Edexcel Chemistry A-level

Topic 9: Kinetics I

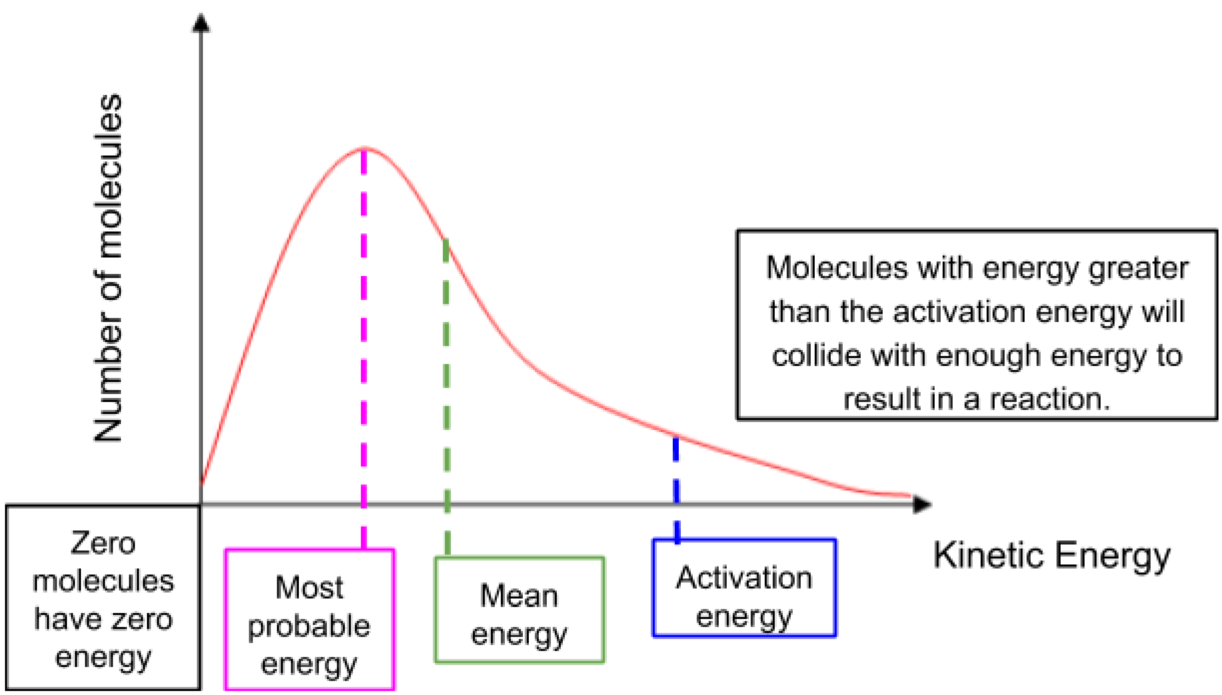
Detailed Notes

Collision Theory

Chemical reactions occur when reactant particles **collide**. For a reaction to occur successfully, these collisions must have energy greater than or equal to the **activation energy** of the reaction, and the **particle orientation** must be correct. The activation energy is the minimum amount of energy required for two particles to react.

Maxwell-Boltzmann Distribution

Not all molecules in a substance have the same amount of energy. Their energies are **distributed** in a pattern called the **Maxwell-Boltzmann distribution**:



Changing the reaction conditions will **alter the shape of the curve,** so that the number of particles with energy greater than the activation energy is different. The total **area under the curve** represents the **total number of molecules** in the sample, and so it **must remain constant**.

Reaction Conditions

The conditions of a reaction impact the collisions of the particles and can be altered to give the particles **more energy**. Therefore, the conditions can be changed to increase the likelihood of a collision occurring with sufficient energy to react. This will lead to a greater rate of reaction.

Rate of Reaction

Rate of reaction can be calculated from empirical data that has been plotted on graphs.

On a **concentration-time graph**, the rate of reaction is equal to the gradient of the curve at a given point. Therefore, the graph can be used to find the rate at a certain time by drawing a tangent to the curve at this given time. Drawing a **tangent** to the curve when **time = 0** finds the **initial rate** of reaction. The tangent at any other position finds the rate of reaction at that moment in time.

Example:

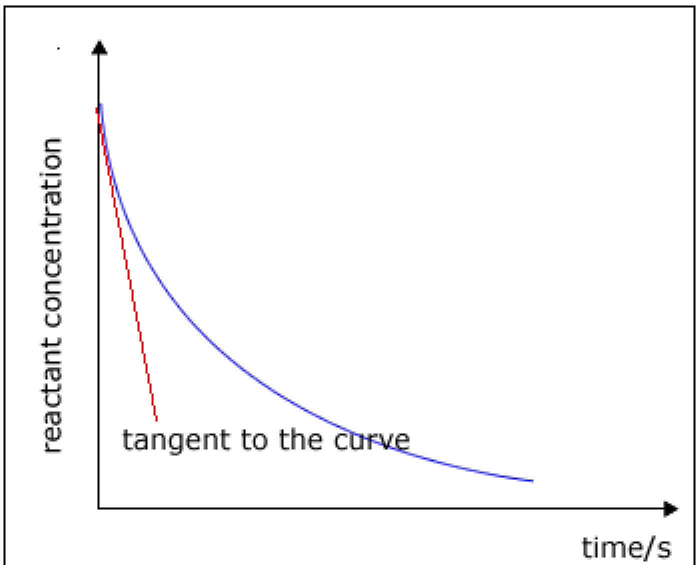


Image courtesy of IBChem

The overall rate of reaction can also be calculated using the following equation:

$$\text{Rate } (s^{-1}) = \frac{1}{\text{Time taken (s)}}$$

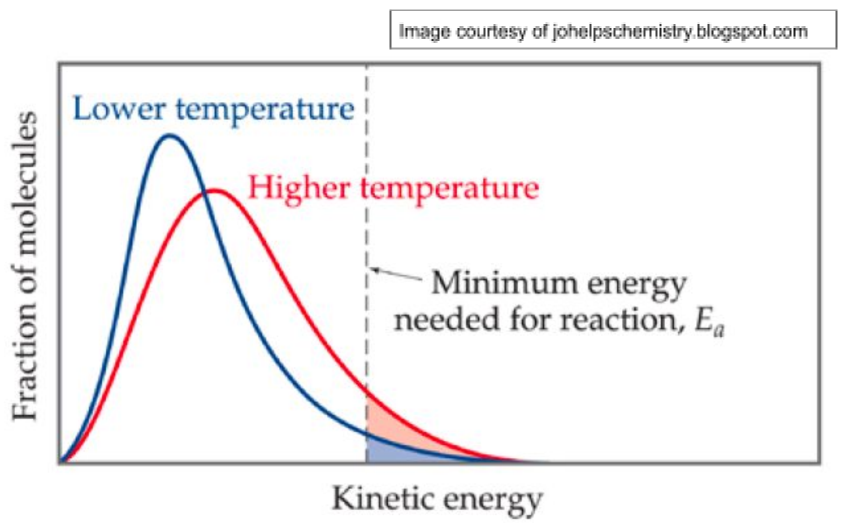
Effect of Temperature

When a substance is heated, **thermal energy** is transferred to it. This energy is converted to **kinetic energy** and the molecules of the substance move **faster and further**. Increased movement of the molecules means **collisions occur more often** and with **greater energy**. As a result, more collisions have energy greater than the activation energy and result in a reaction.

Therefore, **increasing the reaction temperature will increase the rate of reaction** as there are more frequent successful collisions.

The Maxwell-Boltzmann distribution at an increased temperature **shifts to the right** because a **greater proportion** of molecules have energy greater than or equal to the activation energy.

Example:



Effect of Concentration and Pressure

When the concentration of a sample is increased, there are more molecules of substance in the same volume, meaning they are **packed closer together**. Therefore, collisions between molecules become **more likely** and the probability of a collision occurring with energy greater than or equal to the activation energy increases. As a result, the rate of reaction increases.

Increasing the **pressure** of a gas has a similar effect as molecules are **packed closer together** into a smaller volume.

These changes make successful collisions occur more frequently, however, they don't change the **energy** of the **individual particles**. Therefore, the shape of the Maxwell-Boltzmann distribution **does not shift** towards the right as it does with a temperature increase.

Effect of Surface Area

Increasing the surface area of a reactant, for example by crushing it into a powder, increases the **number of exposed reactant particles**. This means there are more frequent, successful collisions, so the rate of reaction **increases**.

As with concentration and pressure changes, it does not change the energy of the individual particles, so the **shape** of the Maxwell-Boltzmann distribution **does not change**.

Effect of Catalysts

A catalyst is a substance that **increases the rate of reaction without being used up** in the reaction. It works by providing an **alternative reaction path** that requires a **lower activation energy** for the reaction to occur.

The Maxwell-Boltzmann distribution curve is **unchanged in shape** but the **position of the activation energy is shifted to the left** so that a greater proportion of molecules have sufficient energy to react.

Example:

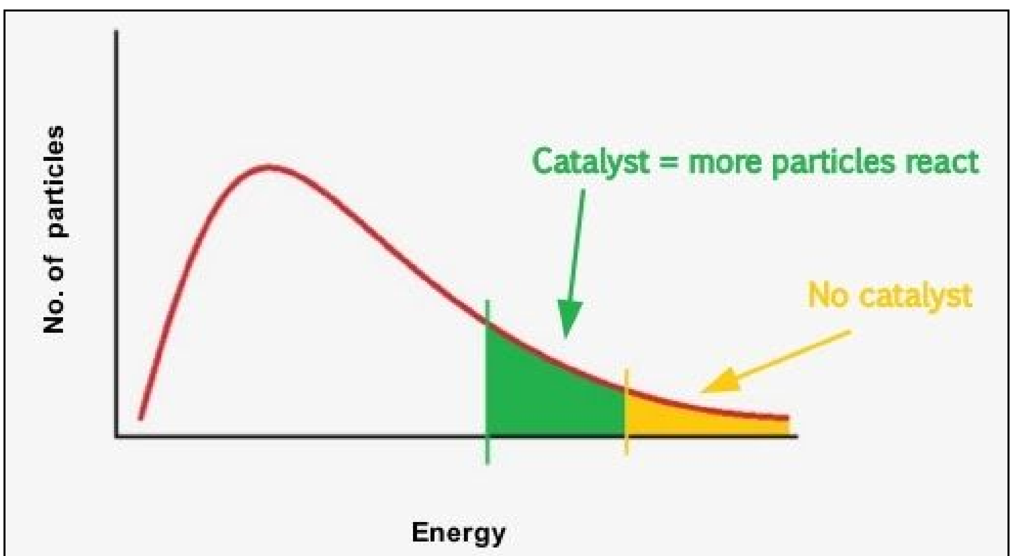
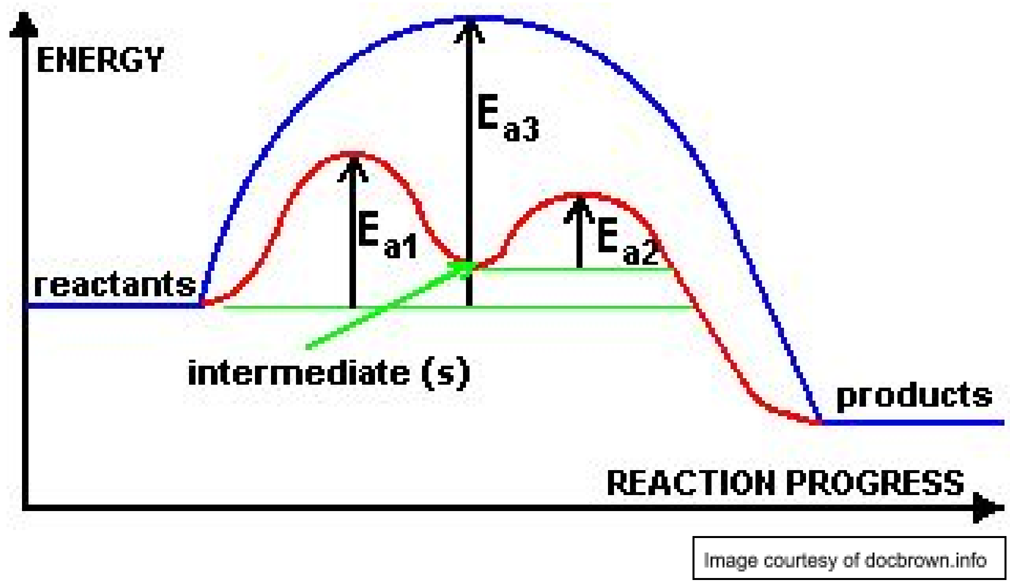


Image courtesy of slideshare.net

Catalysts are used in industry because they **lower the energy costs** of the reaction process. They allow lower temperatures and pressures to be used, whilst still achieving the same rate of reaction. They can also give a **higher atom economy**.

The Reaction Profile of a Catalysed Reaction



The red line shows the pathway for a catalysed reaction, while the blue line shows the pathway for when the reaction occurs without a catalyst.

There is a **dip** in the **energy profile** for the catalysed reaction. This represents the **intermediate** formed during the reaction. The intermediate is **less stable** (and therefore higher in energy) than the reactants and products.

Heterogeneous Catalysts

Heterogeneous catalysts are catalysts that are in a **different phase or state** to the species in the reaction. An example of this is in the Haber Process, where a **solid iron catalyst** is used to speed up the reaction between hydrogen and nitrogen **gases**.

Transition metals make good catalysts as they have variable oxidation states. **Electrons are transferred** to produce a **reactive intermediate** and speed up the reaction rate. An example of this is the **contact process** which uses a vanadium oxide catalyst to speed up the conversion of sulfur dioxide to sulfur trioxide.

In the example below, vanadium is reduced from +5 to +4 and is then **reformed** in its original oxidation state. This indicates that it has acted as a catalyst for the reaction.

Example: The Contact Process

Overall: $2SO_2 + O_2 \xrightarrow{V_2O_5} 2SO_3$

Intermediate Reactions:

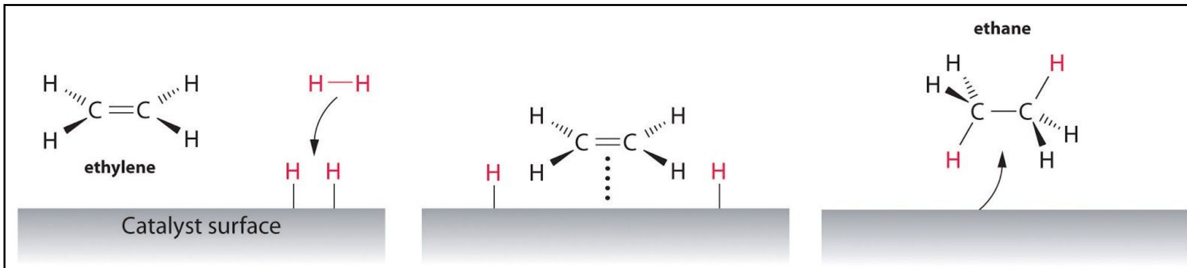
$$V_2O_5 + SO_2 \longrightarrow V_2O_4 + SO_3$$

$$V_2O_4 + 1/2O_2 \longrightarrow V_2O_5$$

Adsorption

A solid catalyst works by **adsorbing** molecules onto an **active site** on the surface of the catalyst. These active sites **increase the proximity** of molecules and **weaken the covalent bonds** in the molecules so that reactions occur more easily and the rate is increased. These catalysts are used in **industry** to give a **surface** for the reaction to occur on.

Example:



(a) Hydrogen (H_2) adsorbs to the catalyst surface (M) to form adsorbed H atoms.

(b) Ethylene (C_2H_4) adsorbs to the catalyst surface.

(c) Ethylene reacts with adsorbed H atoms to give the product ethane (C_2H_6).