Candidate 7 evidence

THE EFFECT OF CONCENTRATION ON REACTION RATE

Aim:

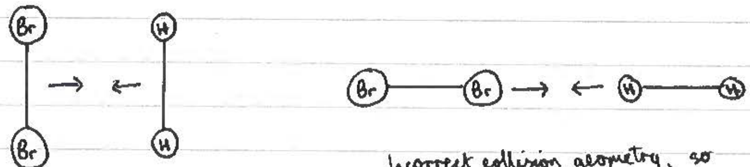
To investigate the effect of increasing the concentration of a reactant, on the rate of reaction.

Underlying Chemistry:

In order for a reaction to take place, collision theory states that reactant particles must collide with enough energy and with the correct alignment. Activation energy (Ea) is the minimum kinetic energy required for a reaction to occur. Reactant particles must collide with enough energy to meet this minimum in order for a collision to potentially be successful and for an activated complex to form.

The reactant particles must also collide with the correct geometry, meaning they must collide at the right angle to allow an activated complex to form.

1)



Correct collision geometry, so potentially successful collision as activated complex may form

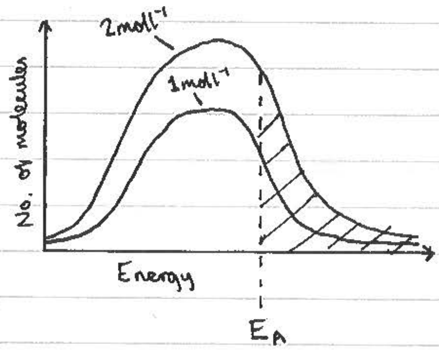
Incorrect collision geometry, so unsuccessful collision as activated complex is unable to form

An activated complex is an unstable arrangement of atoms. It is the 'energy barrier' which must be overcome to allow reactants to become products.

Many factors can impact the rate of a reaction. Increasing temperature, decreasing particle size, increasing pressure, and the use of a catalyst can all speed up a reaction. So these factors must be kept constant when investigating concentration.

By increasing concentration, we can increase the likelihood of successful collisions, thus increasing reaction rate.

This is because increased concentration of a reactant means more reactant particles are moving, meaning more collisions occur, so more successful collisions occur, and the reaction goes faster.



At the higher concentration (2moll⁻¹) there are more molecules present with enough energy to collide successfully.

We can measure the rate of reaction at any given point in time by calculating relative rate:

$$\text{Relative Rate} = \frac{1}{t}$$

The unit for relative rate is s^{-1}.
The relative rate decreases as a reaction goes on. This is because reactants are being used up, so there are less reactant particles and less collisions occur, so less successful collisions occur.

The reaction I am investigating is the reaction between sodium thiosulfate and potassium iodide, commonly known as the 'Iodine Clock' experiment.
The equation for this reaction is:

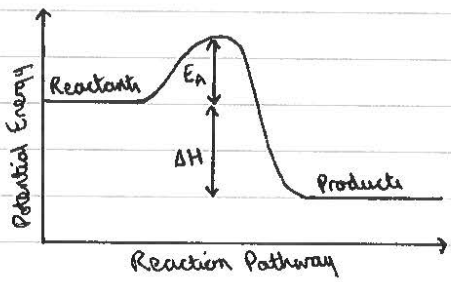
$$H_2O_2 + 2I^- + 2H^+ \longrightarrow I_2 + 2H_2O$$

$$2S_2O_3^{2-} + I_2 \longrightarrow S_4O_6^{2-} + 2I^-$$

In this reaction, iodine ions from the potassium iodide react with hydrogen peroxide, and are oxidised to iodine molecules. As these iodine molecules are produced, they react with thiosulfate ions, and are reduced back to iodine ions. Once all the thiosulfate has been used up, the iodine

molecules react with starch, causing a sudden colour change from colourless to blue / black.
In this reaction, hydrogen peroxide acts as an oxidising agent, and then thiosulfate acts as a reducing agent, and starch acts as an indicator.

This reaction is exothermic, meaning heat is released. This results in the products having less energy than the reactants.



This means the enthalpy change (ΔH) is negative.

Method:
Sulfuric acid, sodium thiosulfate, starch and potassium iodide are added to a beaker. Then hydrogen peroxide is added, and the time taken for the blue / black colour to appear is measured. These steps are then repeated using different concentrations of potassium iodide.
No additional safety measures are required for this experiment.

Results:

Concentration of KI ($mol l^{-1}$)	Time taken for colour change (s)			Average time taken (s)	Relative Rate (s^{-1})
	Attempt 1	Attempt 2	Attempt 3		
0·04	48·5	48·2	47·9	48·2	0·021
0·05	41·2	39·4	39·8	40·1	0·025
0·07	25·7	24·8	25·4	25·3	0·040
0·09	20·6	20·0	19·3	20·0	0·050
0·10	17·4	11·6	16·1	15·0	0·060

Sample Calculations:

$$\text{Average time taken} = \frac{\text{Attempt 1} + \text{Attempt 2} + \text{Attempt 3}}{3}$$

$$= \frac{48.5 + 48.2 + 47.9}{3}$$

$$= 48.2\,s$$

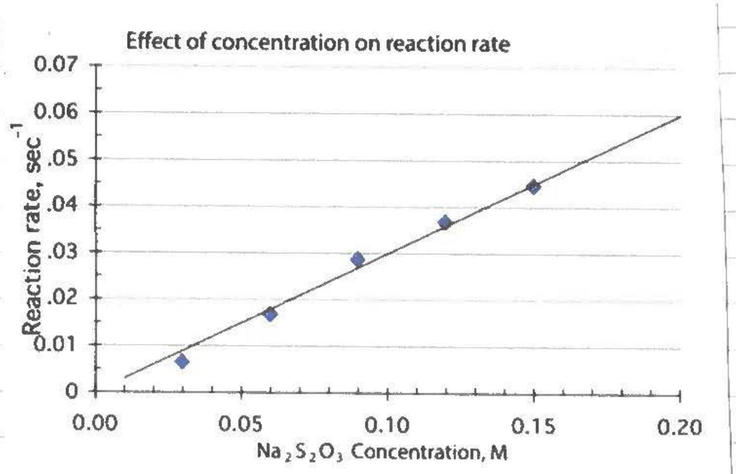
$$\text{Relative Rate} = \frac{1}{t}$$

$$= \frac{1}{48.2}$$

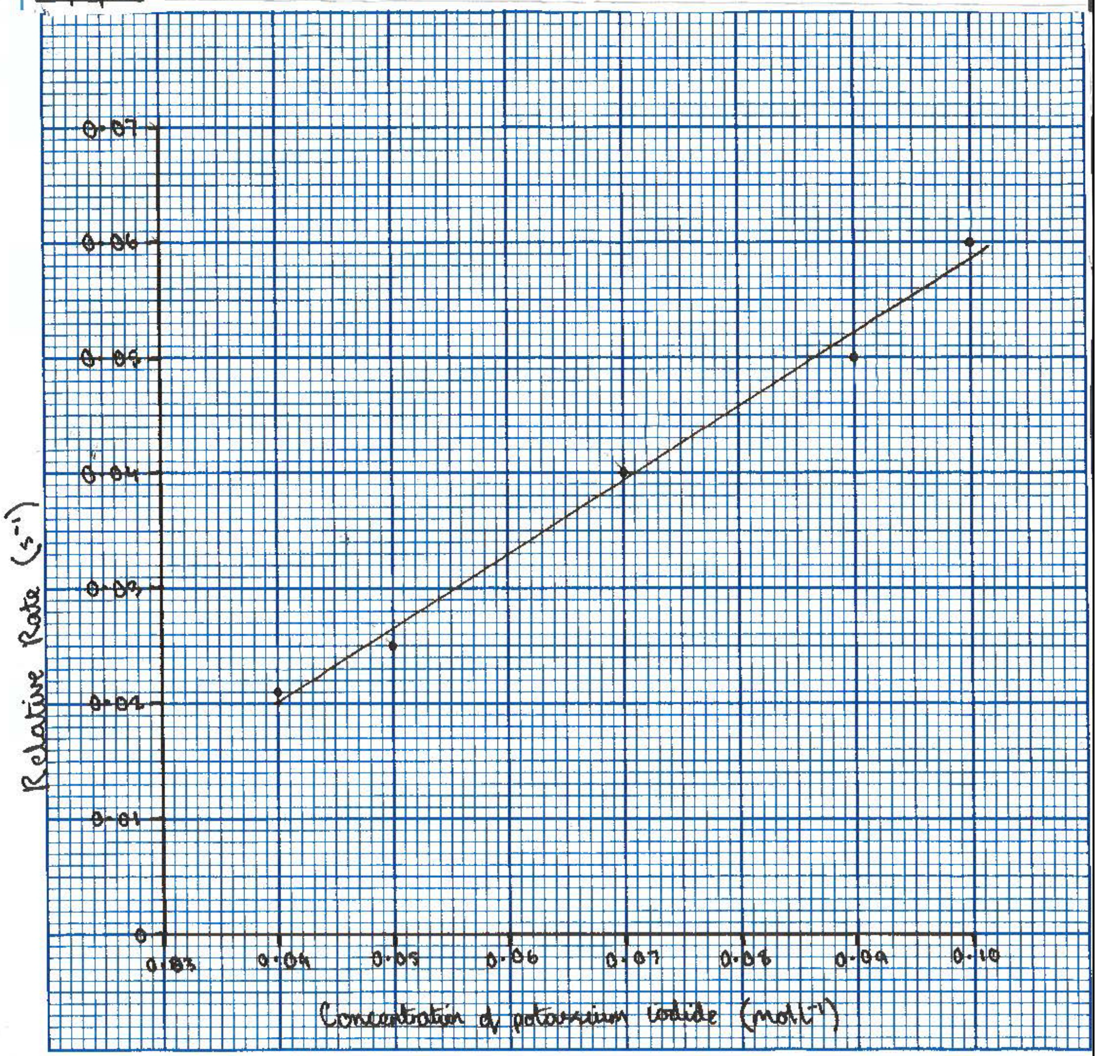
$$= 0.021\,s^{-1}$$

Internet Source:

2)



Graph:



Analysis:

The internet source data shows the general trend of as concentration increases, relative rate also increases. My results show this same general trend. However the internet source investigated a reaction by increasing concentration of $Na_2S_2O_3$ whereas in my reaction I increased concentration of potassium iodide.

My relative rates are higher than the internet source data. For example, at a concentration of 0·09 M the source data has a relative rate of 0·029 s^{-1}, whereas my results produced a relative rate of 0·05 s^{-1} at this concentration. This means my relative rate was 72% higher than the source data's.

Some of my points lie above the line of best fit (Relative rate for 0·04 $mol\,l^{-1}$, 0·07 $mol\,l^{-1}$ and 0·10 $mol\,l^{-1}$) and some points lie below the line of best fit (Relative rates for 0·05 $mol\,l^{-1}$ and 0·09 $mol\,l^{-1}$). This is similar to the source data which also has some results which lie slightly above or below the line of best fit.

<u>Conclusion:</u>

In conclusion, increasing the concentration of a reactant increases the rate of a reaction. This is shown by my results table, graph, and also proven by the internet source, as all these sources show this same general trend of increasing concentration increases reaction rate.

<u>Evaluation:</u>

My internet source data is reliable as it comes from an established chemical company, so the experiment will have been carried out by scientists, using proffessional lab equipment, and will have been repeated and peer reviewed. All making the results very reliable.

Some of my results lie below the line of best fit. The relative rates at 0·05 $mol\,l^{-1}$ and 0·09 $mol\,l^{-1}$ are below the line of best fit. This may be due to water being left in the beaker, so the concentration was lowered and the reaction went slower than expected. To improve this, next time I could dry the beakers in an oven, ensuring no water was left in them and the concentration would not be diluted. This would produce more accurate results.

Some of my repeated attempts produced quite varied results. At 0·10 $moll^{-1}$ Attempt 1 took 17·4s but Attempt 2 took 11·6s. This is a difference of 5·8s. This may be because I used a measuring cylinder to measure the volume of potassium iodide, so the volume measured may have varied slightly between attempts. Next time I could use a pipette to more accurately measure volume of potassium iodide and produce more precise results.

All of my results do not lie exactly on the line of best fit, with 0·05 $moll^{-1}$ and 0·09 $moll^{-1}$ slightly below, and 0·04 $moll^{-1}$, 0·07 $moll^{-1}$ and 0·10 $moll^{-1}$ slightly above. This may be because I did not control temperature during the experiment, so each reaction may have taken place at slightly different temperatures. A higher temperature would produce results higher than expected, and a lower vice versa. To improve this, next time I could use a water bath to ensure the temperature remained constant, and that only one variable (concentration) was being changed. This would improve the validity of my experiment.

References:

1) https://www.bbc.co.uk/bitesize/guides/z2gccdm/revision/2

Date accessed: 15/01/24

2) https://www.flinnsci.com/api/library/Download/78da6c8204aa48a294bd9a51844543ad

Date accessed: 15/01/24