Collision Theory

MYP Chemistry

Topic: Reaction Rates

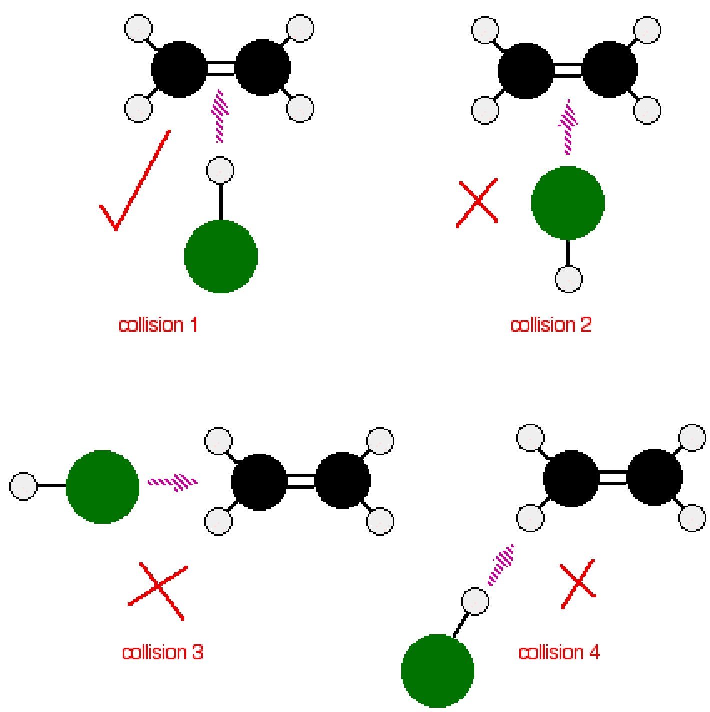
The **Collision theory** explains how chemical reactions occur and why reaction rates differ for different reactions. This theory is based on the idea that reactant particles must **collide** for a reaction to occur, but only a certain fraction of the total collisions have the **energy** to connect effectively and cause the reactants to transform into products. This is because only a portion of the molecules have enough energy and the right **orientation** (or "angle") at the moment of impact to break any existing bonds and form new ones. The minimal amount of energy needed for this to occur is known as **activation energ**y. If the elements react with each other, the collision is called successful, but if the **concentration** of at least one of the elements is too low, there will be fewer particles for the other elements to react with and the reaction will happen much more slowly.

As temperature increases, the average kinetic energy and speed of the molecules increases but this only slightly increases the number of collisions. The rate of the reaction increases with temperature increase because a higher fraction of the collisions overcome the activation energy.

The orientation of collision

Consider a simple reaction involving a collision between two molecules - ethene, $CH_2{=}CH_2$, and hydrogen chloride, HCl, for example. These react to give chloroethane.

$$CH_2{=}CH_2 + HCl \longrightarrow CH_3CH_2Cl$$



As a result of the collision between the two molecules, the double bond between the two carbons is converted into a single bond. A hydrogen atom gets attached to one of the carbons and a chlorine atom to the other.

The reaction can only happen if the hydrogen end of the H-Cl bond approaches the carbon-carbon double bond. Any other collision between the two molecules doesn't work. The two simply bounce off each other.

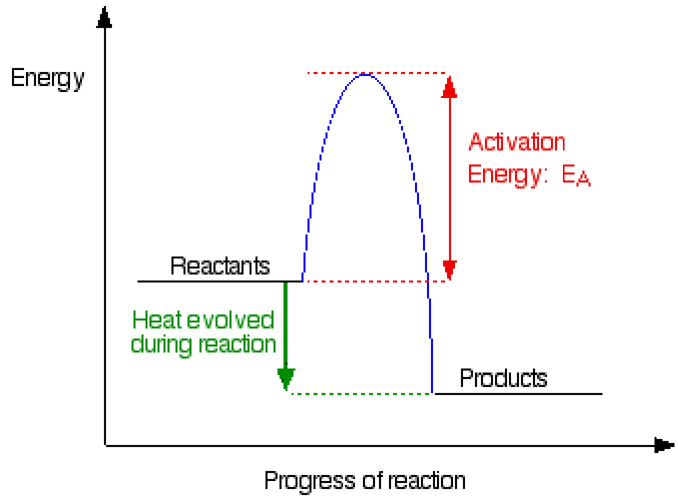
Of the collisions shown in the diagram, only collision 1 may possibly lead on to a reaction.

The energy of the collision

Activation Energy

Even if the species are orientated properly, you still won't get a reaction unless the particles collide with a certain minimum energy called the ***activation energy*** of the reaction.

Activation energy is the minimum energy required before a reaction can occur. You can show this on an ***energy profile*** for the reaction. For a simple over-all exothermic reaction, the energy profile looks like this:



If the particles collide with less energy than the activation energy, nothing important happens. They bounce apart. You can think of the activation energy as a barrier to the reaction. Only those collisions which have energies equal to or greater than the activation energy result in a reaction.

Any chemical reaction results in the breaking of some bonds (needing energy) and the making of new ones (releasing energy). Obviously some bonds have to be broken before new ones can be made. Activation energy is involved in breaking some of the original bonds.

Where collisions are relatively gentle, there isn't enough energy available to start the bond-breaking process, and so the particles don't react.

Collision theory lecture: http://www.youtube.com/watch?v=NO413LiI6QA

Collision Theory Interactive animation: http://www.kscience.co.uk/animations/collision.htm