NAT 5 Revision

Consider the following reaction:

**Magnesium + hydrochloric Acid \rightarrow Magnesium Chloride + Hydrogen**

\[ \text{Mg}(s) + \text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2(g) \]

The volume of hydrogen gas evolved can be measured at fixed time intervals.
Three important aspects of the graph to note:

- A-B Indicates initial reaction rate.

- Graph levels off at C indicating that the reaction was complete (one of the reactants completely used up) and the total volume of hydrogen gas evolved.

- Where slope A-B joins line C is the exact time the reaction is complete.

------------------------ increased reaction rate when:

- Particle size is decreased
- Concentration increases
- Temperature increases

----- -- ---- -- ---- decreased reaction rate when:

- Particle size is increased
- Concentration decreases
- Temperature decreases
Consider the following reaction

CaCO$_3$(s) + HCl(aq) → CaCl$_2$(aq) + H$_2$O(l) + CO$_2$(g)

Results Table
Graph

Mass of $\text{CO}_2(g)$ (g) against Time (s)

Average Rate Calculation

The average rate of a chemical reaction can be calculated from:

$$\text{Average Rate} = \frac{\text{Change in Quantity}}{\text{Change in Time}} = \frac{\Delta Q}{\Delta t}$$
Use the above graph to calculate the average rate of reaction between 90 and 150s

90s → 2.40g
150s → 3.06g

Average Rate = \( \frac{\text{Change in Quantity}}{\text{Change in Time}} \)

\[
\begin{align*}
\Delta Q &= (3.06 - 2.40) g \\
\Delta t &= (150 - 90) s \\
\text{Average Rate} &= \frac{0.66g}{60s} \\
&= 0.011 g s^{-1}
\end{align*}
\]

Graph

*Concentration of Acid against Time*
Average Rate Calculation

Use the above graph to calculate the average rate of reaction between 60 and 180s

\[
\begin{align*}
60 \text{s} & \rightarrow 2.25 \text{moll}^{-1} \\
180 \text{s} & \rightarrow 3.32 \text{moll}^{-1}
\end{align*}
\]

\[
\text{Average Rate} = \frac{\text{Change in Quantity}}{\text{Change in Time}} = \frac{\Delta Q}{\Delta t} = \frac{(3.32 - 2.25) \text{moll}^{-1}}{(180 - 60) \text{s}}
\]

\[
= \frac{1.07 \text{moll}^{-1}}{120 \text{s}}
\]

\[
= 0.011 \text{moll}^{-1} \text{s}^{-1}
\]

The rate of a chemical reaction can be followed by recording the following variables against time:

- Mass
- Volume
- Concentration
- pH
- Colour
**Controlling the Rate of a Reaction**

**Factors Affecting the Rate of a Chemical Reaction**

There are four factors that affect the rate of a chemical reaction:

- Particle Size
- Concentration
- Temperature (Kinetic Energy)
- Catalyst (homogeneous or heterogeneous)

**Collision Theory**

For a chemical reaction to occur some important things have to happen:

1. The reacting particles must collide together.

2. Collisions must have sufficient energy to produce a product.

3. The reacting particles must have the correct geometry.

Therefore anything that increases the number of and kinetic energy of collisions between reactant particles will increase reaction rate.
Consider the following reaction

\[
\text{Calcium Carbonate} + \text{Hydrochloric Acid} \rightarrow \text{Calcium Chloride} + \text{Water} \quad \text{Carbon Dioxide}
\]

(Marble Chips)

**Particle Size**

The smaller the particle size the faster the reaction rate.

**Why? - Theory**

A lump of calcium carbonate (marble chips) when cut into smaller pieces has a greater surface area exposed on which many more collisions can take place.

1. Acid particles can only collide with the outside 'faces' of the marble chip.

2. Now the acid can collide with the inner (exposed) 'faces' of the marble chip when it is cut into smaller pieces.

Increasing the number of collisions increases the chance of successful collisions and therefore increases the reaction rate.
**Temperature**

Increasing the temperature increases the reaction rate.

**Why? - Theory**

At higher temperatures - reactant particles have greater kinetic energy.

The more kinetic energy that a reactant particle has then the more likely its collision will have sufficient energy to be successful and produce product.

Therefore increasing temperature increases the reaction rate.

**Concentration**

Increasing the concentration of the acid increases the reaction rate.

**Why? - Theory**

Concentration is the measure of the number of particles in a certain volume of substance - the more particles present the greater the concentration.

When reactants are mixed and the concentration (number of particles) of one or both is increased then the greater the number of collisions that will occur.

Increasing the number of collisions increases the chance of successful collisions and therefore increases the reaction rate.
**Catalyst**

A catalyst increases (or decreases) the rate of a chemical reaction without being used up or physically changed during the reaction.

Enzymes are biological catalysts that occur in nature.
An Experiment to determine the effects of Concentration on a Chemical Reaction

Aim

To examine the effects of concentration changes on the rate of reaction between hydrogen peroxide and acidified potassium iodide solution by varying the concentration of the potassium iodide solution.

Theory

\[
H_2O_{(aq)} + H^+_{(aq)} + 2I^-_{(aq)} \rightarrow 2H_2(l) + I_2(aq)
\]

The course of the reaction is followed by adding small quantities of starch solution and sodium thiosulphate solution.

The iodine molecules produced immediately react with the thiosulphate ions and are converted back into iodide ions (I^-_{(aq)}).

\[
I_2(aq) + 2S_2O_3^{2-}_{(aq)} \rightarrow 2I^-_{(aq)} + S_4O_6^{2-}_{(aq)}
\]

During this process the reaction mixture is colourless.

Once all the thiosulphate ions (S_2O_3^{2-}_{(aq)}) have reacted, iodine molecules (I_2(aq)) form and then react immediately with the starch molecules to form a blue/black colour.
**Method**

![Image of reaction setup]

- 5 cm³ of H₂O₂(aq)
- 10 cm³ of H₂SO₄(aq)
- 10 cm³ of Na₂S₂O₃(aq)
- 1 cm³ of starch solution
- 25 cm³ of KI(aq)

Repeat the experiment using different concentrations of KI(aq).

**Results**

<table>
<thead>
<tr>
<th>Volume of KI(aq)/cm³</th>
<th>Volume of H₂O/cm³</th>
<th>Time (s)</th>
<th>Rate (1/s⁻)</th>
</tr>
</thead>
<tbody>
<tr>
<td>25</td>
<td>0</td>
<td>23</td>
<td>0.043</td>
</tr>
<tr>
<td>20</td>
<td>5</td>
<td>29</td>
<td>0.034</td>
</tr>
<tr>
<td>15</td>
<td>10</td>
<td>39</td>
<td>0.026</td>
</tr>
<tr>
<td>10</td>
<td>15</td>
<td>60</td>
<td>0.017</td>
</tr>
<tr>
<td>5</td>
<td>20</td>
<td>111</td>
<td>0.009</td>
</tr>
</tbody>
</table>

**Evaluation**

- **Volume of H₂O₂(aq) and H₂O(l) kept at 25 cm³**

**Why?**

To ensure that the concentration of all the reactants (except KI(aq)) are kept constant – fair test.

- **End point of the reaction.**

**Appearance of blue/black colour is instantaneous.**
Graph of Time(s) against KI\(_{(aq)}\)(cm\(^3\))

This graph indicates that the rate of reaction is inversely proportional to time.

Whereas the graph of rate\((1/t)(s^{-1})\) against KI\(_{(aq)}\)(cm\(^3\))

The rate of reaction is directly proportional to 1/t(s\(^{-1}\))
There are two potential questions arising from the graph of rate(1/t)(s⁻¹) against KI_{(aq)}(cm³).

- Working out rate from graph.
- Calculating time from 1/Rate.
An Experiment to determine the effects of Temperature on a Chemical Reaction

Aim

To examine the effects of temperature changes on the rate of reaction between oxalic acid and acidified potassium permanganate solution by varying the temperature of the reaction mixture.

Equation

\[
5(COOH)_2(aq) + 6H^+(aq) + 2MnO_2(aq) \rightarrow 2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O(l)
\]

Oxalic Acid Permanganate ion
(purple) (colourless)

Method

Heat to 40°C then added oxalic acid (COOH)₂. Repeat at different temperatures (50°C, 60°C, 70°C).

Why did the beaker have to be dry?
To ensure the concentrations of all the reactants remained constant.

Why not at room temperature? - The colour change is gradual (too slow) to be accurately determine the end point of the reaction.

Time from adding the oxalic acid to purple solution until solution turns colourless.
Results

<table>
<thead>
<tr>
<th>Temperature/°C</th>
<th>Time (t)/s</th>
<th>Rate (1/t)/s⁻¹</th>
</tr>
</thead>
<tbody>
<tr>
<td>38</td>
<td>87</td>
<td>0.011</td>
</tr>
<tr>
<td>50</td>
<td>35</td>
<td>0.029</td>
</tr>
<tr>
<td>59</td>
<td>18</td>
<td>0.056</td>
</tr>
<tr>
<td>70</td>
<td>8</td>
<td>0.125</td>
</tr>
</tbody>
</table>

Graph of rate(1/t)/s⁻¹ against Temperature/°C

This graph is a curve so the rate of reaction is not directly proportional to temperature.

Note

The graph indicates that for every 10°C rise in temperature the reaction rate appears to double.
The graph for an explosive reaction would look like the one reproduced below.

Explosive reactions come to completion almost instantaneously

**Photochemical Reactions**

In some chemical reactions light energy is used to increase the number of reactant molecules with energies equal to or greater than the activation energy.

Photosynthesis - where light energy is absorbed by chlorophyll to convert carbon dioxide and water into glucose and oxygen.

Black and white photography - when film is exposed to light energy, silver ions are reduced to silver atoms.
**Effects of Temperature on Reaction Rate**

**Theory**

At a given temperature (say $T_1$) individual molecules within a gas have widely differing temperatures (kinetic energies).

![Graph showing the distribution of kinetic energies at $T_1$.]  

Kinetic energies of individual molecules continually change due to collisions with other molecules, but, at a constant temperature the overall distribution of kinetic energies remains the same.

![Graph showing the number of reactant molecules with kinetic energies greater than the activation energy $E_a$ at $T_1$.]  

The shaded area shows the number of reactant molecules that have kinetic energies equal to or greater than the activation energy $E_a$ at temperature $T_1$.

An increase in temperature (from $T_1$ to $T_2$) causes a significant increase in the number of reactant molecules that have kinetic energies equal to or greater than the activation energy.

![Graph showing the number of reactant molecules with kinetic energies greater than the activation energy $E_a$ at $T_2$.]  

The new shaded area shows the number of reactant molecules that now have sufficient kinetic energy to overcome the activation energy, $E_a$, when the temperature is increased from ($T_1$ to $T_2$).
Activation Energy

Colliding particles must have a minimum amount of kinetic energy to enable them to form products.

This minimum amount of kinetic energy is known as the activation energy ($E_a$).
**Collision Geometry**

Some collisions will not produce a successful reaction even if the reactant particles collide with kinetic energies equal to or greater than the activation energy.

Consider the addition reaction between propene and bromine

a) Unfavourable collision geometry - reaction is unlikely.

b) Favourable collision geometry - reaction is likely.
**Enthalpy**

**Exothermic Reactions**

An exothermic reaction releases energy to its surroundings. These reactions include:

- Combusting fuels
- Neutralising Acids with Alkalis
- Neutralising an Acid with a Reactive Metal
- Displacement Reactions
**Potential Energy Diagram for an Exothermic Reaction**

A potential energy diagram shows the energy pathway as the reactants are converted into products.

Enthalpy is the energy difference between the reactants and the products.

Enthalpy Change has the symbol $\Delta H$ and is measured in KJ/mol.$^{-1}$.

Note

The products have less energy than the original reactants.

**Activation Energy Diagram for an Exothermic Reaction**

The activation energy of an exothermic reaction

Note

The rate of the reaction does not depend on the Enthalpy Change, $\Delta H$. The rate of the reaction depends on the extent of the activation Energy, $E_a$ that has to be overcome - higher the activation Energy, $E_a$ the slower the reaction rate.
Endothermic Reactions

An endothermic Reaction absorbs energy from its surroundings. These reactions include:

Neutralisation of Ethanoic Acid with Sodium Hydrogen Carbonate

And

Dissolving certain salts in water.

As the temperature of the solution decreases the water on the wooden block freezes and the beaker sticks to the block.
Potential Energy Diagram for an Endothermic Reaction

For an endothermic reaction
\[ \Delta H = +ve \]

Note
The product has more energy than the reactants.

Activation Energy Diagram for an Endothermic Reaction

Note
The rate of the reaction does not depend on the Enthalpy Change, \( \Delta H \). The rate of the reaction depends on the extent of the activation Energy, \( E_a \) that has to be overcome - higher the activation Energy, \( E_a \) the slower the reaction rate.
**Activation Complex**

As the reaction proceeds from reactants to products an intermediate state is reached at the top of the activation barrier at which a highly unstable complex called an activation complex is formed.

The activation complexes are highly unstable and only exist for a very short period of time.

![Activation Complex Diagram](image)

**Note**

The activation complex can lose energy to form either the reactants or the products.
Catalyst

A catalyst speeds up (or slows down) the rate of a chemical reaction without being used up or changed during the reaction.

A catalyst works by providing an alternative route by which the reaction can take place.

The reaction pathway has a lower activation energy, $E_a$.

By lowering the activation energy more reactant molecules now have kinetic energies equal to or greater than the activation energy.
Types of catalysts

Homogeneous Catalyst

The catalyst and the reactants are in the same state of matter.

Heterogeneous Catalyst

The catalyst and the reactants are in different states of matter.

Catalytic converters contain a ceramic support material covered in expensive transition metal elements such as platinum and rhodium which catalyses the conversion of CO gas into CO$_2$ gas and oxides of nitrogen gases to N$_2$ gas.

Catalytic converters should be fitted to exhaust systems of cars that run on unleaded petrol, otherwise the leaded compounds produced during the combustion of the fuel will poison the catalyst.
**How a Heterogeneous Catalyst Works**

Catalytic Poisoning

Catalytic poisoning occurs when a substance (impurity) forms strong bonds with the activation site on the surface of the catalysts so reducing the catalysts efficiency.

Regeneration of Poisoned Catalysts

Regeneration involves cleaning the catalyst by removing impurities from the activation sites, usually by heating with a gas that reacts with the impurity to form a gaseous product.
Catalysts in Industry

<table>
<thead>
<tr>
<th>Catalyst</th>
<th>Process</th>
<th>Reaction</th>
<th>Importance</th>
</tr>
</thead>
<tbody>
<tr>
<td>Vanadium(V) oxide</td>
<td>Contact</td>
<td>$2\text{SO}_3 + \text{O}_2 \rightarrow 2\text{SO}_4$</td>
<td>Manufacture of sulfuric acid</td>
</tr>
<tr>
<td>Iron</td>
<td>Haber</td>
<td>$\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$</td>
<td>Manufacture of ammonia</td>
</tr>
<tr>
<td>Platinum</td>
<td>Catalytic oxidation of ammonia</td>
<td>$4\text{NH}_3 + 5\text{O}_2 \rightarrow 4\text{NO} + 6\text{H}_2\text{O}$</td>
<td>Manufacture of nitric acid</td>
</tr>
<tr>
<td>Nickel</td>
<td>Hydrogenation</td>
<td>Unsaturated oils + $\text{H}_2 \rightarrow$ saturated fats</td>
<td>Manufacture of margarine</td>
</tr>
<tr>
<td>Aluminium silicate</td>
<td>Catalytic cracking</td>
<td>Breaking down long-chain hydrocarbon molecules</td>
<td>Manufacture of fuels and monomers for the plastics industry</td>
</tr>
</tbody>
</table>

Contrasting Uses of a Catalyst with Heating the Reaction Mixture

Heating speeds up the reaction rate by increasing the number of reactant molecules with kinetic energies equal to or greater than the activation energy.

A catalyst speeds up the rate of a chemical reaction by proving an alternative reaction pathway with lower activation energy.

The former provides energy to overcome the energy barrier, the latter lowers the energy barrier.

This highlights the importance of catalysts in saving energy (and therefore money) in many industrial processes.