and air: alcohol burns in air releasing carbon dioxide and water.
- Alcohol and water: alcohols dissolve in water to form a neutral solution.

**Fermentation**

Ethanol is produced from fermentation.

When sugar solutions are fermented using yeast, aqueous solutions of ethanol are produced. The conditions needed for this process include a moderate temperature (25 – 50⁰C), water (from sugar solution) and an absence of oxygen.

**Alcohols**

**Functional group**

-OH

**For example:** CH₃CH₂OH

**Methanol, ethanol, propanol and butanol**

**Carboxylic acids**

**Functional group**

-COOH

For example: CH₃COOH

**Carboxylic acid reactions**

- Carboxylic acids react with carbonates, water and alcohols.

- Methanoic acid, ethanoic acid, propanoic acid and butanoic acid are the first four of the homologous series.

**Alkenes**

**Hydrocarbons with a double carbon-carbon bond.**

**Structure and formula of alkenes**

- Alkenes are unsaturated because they contain two fewer hydrogen atoms than their alkane counterparts.

**General formula for alkenes**

$C_nH_{2n}$
A pure substance is a single element or compound, not mixed with any other substance.

Pure substances melt and boil at specific temperatures. Heating graphs can be used to distinguish pure substances from impure.

Chromatography

A formulation is a mixture that has been designed as a useful product.

How are formulations made?

By mixing chemicals that have a particular purpose in careful quantities.

Examples of formulations.

Fuels, cleaning agents, paints, medicines and fertilisers.

Involves a mobile phase (e.g. water or ethanol) and a stationary phase (e.g. chromatography paper).

Position solvent reaches

Mixture separated

Mixture

Solvent

Rf Values

The ratio of the distance moved by a compound to the distance moved by solvent.

\[ R_f = \frac{\text{distance moved by substance}}{\text{distance moved by solvent}} \]

This depends on the solvent used. A pure substance will produce a single spot in all solvents whereas an impure substance will produce multiple spots.

Pure substances

Chromatography

Can be used to separate mixtures and help identify substances.

AQA Chemical analysis

Identification of ions (CHEMISTRY ONLY)

Identification of common gases

Gas | Test | Positive result
--- | --- | ---
Hydrogen | Burning splint | ‘Pop’ sound.
Oxygen | Glowing splint | Re-lights the splint.
Chlorine | Litmus paper (damp) | Bleaches the paper white.
Carbon dioxide | Limewater | Goes cloudy (as a solid calcium carbonate forms).

Flame tests (chem only)

Flame emission spectroscopy

Instrumental methods

Methods that rely on machines

Can be used to identify elements and compounds. These methods are accurate, sensitive and rapid.

Instrumental methods

Sulfate ions

When in a solutions they produce a white precipitate with barium chloride solutions in the presence of hydrochloric acid.

Carbonates

React with dilute acids to form carbon dioxide.

Halide ions

When in a solution, they produce precipitates with silver nitrate solution in the presence of nitric acid.

Flame emission spectroscopy

An instrumental method used to analyse metal ions.

The sample solution is put into a flame and the light that is given out is put through a spectroscope. The output line spectrum, can be analysed to identify the metal ions in the solution. It can also be used to measure concentrations.

Carbonates, halides and sulfates (chem only)

Sodium hydroxide

Is added to solutions to identify metal ions.

White precipitates

Aluminium, calcium and magnesium ions form this with sodium hydroxide solution.

Coloured precipitates

Copper (II) = blue
Iron (II) = green
Iron (III) = brown

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Composition and evolution of the atmosphere

Algae and plants
These produced the oxygen that is now in the atmosphere, through photosynthesis.

First produced by algae 2.7 billion years ago.

Reducing carbon dioxide in the atmosphere

Formation of sedimentary rocks and fossil fuels
These are made out of the remains of biological matter, formed over millions of years.

Over the next billion years plants evolved to gradually produce more oxygen. This gradually increased to a level that enabled animals to evolve.

Greenhouse gases
Carbon dioxide, water vapour and methane
Examples of greenhouse gases that maintain temperatures on Earth in order to support life.
Radiation from the Sun enters the Earth’s atmosphere and reflects off of the Earth. Some of this radiation is re-radiated back by the atmosphere to the Earth, warming up the global temperature.

The greenhouse effect

Global climate change

Human activities and greenhouse gases
Carbon dioxide
Human activities that increase carbon dioxide levels include burning fossil fuels and deforestation.

Methane
Human activities that increase methane levels include raising livestock (for food) and using landfills (the decay of organic matter released methane).

Climate change
There is evidence to suggest that human activities will cause the Earth’s atmospheric temperature to increase and cause climate change.

Properties and effects of atmospheric pollutants

Carbon monoxide
Toxic, colourless and odourless gas. Not easily detected, can kill.

Sulfur dioxide and oxides of nitrogen
Cause respiratory problems in humans and acid rain which affects the environment.

Particulates
Cause global dimming and health problems in humans.

Effects of climate change
Rising sea levels
Extreme weather events such as severe storms
Change in amount and distribution of rainfall
Changes to distribution of wildlife species with some becoming extinct

Carbon footprints
The total amount of greenhouse gases emitted over the full life cycle of a product/event. This can be reduced by reducing emissions of carbon dioxide and methane.

Common atmospheric pollutants

Carbon dioxide
Water vapour
Methane

Reducing carbon dioxide in the atmosphere

When the oceans formed, carbon dioxide dissolved into it
This formed carbonate precipitates, forming sediments. This reduced the levels of carbon dioxide in the atmosphere.

Volcano activity
Billions of years ago there was intense volcanic activity
This released gases (mainly CO₂) that formed to early atmosphere and water vapour that condensed to form the oceans.

Composition of gases in the atmosphere

<table>
<thead>
<tr>
<th>Gas</th>
<th>Percentage</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nitrogen</td>
<td>~80%</td>
</tr>
<tr>
<td>Oxygen</td>
<td>~20%</td>
</tr>
<tr>
<td>Argon</td>
<td>0.93%</td>
</tr>
<tr>
<td>Carbon dioxide</td>
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</tr>
</tbody>
</table>

Other gases
Released from volcanic eruptions
Nitrogen was also released, gradually building up in the atmosphere. Small proportions of ammonia and methane also produced.

AQA GCSE Chemistry of the atmosphere

Proportions of gases in the atmosphere

Atmospheric pollutants from fuels

Combustion of fuels
Source of atmospheric pollutants. Most fuels may also contain some sulfur.

Gases from burning fuels
Carbon dioxide, water vapour, carbon monoxide, sulfur dioxide and oxides of nitrogen.

Particulates
Solid particles and unburned hydrocarbons released when burning fuels.
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First produced by algae 2.7 billion years ago.

Reducing carbon dioxide in the atmosphere

Algae and plants
These gradually reduced the carbon dioxide levels in the atmosphere by absorbing it for photosynthesis.

Formation of sedimentary rocks and fossil fuels

Remains of biological matter falls to the bottom of oceans. Over millions of years layers of sediment settled on top of them and the huge pressures turned them into coal, oil, natural gas and sedimentary rocks. The sedimentary rocks contain carbon dioxide from the biological matter.

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The Earth's early atmosphere

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Particulates
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Carbon dioxide + water → glucose + oxygen
6CO₂ + 6H₂O → C₆H₁₂O₆ + 6O₂
Corrosion: The destruction of materials by chemical reactions with substances in the environment. An example of this is iron rusting; iron reacts with oxygen from the air to form iron oxide (rust) water needs to be present for iron to rust.

Preventing corrosion: Coatings can be added to metals to act as a barrier. Examples of this are greasing, painting and electroplating. Aluminium has an oxide coating that protects the metal from further corrosion.

Sacrificial corrosion: When a more reactive metal is used to coat a less reactive metal. This means that the coating will react with the air and not the underlying metal. An example of this is zinc used to galvanise iron.

NPK fertilisers: These contain nitrogen, phosphorous and potassium. Formulations of various salts containing appropriate percentages of the elements.

Fertiliser examples: Potassium chloride, potassium sulfate and phosphate rock are obtained by mining. Phosphate rock needs to be treated with an acid to produce a soluble salt which is then used as a fertiliser. Ammonia can be used to manufacture ammonium salts and nitric acid.

The Haber process – conditions and equilibrium:
- The reactants side of the equation has more molecules of gas. This means that if pressure is increased, equilibrium shifts towards the production of ammonia (Le Chatelier’s principle). The pressure needs to be as high as possible.
- The forward reaction is exothermic. Decreasing temperature increases ammonia production at equilibrium. The exothermic reaction that occurs releases energy to surrounding, opposing the temperature decreases. Too low though and collisions would be too infrequent to be financially viable.

Polymers: Thermosetting (polymers that do not melt when they are heated.) Thermosoftening (polymers that melt when they are heated.)

A mixture of two elements, one of which must be a metal e.g. Bronze is an alloy of copper and tin and Brass is an alloy of copper and zinc.

Gold jewellery is usually an alloy with silver, copper and zinc. The carat of the jewellery is a measure of the amount of gold in it e.g. 18 carat is 75% gold, 24 carat is 100% gold.

Alloys of iron, carbon and other metals.
- High carbon steel is strong but brittle.
- Low carbon steel is softer and easily shaped.
- Steel containing chromium and nickel (stainless) are hard and corrosion resistant.
- Aluminium alloys are low density.

Composite materials: A mixture of materials put together for a specific purpose e.g. strength
- Soda-lime glass, made by heating sand, sodium carbonate and limestone.
- Borosilicate glass, made from sand and boron trioxide, melts at higher temperatures than soda-lime glass.
- MDF wood (woodchips, shavings, sawdust and resin)
- Concrete (cement, sand and gravel)

Ceramic materials: Made from clay
- Made by shaping wet clay and then heating in a furnace, common examples include pottery and bricks.

Polymers: Many monomers can make polymers
- These factors affect the properties of the polymer. Low density (LD) polymers and high density (HD) polymers are produced from ethene. These are formed under different conditions.
- Ammonia is used to produce fertilisers
- Nitrogen + hydrogen \( \rightarrow \) ammonia

Raw materials:
- Nitrogen from the air while hydrogen from natural gas
- Both of these gases are purified before being passed over an iron catalyst. This is completed under high temperature (about 450°C) and pressure (about 200 atmospheres).

Catalyst: Iron
- The catalyst speeds up both directions of the reaction, therefore not actually increasing the amount of valuable product.