RATE AND EXTENT OF REACTION

Rate of reaction: Change in concentration of a reactant/product per unit time.
\[
\text{Rate of reaction} = \frac{\Delta c}{\Delta t}
\]

Collision Theory: This theory is based on the assumption that for a reaction to occur, it is necessary for the reacting species (atoms or molecules) to come together or collide with one another. This theory states that three conditions must be met for a reaction to occur:

1. The reacting particles must **collide effectively** with one another in order to react.
2. The reacting particles must collide with **minimum energy** that is equal or higher than the **activation energy** to start the process of breaking and forming bonds.
3. The reacting particles must collide with a **correct orientation** that can lead to re-arrangement of atoms and the **formation of products**.

The reaction rate \(\alpha\) **number of effective collisions per second**

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**Definitions**

- **Activation energy**: The minimum energy needed for a reaction to take place.
- **Activated complex**: The unstable transition state from reactants to products.
- **Catalyst**: A substance that increases the rate of a chemical reaction without itself undergoing a permanent change.

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**Factors affecting the rate of a reaction**

1. **Surface area (for solids)**: The reaction rate increases with an increase in the surface area of the reactants.

Reactants only react if their particles **collide effectively**. Increasing the surface area of the reactants results in **more reaction sites**. (Reaction sites - specific sites on molecules at which reactions occur).

Increasing the number of reaction sites **increases the frequency of collisions**. More effective collisions per unit time will occur which will increase the reaction rate.

**NB**: We can increase the surface area of a solid reactant by grinding it into powder.

2. **Concentration (for solutions)**: The higher the concentration of a reactant in solution, the higher the rate of the reaction.

Concentration is the number of moles per unit volume. So when the concentration increases, the number of reacting particles also increases. This results in more effective collisions per unit time and hence higher rate of reaction.
3. **Pressure (for gases):** Increasing the pressure on reacting gases increases the rate of reaction.

When we increase the pressure, the molecules have less space in which they can move. This *increases* the chance of number of effective collisions per unit time.

4. **Temperature:** An increase in temperature increases the reaction rate.

Higher temperature means higher average kinetic energy of reacting molecules. More molecules have minimum energy equal to or higher than the activation energy. This will cause more effective collisions per unit time. That will speed up the rate of reaction.

5. **Catalyst:** A suitable catalyst increases the rate of reaction.

Collisions only result in a reaction if the particles collide with a certain minimum energy called the activation energy. **Adding a catalyst provides an alternative route with lower activation energy for the reaction.** The majority of particles will now react via the easier catalysed route with lower energy. More effective collisions per unit time will occur which will increase the reaction rate.

![](image)

**Measuring rate of reaction**

The following are some methods which can be used to measure reaction rate.

1. **Measuring the volume of a gas produced.**

The reaction is carried out in **Erlen Meyer (conical) flask** and the flask is connected to **a syringe**, as shown in the diagram. The volume of gas indicated on the syringe scale is read off at regular **time intervals**. A graph can be drawn with Volume vs time as shown below.
The gradient of the graph \( \frac{\Delta V}{\Delta t} \) gives the rate of reaction. Here the unit of reaction rate is \( \text{dm}^3\cdot\text{s}^{-1} \) or \( \text{cm}^3\cdot\text{s}^{-1} \).

A burette or an inverted measuring cylinder may also be used to measure the volume of a gas. The gas is collected over water.
2. Measuring the changes in mass (For gas forming reaction)

If a gas is formed during the reaction, the total mass of the reaction mixture will decrease as the gas is released, thus the loss in mass of the mixture can be measure at regular intervals during the reaction.

\[
\frac{\Delta m}{\Delta t}
\]

The change in mass per unit time (\( \frac{\Delta m}{\Delta t} \)) gives the rate of reaction. Here the unit of reaction rate is g·s\(^{-1}\)

Eg. The reaction between marble (CaCO\(_3\)) and HCl in a conical flask.

\[
\text{CaCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{CaCl}_2(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g)
\]

The mass of the flask + contents decreases as the reaction progress due to the fact that CO\(_2\) escapes from the flask.

3. Observing the change in colour

In some reactions there is a change in colour which tells us that the reaction is occurring. The faster the colour change the faster the reaction rate.

When ethanoic acid (acetic acid) is titrated against sodium hydroxide, an indicator such as phenolphthalein is added. The solution is clear in an acidic solution and changes to pink when the reaction is completed (end point is reached). If the concentration of the base is increased, the colour change occur faster, showing that a higher concentration of base increased the reaction rate.
**Change in turbidity (cloudiness)**

Some reactions form a precipitate which becomes more opaque (cloudy) as more of the fine solid product is formed in the reaction mixture.

*eg. Experiments with sodium thiosulphate and hydrochloric acid*

The reaction between sodium thiosulphate and hydrochloric acid produces a yellow precipitate of fine sulphur (S).

\[
\text{Na}_2\text{S}_2\text{O}_3(\text{aq}) + 2\text{HCl}(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\ell) + \text{SO}_2(\text{g}) + \text{S(s)}
\]

This makes the reaction mixture cloudy and prevents us from seeing clearly through the mixture.

Take sodium thiosulphate solution in a conical flask and kept over a white paper with cross. Pour dilute HCℓ into it and measure the time (t) it takes the cross to disappear.

\[
\frac{1}{t}
\]

Here \( \frac{1}{t} \) gives the rate of reaction.
Unit is s\(^{-1}\).

**Maxwell-Boltzmann Distribution Curves**

This is a curve that represents the fraction of molecules against energy. A typical curve is shown below.

Only particles under this part of the graph has enough energies to react (\( E_k \geq E_a \))
Effect of a catalyst on reaction rate explained using Maxwell-Boltzman Distribution Curves

When a catalyst is added the catalyst provides an alternative pathway with a lower activation energy. (new catalysed $E_a < E_a$). More molecules now have enough energy ($E_k \geq E_a$) to undergo effective collisions. So more effective collisions per unit time will occur and the reaction rate will increase.

Effect of temperature on reaction rate explained using Maxwell-Boltzman Distribution Curves

When the temperature is increased to $T_2$, the average kinetic energy of the molecules have increased. Now more particles have enough kinetic energy ($E_k \geq E_a$) to undergo effective collisions. Hence more effective collisions per second and the rate of reaction will increase.

**NB**: Note that the total area of the two graphs at different temperatures remains the same. This is because the area under the graph represents the total number of molecules and the number of molecules remains the same at different temperatures.