Reaction Rates

A **reaction rate** describes **how quickly** a chemical reaction takes place. It is measured in moles per second (mol.s^{-1}) and calculated thus:

$$\frac{No.\,of\ moles\ formed\ or\ used}{Reaction\ time\ in\ seconds}$$

It can also be measured as the change in volume per unit time.

Molecular Collision Theory:

Molecular collision theory dictates that when substances react their molecules or atoms must collide. In order for this to happen, they must be on a collision path and must be travelling fast enough to overcome the repulsive forces of their outer electrons.

In order for an **effective collision** to occur:

- Particles must be orientated on a collision path
- The kinetic energy of the particles must allow them to overcome repulsive forces

Factors affecting reaction rates:

There are **five factors** that affect reaction rates:

1. Temperature
    - In hotter substances, the **kinetic energy** of the particles is greater. This means that the probability of **effective collision** occurring is higher
2. Concentration (moles per unit volume)
    - More concentrated substances have **more particles per unit volume**, increasing the probability of effective collisions occurring
3. Degree of fineness
    - The finer a substance, the greater surface area is exposed. Therefore the probability for collisions increases
4. Catalyst
    - A catalyst is a substance that increases (sometimes decreases) the reaction rate without itself being changed
5. Kind of substance
    - Some substances react faster than others

Energy changes during reactions:

- Thermodynamics:
  - Laws:
    - Energy in = energy out
    - Hot things get cold

1. **Exothermic reactions**
   - **Energy is released** in the form of heat, light, electricity or sound
   - They can be recognised by a **rise in temperature**
   - The **reactants** are at a higher energy state than the **products**
     - When the energy is **liberated**, the energy falls to a **lower level**
   - Heat of reaction (ΔH) = Energy of products – energy of reactants
     - ΔH always refers to the **forward reaction**
     - For exothermic reactions, $\Delta H < 0$
2. **Endothermic reactions**
   - **Energy is taken in continuously** in order for the reaction to proceed
   - Can be recognised by a **drop in temperature** or need for **continuous supply** of heat
   - $\Delta H > 0$
3. **Activation energy**
   - Energy needed to start a reaction
   - Visible on graphs as the 'energy hump'
     - Measured from energy of reactants to energy of activated complex
   - Catalysts lower the energy hump