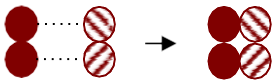
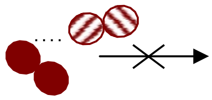
Rates of reaction

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**, Ea. 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.

This means that any factor that

- increases the frequency* of collisions (collision rate) **and/or**
- the proportion of these collisions that have the required Ea,

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/ unit time – i.e. the RATE …. Number just means how many occur – and doesn’t imply a time period. 100 collisions is just 100 collisions…. They might take a millisecond, a second, an hour, a week….. 100 collisions / second is an example of rate.

Factors affecting rates of reaction

- changes in concentration - an increase in concentration increases the frequency of collisions.

  E.g. 100 mL of 2.0 mol L^{-1} HCl reacts at a faster rate with a 5 cm length of Mg ribbon that 100 mL of 1.0 mol L^{-1} 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 concentration increases the frequency of collisions.

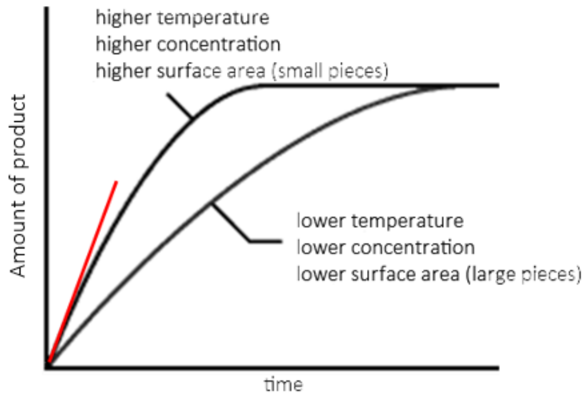
  E.g. 2.0 g of powdered calcium carbonate reacts at a faster rate than a 2.0 g marble chip with 100 mL of 2.0 mol L^{-1} HCl.

- changes in temperature - an increase in temperature increases the frequency of collisions AND increases the proportion of these collisions that have the required Ea to be effective collisions. (Actually, the increase in collision rate is fairly insignificant compared to the effect on the proportion of molecules with energy > or = to Ea which results because at higher temperatures the particles have more kinetic energy).

  E.g. 3.0 g of powdered zinc reacts at a faster rate with 1.0 mol L^{-1} HCl at 40°C than it does with 1.0 mol L^{-1} HCl at at 20°C.

  (Similarly, any factor that decreases the concentration, surface area or temperature 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 (gradient of the) line, the greater the rate of reaction.

Reactions are fastest at the beginning when the concentration of reactants is greatest.

When the line becomes horizontal the reaction has stopped. This is because one of the reactants (known as the limiting reagent) has run out / all reacted.

Note: if there are equal amounts of reactants in both cases then the same amount of product will be made – all that is different is the rate.

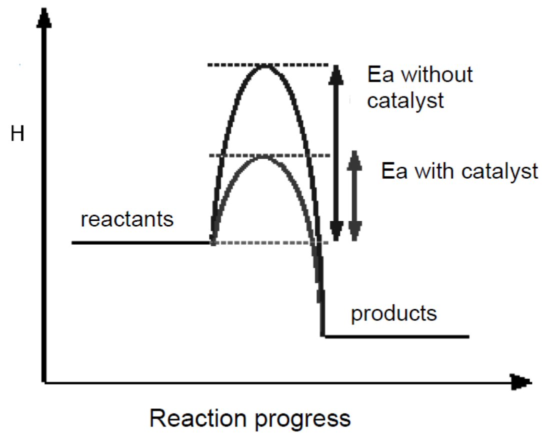
- the presence of a catalyst – a catalyst increases the rate of a chemical reaction without undergoing any overall change. It doesn't get used up. It doesn't turn into a new substance. A catalyst provides an alternative route (or reaction mechanism) for the reaction, with a lower activation energy. Now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

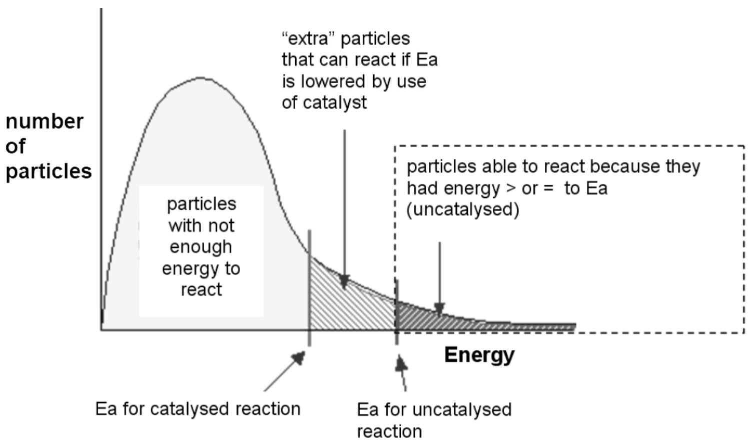
  E.g. Hydrogen peroxide is stable at room temperature. The presence of a catalyst may cause it to decompose.

  Hydrogen peroxide $\rightarrow$ oxygen + water $2H_2O_2 \rightarrow O_2 + 2H_2O$

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

  The catalyst used is manganese(IV) oxide - $MnO_2(s)$





Be very careful if you are asked about how catalysts increase the rate of reaction in an examination. "A catalyst provides an alternative reaction pathway 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.