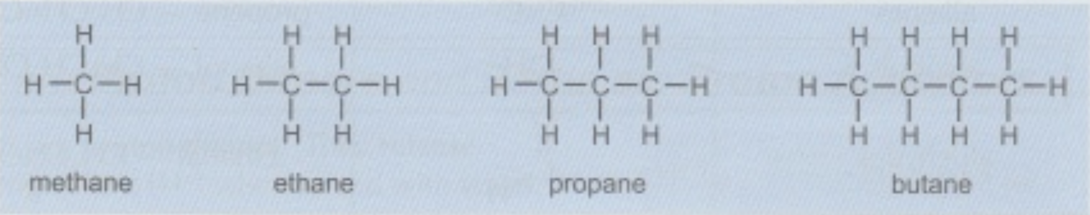
Alkanes

Structure and **Bonding** in Alkanes

1) Alkanes are **hydrocarbons** — they **only** contain hydrogen and carbon atoms.
2) Alkanes contain **two types** of bond. All of the **carbon-carbon** bonds are **single covalent bonds**. All the other bonds are **carbon-hydrogen covalent bonds** (which are always single).
3) All of the available bonds have been formed, so we call alkanes **saturated** molecules.
4) **Carbon** always forms **four** covalent bonds and **hydrogen** makes **one** covalent bond.
5) The diagrams below show the structures of the first four **straight-chain alkanes**: methane, ethane, propane and butane.



It is important to realise that these structures are only **2D representations** of the **3D molecules**. The molecules are not rigid. There is **free rotation** around a carbon-carbon single bond. This means that the carbon chains are quite **flexible** and gives the molecules the ability to **change shape**, particularly as the chain length increases.

Properties of Alkanes

The **bonds** in alkanes are very **strong** and it requires a **lot of energy** to break them. This can be used to explain some of their **properties**:

1) They are very **unreactive**.
2) They are **not** able to form **polymers**.
3) They **burn cleanly**, tending to undergo **complete combustion** to form **carbon dioxide** and **water** (see page 28). The flame is usually a faint blue colour.
   For example, the combustion of ethane:
   $2C_2H_{6(g)} + 7O_{2(g)} \rightarrow 4CO_{2(g)} + 6H_2O_{(g)}$ (+ energy)

Also:

4) **Boiling point increases** as the **length** of the carbon chain increases.
5) **Viscosity** (resistance to flow) **increases** as chain length increases.
6) **Volatility** (ease of evaporation) **decreases** as chain length increases.

These last three properties are explained by the fact that the **attractive forces** between molecules get stronger as the chain length **increases** (page 9).

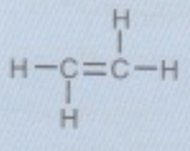
Alkanes are like the weather in the UK — completely saturated...

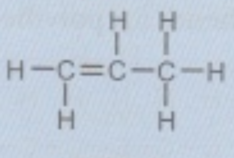
1) Draw out the structures of the next two alkanes, pentane (C_5H_{12}) and hexane (C_6H_{14}).
2) a) Write out the molecular formulae of the first four alkanes.
   b) We can work out a general formula for the alkanes of the form $C_nH_?$, where n is the number of carbon atoms. Work out, in terms of n, what should be in place of the ?.
3) Write a balanced equation for the complete combustion of propane in oxygen.

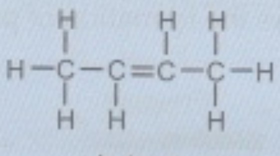
Alkenes

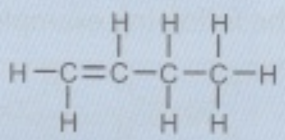
Structure and Bonding in Alkenes

1) **Alkenes** are similar to alkanes in that they are also **hydrocarbons**. The difference is in the presence of a **carbon-carbon double covalent bond** (C=C) somewhere in the carbon chain.
2) This means **not** all possible single bonds have been made — these molecules are **unsaturated**.
3) As in all compounds the carbon atoms must have **four** bonds, and hydrogen only **one**.
4) The structures of the first three alkenes (ethene, propene and butene) are shown below:


ethene


propene


butene


butene (again)

As you can see from butene, the presence of the C=C bond means that most alkenes have **more than one** possible structure.
The C=C bond can be in various different **positions** along the chain.

Molecules with the same **molecular formula** but different **structures** are called **isomers**.

The C=C bond does not allow the same **free rotation** and flexibility around itself as a C–C bond. It is a **rigid** bond. But the rest of the carbon chain is the same as in an alkane molecule, so rotation is allowed around the **single** bonds.

Properties of Alkenes

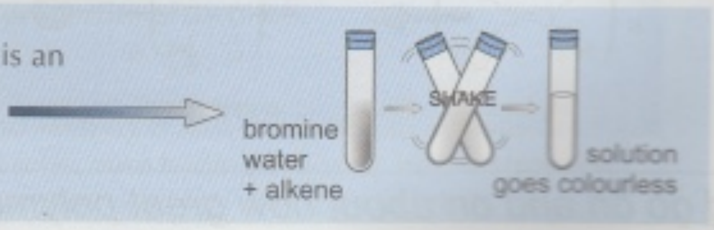
The presence of the C=C bond dictates the chemical properties of alkenes.

1) They are **reactive** compounds, undergoing many different types of chemical reaction.
2) They are used extensively to form **polymers**, e.g. poly(ethene) (see next page).
3) They **do not** burn **cleanly**, giving very **yellow** flames and lots of **soot**.

And as for alkanes, when you increase the **chain length** of an alkene:

4) The **boiling point** increases.
5) The **viscosity** increases.
6) The **volatility** decreases.

You can **test** for whether a compound is an alkene by adding it to **bromine water**. Alkenes **decolourise** bromine water, turning it from **orange** to **colourless**.



'Sleeping Butene' — coming soon to a cinema near you...

1) Draw out a structure for the next alkene: pentene (C_5H_{10}).
2) Draw out two alternative structures for hexene (C_6H_{12}).
3) Work out the general formula for the alkenes of the form C_xH_y.

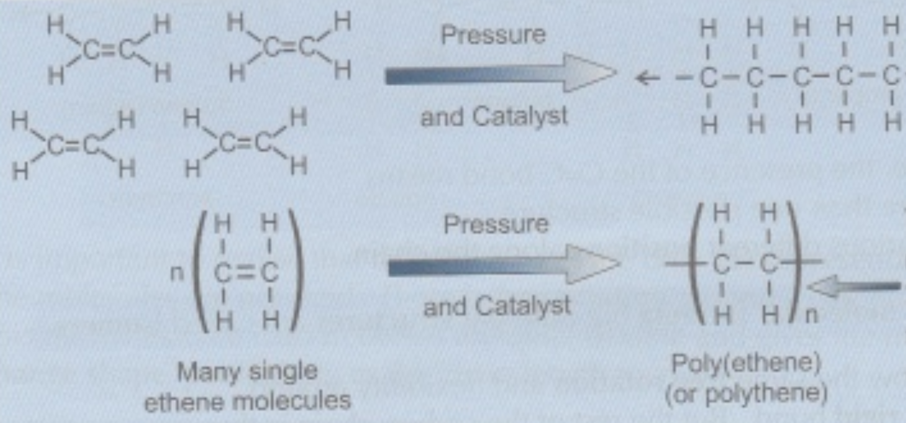
Polymerisation

*Alkenes Can Form **Polymers***

The presence of the double bond in alkene molecules means that they are capable of forming **polymers**. A polymer is a long, chain-like molecule built up from lots of **repeating units**. In this case the repeating units, called **monomers**, are alkene molecules.

Under the right conditions (these depend on the alkene and the desired properties of the polymer), many small alkenes (like ethene and propene) will **open** up their double bonds and **link together** to form these long chain polymers.

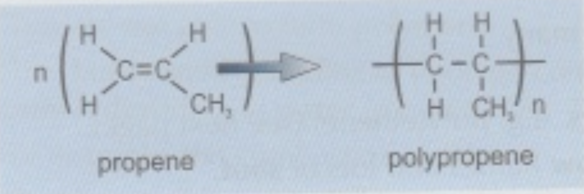
The following example shows the formation of **poly(ethene)** (or **polythene** for short):

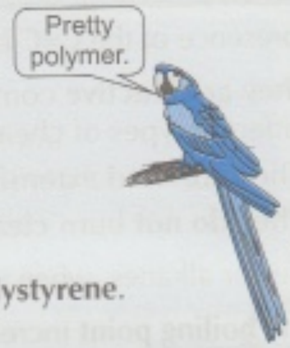


This is the standard way of representing polymers. The group of atoms inside the brackets is the **monomer** (or **repeating unit**).

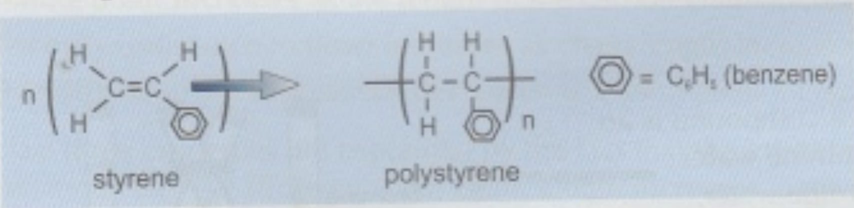
*Other Small Alkenes do a **Similar Thing***

1) **Propene** polymerises to form **polypropene**.

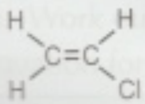
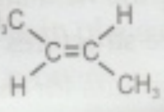




2) **Styrene**, which has a **benzene** ring in it, polymerises to form **polystyrene**. (Benzene is just a ring of six carbon atoms in which the bonding electrons are shared between all six carbons.)



I'd go on and on about how great polymers are, but it'd get repetitive...

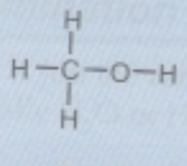
1) What property of alkenes allows them to form polymers?
2) Using the standard way of representing polymers, shown above, draw the polymers formed by the following alkenes:
   a) chloroethene:  b) this isomer of butene: 

Alcohols

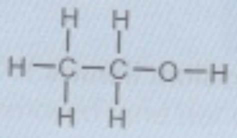
Alcohols Contain an -OH Group

The **alcohols** are a group of compounds that all contain an -OH group (an oxygen atom covalently bonded to a hydrogen atom).

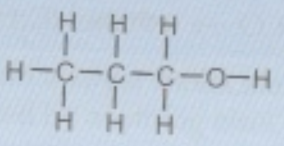
The first three alcohols are called **methanol**, **ethanol** and **propanol**. Their structures are:



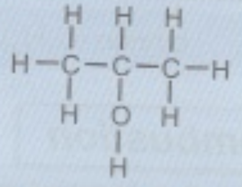
methanol



ethanol



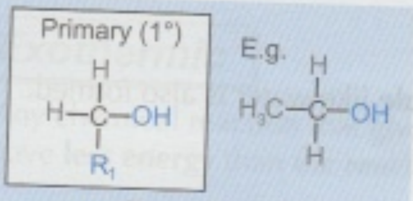
propanol



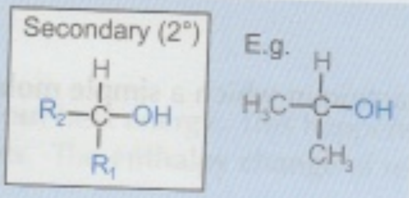
propanol again

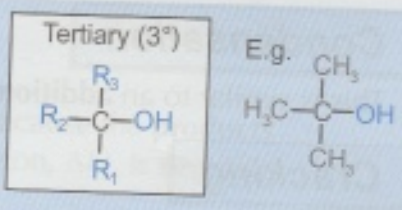
For many alcohols, the -OH group can be put in different **positions** along the chain, so they are able to form **isomers** — just like in the example with propanol above.

Alcohols can be called **primary**, **secondary** or **tertiary**. The type of alcohol depends on what other groups surround the carbon atom that the -OH group is attached to.



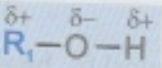
R = carbon chain





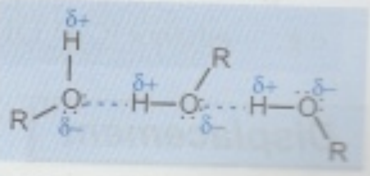
The Properties of Alcohols

Oxygen is an **electronegative** element (see page 10), so it draws the bonding electrons towards itself in the C–OH bond, meaning that alcohols are normally **polar** molecules.



$$\overset{\delta+}{R_1}-\overset{\delta-}{O}-\overset{\delta+}{H}$$

The electronegative oxygen also draws electrons away from the **hydrogen atom** in the -OH group, giving the hydrogen atom a **slightly positive charge**. This charge attracts the **lone pairs** of **electrons** on oxygen atoms in other nearby alcohol molecules, which forms a **hydrogen bond** (page 10).



Hydrogen bonds have a big effect on the **properties** of alcohols.

1) Alcohols are **soluble** in water.
2) Alcohols have **high boiling** and **melting** points compared to alkanes or alkenes of a similar size. This is because hydrogen bonds are the **strongest** type of intermolecular bond, so they need lots of energy to break.

Alcohols — always wine-ing about their rum luck. I cava beer it...

1) Draw two different isomers of butanol, C_4H_9OH.
2) Work out the general formula of alcohols, using the form C_nH_xOH.
3) Predict, with reasoning, whether ethane or ethanol will have a higher melting point.

Reaction Types

There are lots of types of **chemical reaction**. You will need to know all of them quite well. These pages give you types, explanations and examples (in alphabetical order).

Addition

This is a reaction in which atoms are **added** to an **unsaturated** bond so that the bond becomes **saturated**.

e.g. ethene (C_2H_4) + $H_2O \rightarrow$ ethanol (C_2H_5OH)

Combustion

This is the chemical reaction between a **fuel** and **oxygen**. Normally the fuel is an organic compound and the products are **carbon dioxide** and **water** — this is **complete combustion**.

e.g. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

Without enough oxygen, **incomplete combustion** occurs, producing poisonous carbon monoxide.

e.g. $C_3H_8 + 3½O_2 \rightarrow 3CO + 4H_2O$

Condensation

This is similar to an **addition** reaction in which a **simple molecule** like water is also formed.

Cracking

This is the (thermal) **decomposition** of **long-chain** hydrocarbon molecules from crude oil into **shorter-chain** alkanes and alkenes. This requires **high temperatures** and **pressures** and a **catalyst** (usually aluminium oxide), and makes hydrocarbons that are more useful.

e.g. decane ($C_{10}H_{22}$) $\rightarrow$ octane (C_8H_{18}) + ethene (C_2H_4)

Dehydration

This is the removal of **water** from a compound by **heating**.
In organic molecules it usually results in the formation of a **C=C** bond.

e.g. ethanol (C_2H_5OH) $\rightarrow$ ethene (C_2H_4) + H_2O

Displacement

This is a reaction where one element **displaces** another, **less reactive**, element from a compound. This usually takes place between **metals**, but also with **halogens**.

e.g. $2Al_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + 2Fe_{(s)}$ (the Thermite Reaction)

Disproportionation

This is a rare type of chemical reaction where an **element** in a reactant is **oxidised** and **reduced** at the same time. **Chlorine** can undergo disproportionation reactions.

e.g. $Cl_{2(aq)} + H_2O_{(l)} \rightarrow HOCl_{(aq)} + HCl_{(aq)}$

Chloric(I) acid Hydrochloric acid

The chlorine has been: **oxidised** **reduced**

Reaction Types

Electrolysis

This is a process that uses **electricity** to **break down** a compound. The reactant or reactants must be in the **liquid** state — either **molten** or in **solution**. The particles have to be able to move.
An example is the electrolysis of bauxite to obtain pure aluminium.

Elimination

This is just the **removal** of a **small molecule** from a larger molecule.
Usually H_2O or H_2 is removed (and not replaced by anything else).

e.g. propanol ($CH_3CHOHCH_3$) + sulfuric acid catalyst → propene ($CH_2{=}CHCH_3$) + water

Endothermic

Any chemical reaction that **takes in** heat energy. This means that the **reactants** will have **less energy** than the **products**. The **enthalpy change** of reaction, ΔH (see page 41), is **positive**.

Exothermic

Any chemical reaction that **gives out** heat energy. This happens because the **products** have **less energy** than the **reactants**. The **enthalpy change** of reaction, ΔH, is **negative**.

Hydrogenation

This is the **addition** of a molecule of **hydrogen** (H_2) across a **C=C** bond.
One atom attaches to each carbon.

e.g. ethene (C_2H_4) + H_2 → ethane (C_2H_6)

Neutralisation

This is the reaction between a **basic compound** and an **acid**. The products always include the **salt** of the acid, **water** and other products dependent on the acid and base.

e.g. $2KOH_{(aq)} + H_2SO_{4(aq)} \rightarrow K_2SO_{4(aq)} + 2H_2O_{(l)}$

$Na_2CO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Oxidation

There are two possible definitions for this — the best is the **loss of electrons**.
Another useful one is the **gain of oxygen**. It is the opposite of reduction.

Precipitation

A precipitate is a **solid** that is formed in a **solution** by a chemical reaction or by a change in temperature affecting solubility. Precipitates are **insoluble** in the solvent.
A precipitation reaction is simply any reaction that **produces a precipitate**.

Reaction Types

Radical (Chain) Reactions

Reactions involving radicals — an atom or compound with an **unpaired electron.** Often, one of the **products** of the reaction is also a radical which can perform further reactions. This makes the process a **chain reaction.**

Redox

This is the name for a reaction that involves both **reduction** and **oxidation** processes. It is usually used to describe reactions that just involve **electron transfer**.

reduction

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

oxidation

Reduction

There are two possible definitions for this — the best is the **gain of electrons**. The other useful one is the loss of oxygen. Important point: oxidation and reduction **ALWAYS** happen **together** — it is impossible to have one without the other.

Reversible

This is the name given to any chemical reaction that can go **forwards** and **backwards** at the **same time**. That means that the reactants will form the products, but that the products will also react (or decompose) to give the reactants.

e.g. $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

Substitution

This is simply a reaction in which an atom (or group of atoms) in a molecule is **swapped** for a different atom (or group of atoms).

Thermal Decomposition

This is where one compound **breaks down**, under **heating**, into two or more simpler compounds. A classic example is the breakdown of any carbonate compound,

e.g. $CaCO_3 + \text{heat} \rightarrow CaO + CO_2$

Cracking of hydrocarbons is also an example.

I'm in the middle of a chain reaction...

1) Write down all the different types of reaction that each of the following could be classed as.
   a) burning ethanol
   b) iron + copper sulfate → iron sulfate + copper
   c) hydrochloric acid + sodium hydroxide → sodium chloride + water + heat
   d) propene (C_3H_6) + H_2 → propane (C_3H_8)

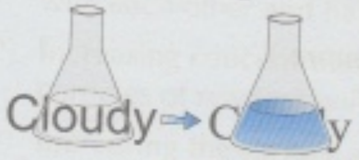
Reaction Rates

Measuring the Rate of a Reaction

The **rate** of reaction is just a measure of how **fast** a particular reaction is going. You need to know some of the ways that you can follow the rates of different reactions. They're all about measuring how fast the **reactants** are being **used up**, or measuring how fast the **products** of the reaction are **forming**.

There are lots of ways of measuring the rate of a reaction:

1) You can measure the **change in mass** that occurs during a reaction where gas is released as one of the products.

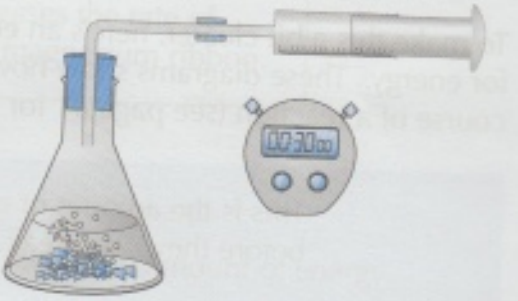


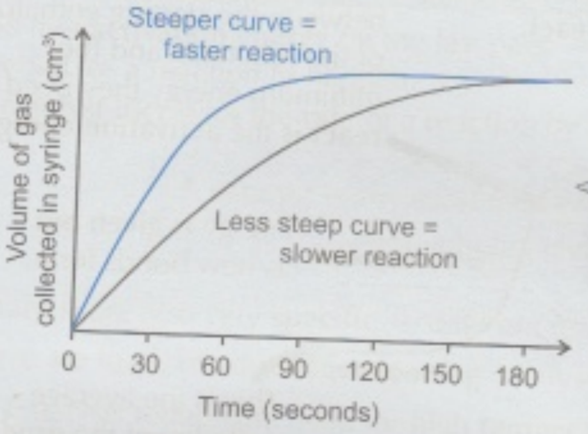
2) You can follow the **colour change** of a reaction. This includes precipitation reactions, where the solution turns cloudy as more of the product is made.
3) You can measure changes in **temperature** or **pH** that occur during the reaction.
4) You can measure the **volume of gas** produced during a reaction.

EXAMPLE: Measuring the rate of reaction between hydrochloric acid and magnesium metal.

magnesium + hydrochloric acid → magnesium chloride + hydrogen

- Use a **gas syringe** to collect the hydrogen gas that is given off during the reaction.
- Use a **stopwatch** to **time** the reaction.
- At **timed intervals**, say every 30 seconds, **record** how much hydrogen gas has been produced.





Plotting graphs lets you compare rates of reactions.

(Another way to measure the rate of this reaction would be to measure the decrease in **mass** as hydrogen gas is lost from the reaction container.)

My rate of chocolate biscuit consumption is worryingly high...

1) Describe how you could measure the rates of the following reactions:
   a) The endothermic reaction between citric acid and sodium bicarbonate to give carbon dioxide, water and a sodium salt.
   b) The precipitation reaction between sodium thiosulfate and hydrochloric acid to form a sulfur precipitate, sulfur dioxide gas, sodium chloride and water.
   c) The reaction between solid calcium carbonate and hydrochloric acid to produce calcium chloride and carbon dioxide gas.

Collision Theory

*Particles Need to **Collide** in Order to **React***

Reaction rates are explained by **collision theory**. It's based on the idea that particles in liquids and gases are always **moving around** and **colliding** with each other.
Not every **collision** results in the particles **reacting**. The following **conditions** need to be right:

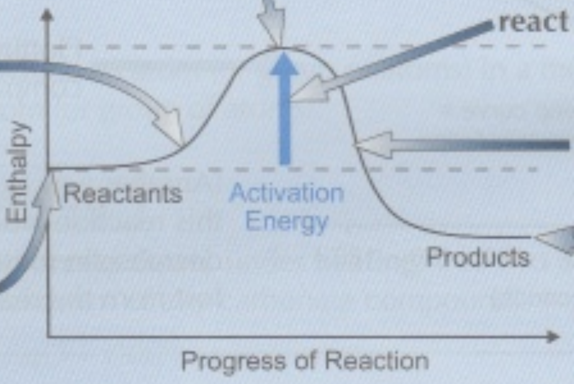
- The particles need to collide in the **right direction**. They need to be **facing** each other the right way.
- The particles need to collide with at least a certain **minimum** amount of **energy**.

Collision theory states that the **more collisions** there are, and the **more energy** these collisions have, the **more likely** particles are to react.

*Particles Need Enough **Energy** to React*

1) The **minimum amount of kinetic** (movement) **energy** particles need to react is known as the **activation energy**. This energy is used to **break the bonds** to start the reaction.
2) Reactions with **low activation energies** often happen **pretty easily**, but reactions with **high activation energies** don't — you have to give the particles **extra energy** (e.g. by **heating** them).

To make this a bit clearer, here's an **enthalpy profile diagram** — enthalpy is just a fancy word for energy. These diagrams show how the **enthalpy** of the reacting particles changes over the course of a reaction (see page 41 for more on enthalpy changes).



This is the amount of energy the particles need before they can **break bonds** and **react**.

The difference in energy between the starting enthalpy of the **reactants** and the minimum energy they need to **react** is the **activation energy**.

The particles have to **take in** energy so that their **bonds** will **break**, allowing them to react.

Energy is **given out** as new bonds form.

This is the **average enthalpy** of the reactants before they collide.

This is the average enthalpy of the **products** of the reaction.

Toast and a large cup of tea — that's my morning activation energy...

1) Two particles in a reaction vessel collide but don't react.
Give two reasons why the reaction may not have happened.
2) What is the activation energy of a reaction?
3) Draw an enthalpy profile diagram for a reaction.
On your diagram, label the reactants, products and activation energy.

Reaction Rates and Catalysts

Changing the Rate of Reaction

The rate of reaction depends on **how often** particles **collide** (see page 32) and how **likely** the collisions are to be **successful**.
More frequent **successful** collisions mean a **faster** rate of reaction.

These factors all **increase** the rate of reaction:

1) **Increasing temperature** — the particles tend to have more kinetic energy. This means that they move around faster, and so are more likely to collide with each other **and** have enough energy to react.
2) **Increasing concentration (or pressure in gases)** — this means that the particles of reactant will be closer together, so they will be more likely to collide.
3) **Increasing the surface area of a solid reactant** — this increases the number of particles of the solid reactant able to come into contact with other reactants.

EXAMPLE: Predict whether magnesium dust or magnesium ribbon will react faster with hydrochloric acid.

Magnesium dust has a larger **surface area** than magnesium ribbon.
Increasing the surface area of a solid reactant increases the rate of reaction, so **magnesium dust** will react faster than magnesium ribbon.

Catalysts Speed Up Reactions

You met **activation energy** on the last page — it's just the **minimum** amount of energy needed for a reaction to happen.
A **catalyst increases** the **rate** of a reaction by **lowering** its activation energy.

A **catalyst** is any substance which changes the **rate** of a reaction, without being **changed** or **used up** itself.

Catalysts are also very **specific** — different reactions will only be sped up by **certain catalysts**.

There are loads of advantages to using catalysts:

1) Catalysts reduce the need for **high temperatures** and **pressures** in industrial reactions, like hydrocarbon cracking (see page 28) and ethanol production (see page 36). This makes these processes **cheaper** to run.
2) Using lower temperatures also means less **energy demand**, and so lower CO_2 emissions.

Tabbys are number one on my cat list...

1) Describe two things you could do to increase the rate of a reaction between aqueous species.
2) Why does increasing the pressure increase the rate of a reaction between gases?
3) What's a catalyst?
4) Give two advantages of using a catalyst in industrial reactions.