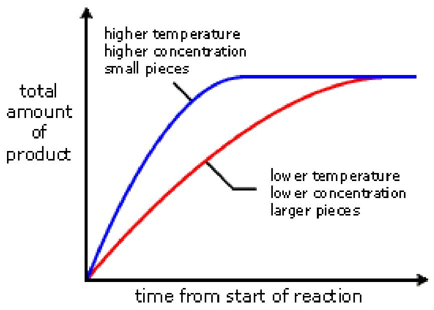
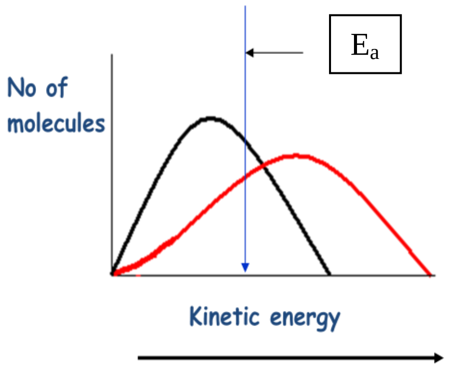
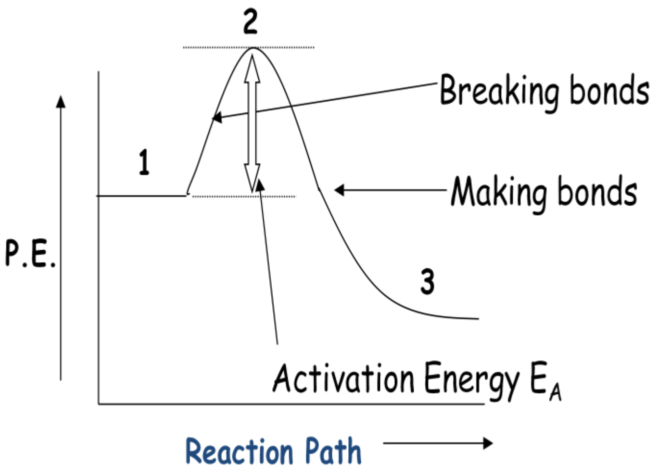
Cathkin High School
CfE Higher Chemistry

Unit 1 : Part 1

Chemical Changes & Structure
Controlling the Rate







Learning Outcomes – Controlling the Rate

Circle a face to show how much understanding you have of each statement: ☺ if you fully understand enough to do what the outcome says, 😐 if you have some understanding of the statement, and ☹ if you do not yet understand enough to do what the statement says. Once you have completed this, you will be able to tell which parts of the topic that you need to revise, by either looking at your notes again or by asking for an explanation from your teacher or classmates.

By the end of this topic I will be able to:

1. Explain the effect of temperature, concentration and particle size in terms of the energy and number of collisions (Collision Theory).	☺	😐	☹
2. State which reactions are slowest or fastest at different points using the slope of rates graphs.	☺	😐	☹
3. State that activation energy is the minimum energy required for particles to react.	☺	😐	☹
4. Draw a graph showing the effect of temperature on the kinetic energy of particles .	☺	😐	☹
5. Use activation energy on this graph to explain why higher temperatures speed up reactions.	☺	😐	☹
6. State that catalysts speed up reactions by providing an alternative reaction pathway with lower activation energy.	☺	😐	☹
7. Describe the difference between a homogeneous catalyst and a heterogeneous catalyst.	☺	😐	☹
8. Explain the adsorption, reaction and desorption stages in the action of a heterogeneous catalyst.	☺	😐	☹
9. State that catalyst poisons occupy the active site in a catalyst and prevent it working.	☺	😐	☹
10. State that enzymes are biological catalysts and give some examples of enzymes.	☺	😐	☹
11. Explain why enzymes operate at optimum temperatures and pH values.	☺	😐	☹
12. Draw potential energy diagrams for exothermic and endothermic reactions.	☺	😐	☹

13. State that enthalpy change represents the difference: $\Delta H = H_{(products)} - H_{(reactants)}$. ☺ 😐 ☹
14. State that the activated complex is an unstable arrangement of atoms formed at the maximum of the potential energy barrier, during a reaction. ☺ 😐 ☹
15. Use potential energy diagrams to illustrate the effect catalysts have on the activation energy and reaction pathway. ☺ 😐 ☹
16. Give definitions for the Standard Enthalpies of Combustion, Solution and Neutralisation. ☺ 😐 ☹
17. Use the equation for heat given out ($E_h = cm\Delta T$) from National 5 to calculate the Enthalpy of Reaction for a range of different reactions. ☺ 😐 ☹

National 4/5 Revision

The Rate of Chemical Reactions

Everyday reactions have different speeds; some are over in a fraction of a second (fast: like a gas explosion) while others can take years (slow: like the rusting of iron). Most reactions occur at rates between these two extremes (medium: like a cake baking).

Collision Theory

For a chemical reaction to occur some important things have to happen:

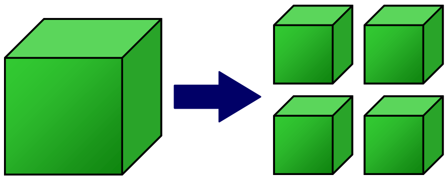
1. The reacting particles must collide together.
2. Collisions must have sufficient energy to produce a product.
3. The reacting particles must have the correct orientation.

Therefore anything that increases the number of and energy of collisions between reactant particles will speed up a reaction.

Factors Affecting the Rate of a Reaction

There are three main factors affecting the rate of a chemical reaction:

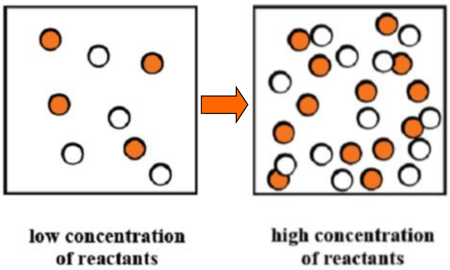
a) Particle Size:



The smaller the particles, the faster the reaction. This is because smaller particles provide more surface area for collision.

Example – Marble powder reacts faster with acid than marble chips.

b) Concentration:



The higher the concentration, the faster reaction. The higher the concentration of solutions, the more particles you have crowded into a small volume of liquid. Hence, the more likely they are to collide with each other.

Example – 2 mol/l hydrochloric acid reacts faster with magnesium ribbon than 1 mol/l hydrochloric acid.

c) Temperature:

Although a higher temperature will cause molecules to move faster, and there may be more collisions, this is not the main reason why higher temperature increases reaction rate. The main reason is that more of the collisions which occur will lead to a successful reaction. This is because at higher temperature, more particles have the **activation energy** required for a reaction to happen.

As a rough guide, the rate of reaction doubles for every increase in temperature by $10^{O}C$.

Example - Benedicts solution reacts faster with glucose solution at $50^{O}C$ than at $25^{O}C$.

Catalysts

Even when particle size is decreased and concentration and temperature are increased, many chemical reactions are still too slow. **How can the rate of these reactions be increased?** This is especially important in today's competitive market: companies are constantly trying to produce more cost effective products by increasing the rate of industrial reactions.

A catalyst is a substance which can be used to increase the rate of a chemical reaction. The 'amount' of catalyst at the end of the reaction is the same as at the start, i.e. **the catalyst is not used up in the reaction and the catalyst can be recovered chemically unchanged at the end of reaction**. Different reactions require different catalysts and not all reactions have a suitable catalyst. Catalysts provide an 'easy route' from reactants to products and lower the activation energy required for a reaction to occur.

Collision Theory and the Activated Complex

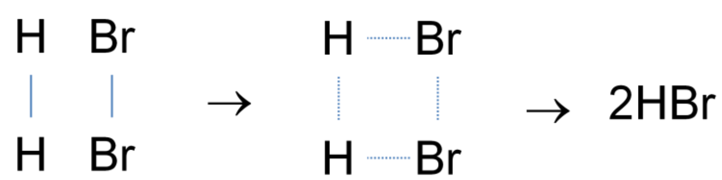
In order to react particles must collide.

A chemical reaction will only occur if the reacting particles collide with enough **kinetic energy** or **speed.** The energy is required to overcome the repulsive forces between the atoms and molecules and to start the breaking of bonds.

The **minimum kinetic energy required** for a reaction to occur is called the **activation energy (E_A).**

When the reactant particles collide with the required activation energy they form an **activated complex.** This unstable intermediate breaks down to form the products of the reaction.

E.g. **The reaction of hydrogen and bromine**

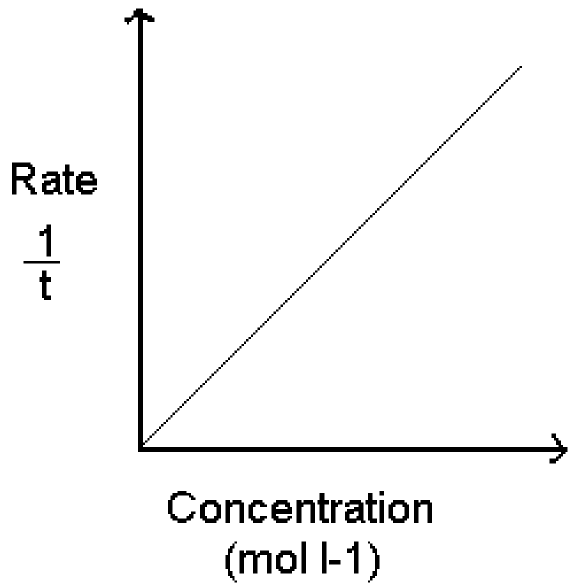


Sometimes the collisions do not result in a reaction, despite having the minimum kinetic energy.
This is thought to be because the particles have not collided with the **correct geometry** (angle) to allow the activated complex to be formed.
In the above reaction of hydrogen and bromine the particles collided side on but if they collided end on…

H-H + Br-Br → H----H-----Br----Br

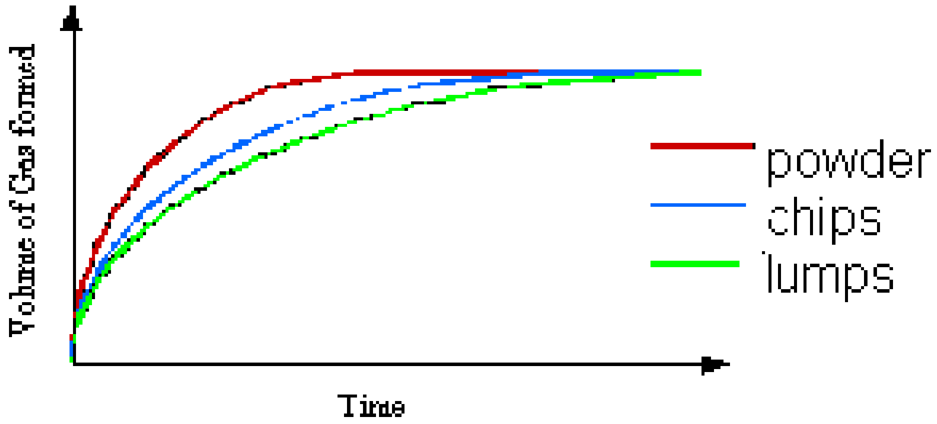
no reaction occurs as the activated complex cannot be formed if only 2 of the atoms come into contact with one another.

Collision Theory and Concentration



The straight line graph means **rate is directly proportional to the concentrations of the reactants**, i.e. double the concentration and you double the rate. This is true of many reactions.
The faster rate is due to the increased number of collisions which must occur with higher concentrations of reactants.

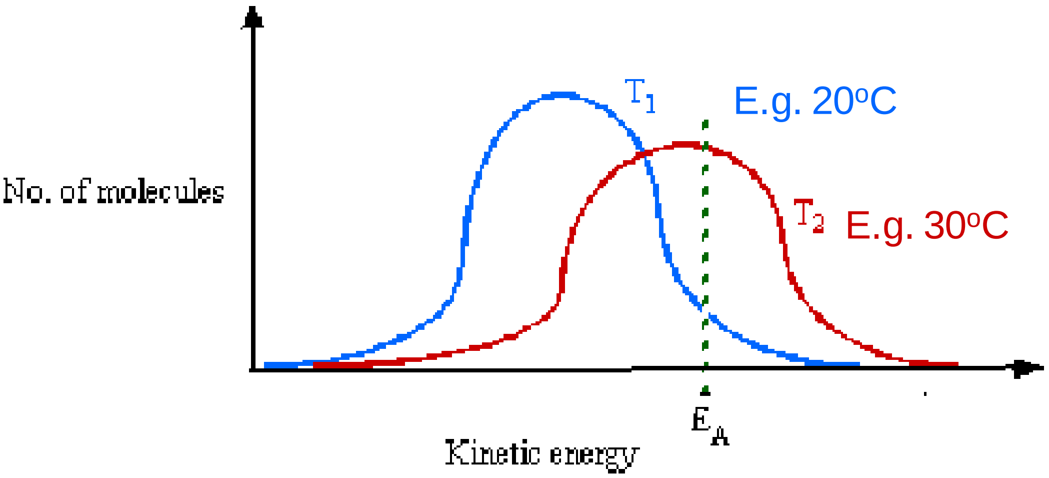
Collision Theory and Particle Size



The smaller the particle size, the faster the reaction as the total surface area is larger so more collisions will occur.

KINETIC Theory and Temperature

Temperature is a measure of the average **kinetic energy** or speed of the particles of a substance.
At any given temperature, the particles of a substance will have a range of kinetic energies and this can be shown on an **energy distribution graph.**



NB The maximum height of T_2 is always lower than T_1

The graph above shows the kinetic energy distribution of the particles of a reactant at two different temperatures.
It shows that at the **higher temperature** (T_2), many **more molecules have energies greater than the activation energy** and will be able to react when they collide.

Photochemical Reactions

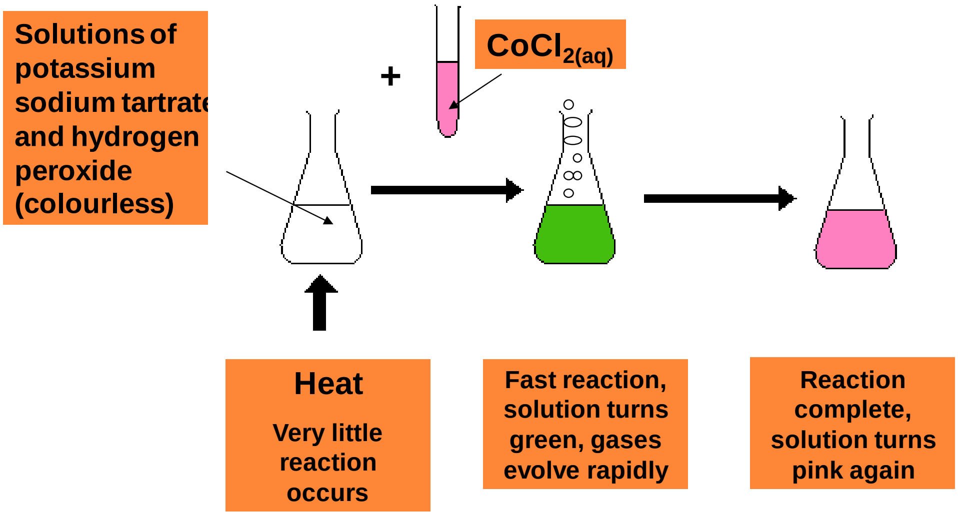
Photochemical reactions are speeded up by the presence of light.
In these reactions, the light energy helps to supply the activation energy, i.e. it increases the number of particles with energy equal to or greater than the activation energy.
Photosynthesis and the reaction of photographic film are photochemical reactions – light sustains the reaction.
Some reactions are 'set off' or initiated by light. For example when a mixture of chlorine and hydrogen gases are activated by U.V. light, a rapid and explosive reaction occurs (chain reaction):

$$H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$$

Catalysts and Reaction Rate

A catalyst is a substance which changes the speed of a chemical reaction without being permanently changed itself.
Catalysts speed up chemical reactions by providing an **alternative reaction pathway** which has a **lower activation energy**.
There are 2 main types of catalyst:
Homogeneous Catalysts
Heterogeneous Catalysts

Homogeneous catalysts are in the **same state as the reactants.**



Heterogeneous catalysts are in a **different state** to the reactants.
e.g. Decomposition of hydrogen peroxide (solution) using manganese (IV) oxide (solid) as a catalyst.

$$2H_2O_{2\,(aq)} \xrightarrow{\text{Manganese (IV) oxide (s)}} 2H_2O_{(l)} + O_{2(g)}$$

Common Catalysed Reactions

Process	Reactants	Products	Catalyst
Haber	Nitrogen & Hydrogen	Ammonia	Iron
Ostwald	Ammonia & Oxygen	Nitric acid	Platinum
Cracking	Long-chain Hydrocarbons	Short -chain hydrocarbons	Aluminium oxide or silicate
Contact	Sulphur Dioxide & oxygen	Sulphuric acid	Vanadium (V) oxide
Catalytic Converter	Carbon Monoxide	Carbon Dioxide	Platinum
Brewing	Maltose & glucose	Alcohol & Carbon dioxide	Zymase

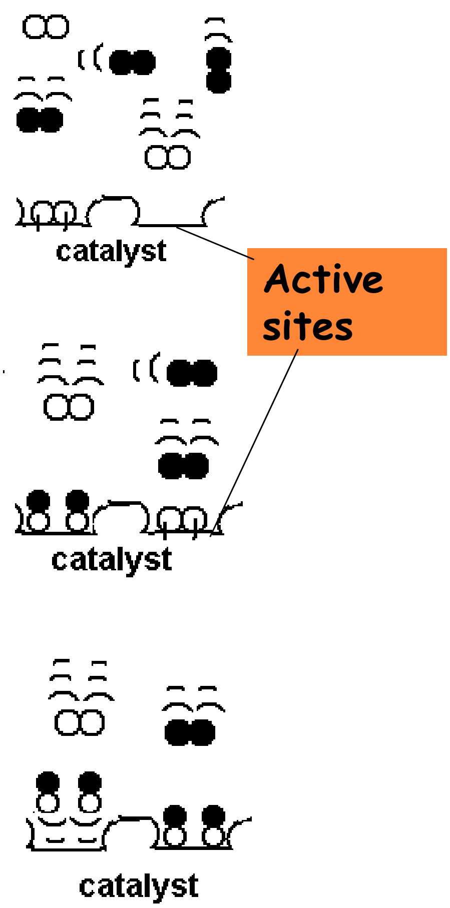
Homogeneous catalysis - hydroxide ions used in manufacture of soap from fats & oils.

How Heterogeneous Catalysts Work

This type of catalyst is called a **surface catalyst**.
It works by **adsorbing** the reacting molecules on to **active sites** and holding them with weak bonds on its surface.

This not only causes the **bonds within the molecule** to **weaken** but also helps the collision geometry. The reaction occurs on the surface with less energy needed to form the activated complex (lower activation energy).

The products are formed and leave the catalyst surface free for further reactions



Catalyst Poisoning

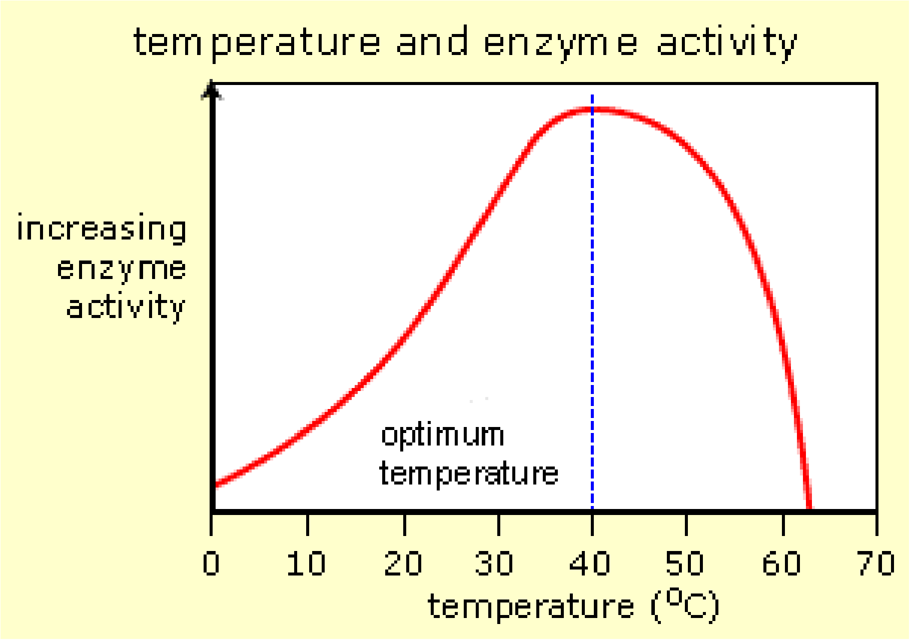
A **surface catalyst** can be **poisoned** when another substance attaches itself to the 'active sites'. This is very often irreversible so **prevents reactant molecules from being adsorbed onto the surface.**

For this reason, catalysts have to be regenerated or renewed.
E.g. **Lead** and its compounds are **poisons** of **transition metal catalysts**. This is why unleaded petrol must be used in cars with catalytic converters.
Arsenic and its compounds are also common poisons.

Catalysts can also be made ineffective by **side-reactions**.
E.g. the iron catalyst used in the Haber Process rusts due to the presence of air and water, so needs to be replaced every so often.

Enzymes

Enzymes catalyse the chemical reactions which take place in living cells.
Enzymes are complex protein molecules which are very specific- they usually only speed up one particular reaction and work best at specific temperatures and pH (**optimum)**.



Examples in nature are:
Amylase – breaks down starch during digestion.
Catalase – breaks down hydrogen peroxide

Many enzymes are used in industry:
Invertase – used in chocolate industry for the hydrolysis of sucrose to form fructose and maltose.
Zymase - converts glucose into alcohol in the brewing industry.
Protease (and others) – used in biological washing powders to dissolve natural stains like protein .

Energy Changes in Chemical Reactions

Chemical reactions involve a change in energy which often results in the loss or gain of heat energy (**exothermic/endothermic reactions)**
The **heat energy** stored in a substance is called its **Enthalpy (H)**.
The difference between the enthalpy of the reactants and the enthalpy of the products in a reaction is the **Enthalpy Change** (ΔH):

$$\boxed{\Delta H = H_{(products)} - H_{(reactants)}}$$

ΔH is measured in kJ per mole (kJ mol^{-1})

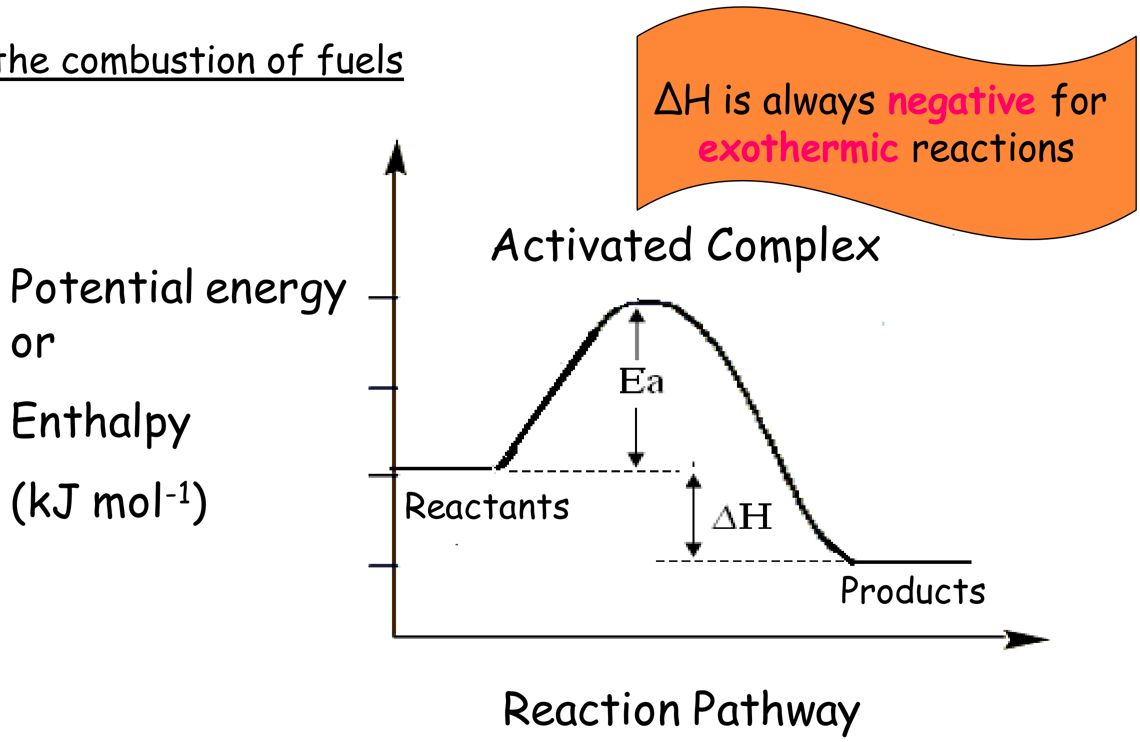
Potential Energy Diagrams

We can show the energy changes involved in exothermic and endothermic reactions by using **potential energy diagrams.**

Exothermic Reactions

Reactions which **give out heat energy** are called exothermic reactions.
The products have less enthalpy (potential energy) than the reactants and the temperature of the surroundings increases.

E.g. the combustion of fuels



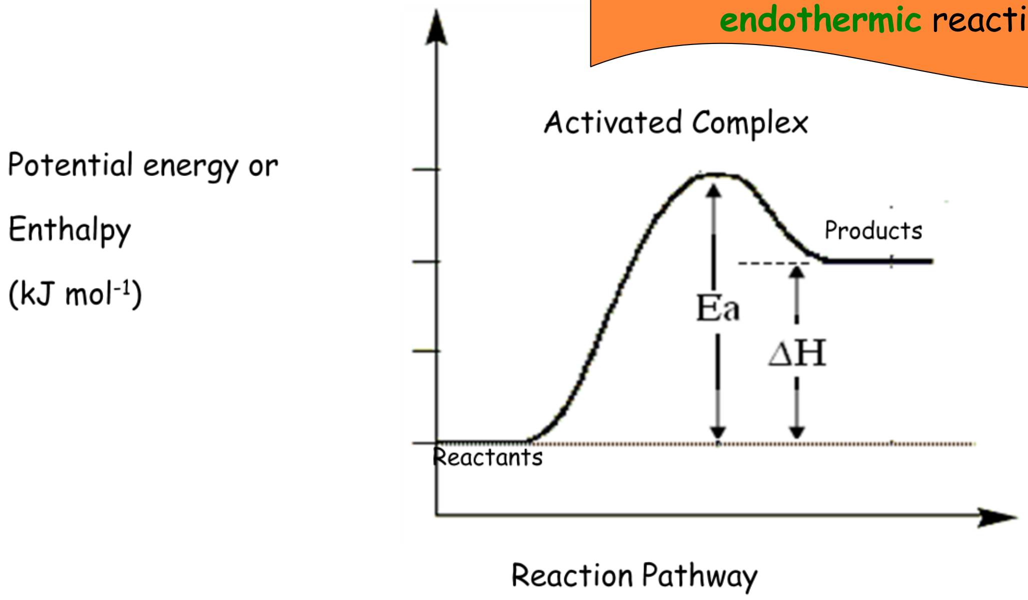
From this diagram we can work out:

- The activation energy (Ea) which is needed to start the reaction.
- The change in enthalpy between the reactants and products (ΔH) = the energy given out by the reaction

Endothermic Reactions

Reactions which absorb heat energy from the surroundings are called endothermic reactions.
The products have more enthalpy than the reactants and the temperature of the surroundings decreases.

ΔH is always **positive** for **endothermic** reactions



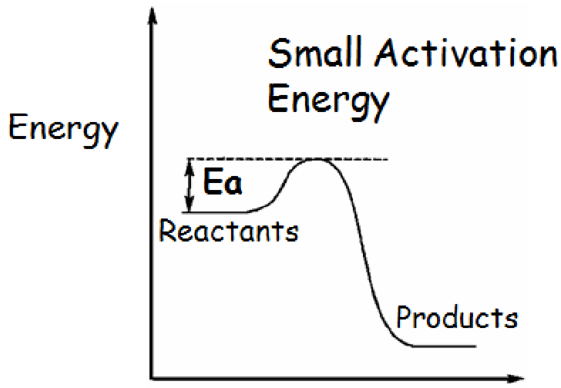
From this diagram we can work out:

- The activation energy (Ea)
- The change in enthalpy between the reactants and products (ΔH) = the energy taken in by the reaction.

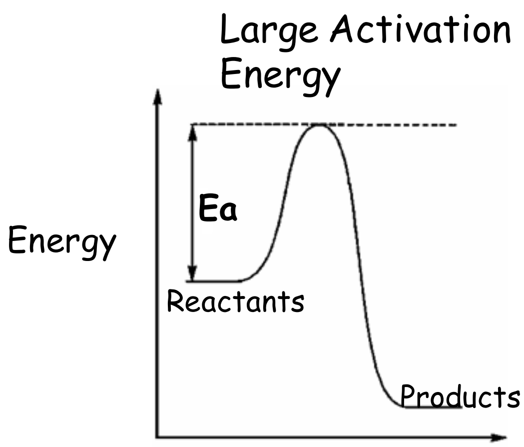
Activation Energy

The Activation Energy is the 'energy barrier' which must be overcome before the reactants can change into products.

The size of the Activation Energy will control how fast or slow a reaction is. The higher the Activation Energy (or 'barrier') the slower the reaction.



Fast Reaction



Slow Reaction

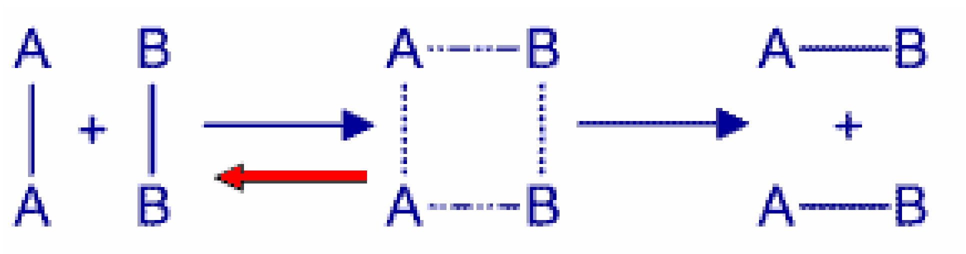
If the activation energy is high, very few molecules will have enough energy to overcome the energy barrier and the reaction will be slow.
e.g. Combustion of Methane

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

This is a very exothermic reaction. At room temperature, no reaction occurs as too few reactant molecules have sufficient energy to react when they collide. Striking a match provides the molecules with enough energy to overcome the barrier- it supplies the Activation Energy. Once started, the energy given out by the reaction keeps it going.

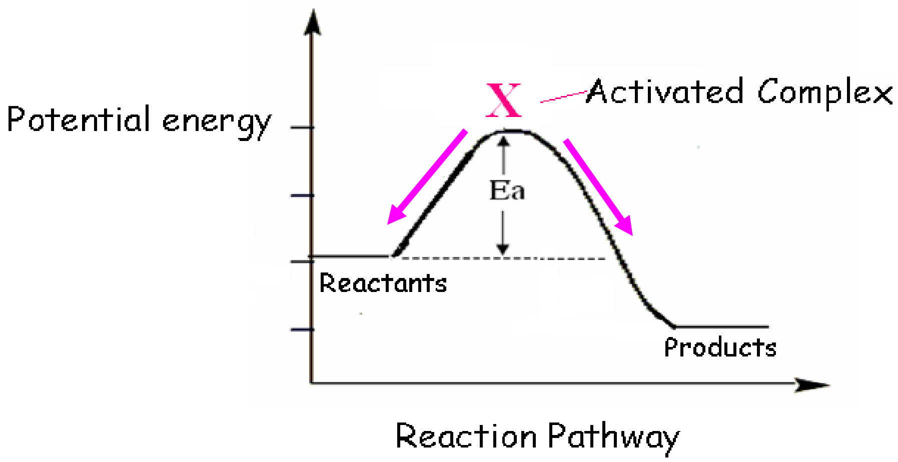
The Activated Complex

When particles collide with the required Activation Energy (& geometry), the activated complex is formed.
The activated complex is an unstable intermediate arrangement of atoms formed as old bonds are breaking and new bonds are forming.
Energy is needed to form the activated complex as bonds in the reactants may need to be broken, or charged particles brought together.



Reactants Activated Complex Products

As the activated complex is very unstable it exists for a very short period of time. From the peak of the energy barrier the complex can lose energy to form either the products or the reactants again.



The higher the enthalpy change (ΔH), the more unstable the activated complex.

Rate of Reactions - Glossary

Word	Meaning
Activation energy	The minimum amount of energy needed for a reaction to begin.
Catalyst	A chemical which speeds up a chemical reaction without being used up itself and which can be removed chemically unchanged at the end of the reaction.
Catalytic converter	A catalyst found in the exhaust of cars. It changes harmful gases into less harmful gases. It is usually made of platinum.
Chemical reaction	An interaction between substances (chemicals) in which their atoms re-arrange to form new substances.
Concentration	The amount of particles in a given volume.
Enzyme	A biological catalyst (ie. a catalyst found in living things).
Fair test	When only one experimental variable is altered at a time.
Products	The substances (chemicals) at the end of a chemical reaction.
Rate of reaction	How quickly a reactant is used up OR how quickly a product is created.
Reactants	The substances (chemicals) at the start of a chemical reaction.
Surface area	Total area of a substance which is exposed to the surroundings.
Variable	Something which can be changed in a chemical reaction.