

## AS90310 Version 2

### Describe thermochemical and equilibrium principles

#### Rates of reaction

##### Particles must collide to react (collision theory).

In order for a chemical reaction to occur the reacting particles (molecules, atoms, ions) must collide with a certain minimum amount of energy known as the **activation energy**,  $E_a$ . They must also collide with the correct orientation.



The greater the rate of effective collisions between molecules in a reaction, the faster the reaction rate.

Any factor that

- increases the frequency\* of collisions (collision rate) **and/or**
- the proportion of these collisions that have the required  $E_a$ ,  
will increase the rate of the reaction.

\* frequency of collisions is not the same as “the number of collisions” or “more collisions” – frequency means the number of collisions/time – ie the RATE .... Number just means how many occur – and doesn’t imply a time period. 100 collisions / second = rate. 100 collisions is just 100 collisions.... They might take a second, an hour, a week.....

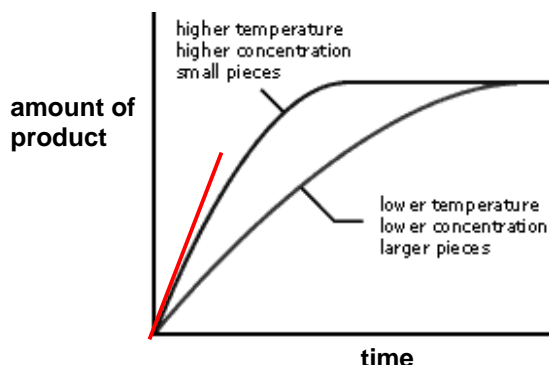
##### Factors affecting rates of reaction

- changes in concentration - an increase in concentration increases the frequency of collisions.  
**Eg 100 mL of 2.0 mol L<sup>-1</sup> HCl reacts at a faster rate with a 5 cm length of Mg ribbon than 100 mL of 1.0 mol L<sup>-1</sup> HCl. For gases an increase in pressure is equivalent to an increase in concentration as the particles are closer together and collide more often).**
- changes in surface area - an increase in surface area increases the frequency of collisions.  
**Eg 2.0 g of powdered calcium carbonate reacts at a faster rate than a 2.0 g marble chip with 100 mL of 2.0 mol L<sup>-1</sup> HCl.**
- changes in temperature - an increase in temperature increases the frequency of collisions AND increases the proportion of these collisions that have the required  $E_a$ . (Actually the increase in collision rate is fairly insignificant compared to the effect on the proportion of molecules with energy  $>$  or  $=$  to  $E_a$  which results because at higher temperatures the particles have more KE)  
**Eg 3.0 g of powdered zinc reacts at a faster rate with 1.0 mol L<sup>-1</sup> HCl at 40°C than it does with 1.0 mol L<sup>-1</sup> HCl at 20°C.**

*(Similarly any factor that decreases these will decrease the reaction rate).*

The graph summarises the differences in the rate of reaction at different temperatures, concentrations and size of pieces (surface area). The steeper the line, the greater the rate of reaction.

Reactions are usually **fastest at the beginning** when the concentration of reactants is greatest. When the line becomes horizontal the reaction has stopped.

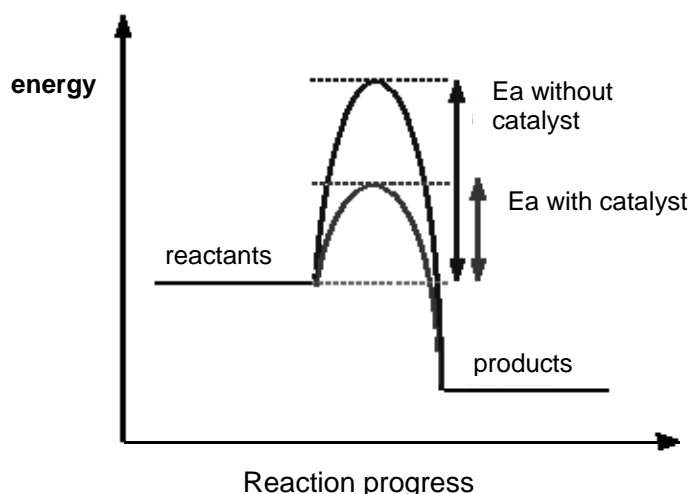


- the presence of a catalyst – a catalyst increases the rate of a chemical reaction without undergoing any overall change. It doesn't get used up. It doesn't turn into a new substance. A catalyst provides an alternative route (reaction mechanism) for the reaction with a lower activation energy. Now a greater proportion of the collisions will have the required  $E_a$  (or more). Although there is an alternative reaction mechanism with a reduced activation energy, the overall enthalpy change is the same for both the catalysed or uncatalysed reaction.

**Eg Hydrogen peroxide is stable at room temperature. The presence of a catalyst may cause it to decompose.** Hydrogen peroxide  $\rightarrow$  oxygen + water  $2\text{H}_2\text{O}_2 \rightarrow \text{O}_2 + 2\text{H}_2\text{O}$

The rate of the reaction can be followed by recording the volume of oxygen produced

The catalyst used is manganese(IV) oxide -  $\text{MnO}_2(\text{s})$



Be very careful if you are asked about this in an exam.

"A catalyst provides an alternative route for the reaction with a lower activation energy."

A catalyst does **not** "lower the activation energy of the reaction". There is a subtle difference. If particles collide with enough energy they can still react in exactly the same way as if the catalyst wasn't there. It is simply that the majority of particles will react via the easier catalysed route.

