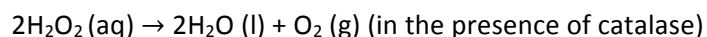


Research Question: What is the activation energy (kJmol^{-1}) of the decomposition of hydrogen peroxide (H_2O_2) to oxygen (O_2) and water (H_2O) by catalase (0.1%), by measuring the time taken for 10cm^3 of oxygen gas to be evolved (s) at different temperatures (K)?

1: Introduction

When we studied catalysts as part of chemical kinetics, I was fascinated by how enzymes function as biological catalysts and I was drawn into the roles enzymes play in biological systems. I found that a particular enzyme, catalase, which is found in animals, catalyses the decomposition of hydrogen peroxide (H_2O_2) in the blood. H_2O_2 is secreted by white blood cells as a defence mechanism against external pathogens. Hence, in order to reduce the exposure of body (somatic) cells to the toxic hydrogen peroxide, catalase decomposes H_2O_2 . (GMO Compass, 2010).



This sparked my curiosity about this particular reaction. I then decided to delve further into reactions involving catalase and found that catalase is also used to preserve egg products by producing oxygen gas when catalysing the decomposition of hydrogen peroxide (GMO Compass, 2010). This oxygen is utilised by glucose oxidase in the egg to catalyse the acidification of glucose to gluconic acid, reacting with all the available glucose in the process. Glucose, in egg products, leads to browning because of its reactions with amino acids present in the albumen of the eggs (Tucker, 1995). Given the importance of the decomposition of hydrogen peroxide by catalase, I questioned the value of the activation energy of the catalysed reaction (E_A). This led me to my research question; ***What is the activation energy (kJmol^{-1}) of the decomposition of hydrogen peroxide (H_2O_2) to oxygen (O_2) and water (H_2O) by catalase (0.1%), by measuring the time taken for 10cm^3 of oxygen gas to be evolved (s) at different temperatures (K)?***

2: Investigation

2.1: Reaction under study

$2\text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$ (in the presence of catalase) at 298.0K, 300.5K, 303.0K, 305.5K and 308K.

2.2: Background Information

Previous research has shown that the rate expression for decomposition of H_2O_2 in the presence of catalase is $rate = k[\text{H}_2\text{O}_2][\text{catalase}]$ (Tao, 2009), where k is the rate constant. The rate refers to the rate of reaction, which is defined in this investigation as the change in the concentration of H_2O_2 per second ($\text{mol dm}^{-3}\text{s}^{-1}$).

The E_A of a reaction is the minimum amount of energy with which reactant molecules need to collide successfully, forming the transition state in the process. It is axiomatic that the E_A of a reaction would be significantly reduced if a catalyst was present, because a catalyst, such as the aptly named catalase, provides an alternative reaction pathway by bringing reactant molecules closer together. Through a series of stochastic collisions, H_2O_2 molecules move into the active site of catalase molecules. Therefore, H_2O_2 molecules are brought closer together by catalase. In the process, it provides an alternative reaction pathway with a lower E_A . Hence catalase, acts as a biological catalyst, reducing E_A , as illustrated on the Maxwell-Boltzmann Distribution and the enthalpy change diagram on the next page.

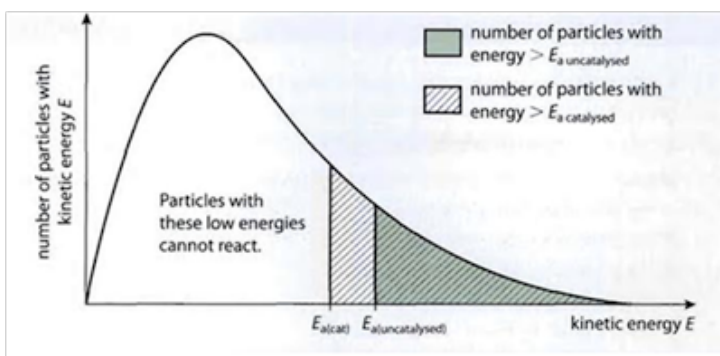


Figure 1: A Maxwell-Boltzmann distribution displaying how there are an increased number of particles with energies more than or equal to the E_A of the catalysed reaction (Gems, 2011)

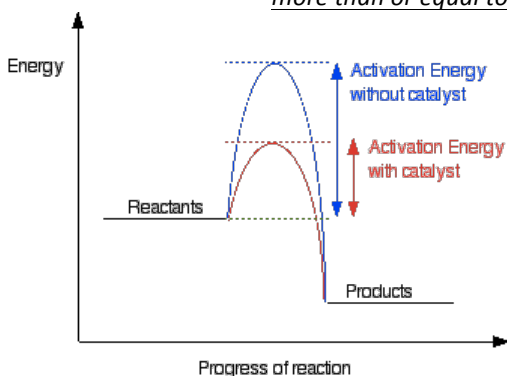


Figure 2: An enthalpy change diagram illustrating the reduction in activation energy for an exothermic reaction when a catalyst is used (Clark, 2013)

2.3: Calculations

The E_A of the decomposition was found through a clock reaction. A stopwatch was started when the reaction began and was stopped when 10cm^3 of oxygen gas was evolved. The number of moles of oxygen evolved was ascertained through the use of the ideal gas law, which is $PV = nRT$, where P is pressure in Pascal, V is the volume of oxygen evolved in m^3 , T is the temperature of the surrounding air in Kelvin (K), n is the number of moles of oxygen evolved and R is the gas constant ($8.3145\text{JK}^{-1}\text{mol}^{-1}$) (John, 2013). Using this equation, the number of moles of oxygen evolved at each temperature was found. With this in mind, the number of moles of H_2O_2 consumed was determined, using the molar ratio between O_2 and H_2O_2 , which is 1:2, as seen from the equation: $2\text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$. The number of moles of H_2O_2 consumed was subsequently divided by the time taken to evolve 10cm^3 of oxygen gas as well as the volume of the solution in dm^3 to produce a value for the rate of reaction in $\text{mol dm}^{-3}\text{s}^{-1}$. Using the equation $R = k[\text{H}_2\text{O}_2][\text{catalase}]$, the rate of reaction was divided by the concentrations of H_2O_2 and catalase to produce a value for the rate constant (k) in $\text{dm}^3\text{mol}^{-1}\text{s}^{-1}$.

To find the activation energy, the Arrhenius equation, which is shown below, was used, where k is the rate constant of the reaction, A is the frequency of successful collisions, R is the gas constant ($8.3145\text{JK}^{-1}\text{mol}^{-1}$) and T is temperature in Kelvin. Please note that $\ln k$ is simply the natural logarithm of k ($\log_e k$).

$$\ln k = \ln A - \frac{E_A}{RT}$$

This equation was plotted using the k values found at each temperature, with $\ln k$ on the y -axis and $\frac{1}{T}$ on the x -axis. The graph obtained was similar to that shown below.

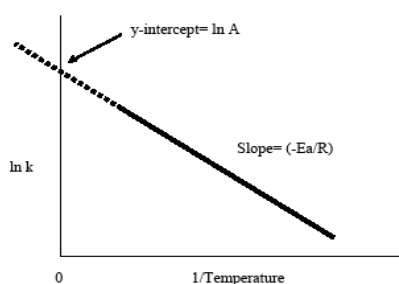


Figure 3: A graph displaying the shape of an Arrhenius graph ($\ln k = \ln A - \frac{E_A}{RT}$) (John, 2013)

Using Microsoft Excel, the line of regression for this graph was sketched and its equation was found, to provide a value for the gradient of the line, which is equal to $-\frac{E_A}{R}$, the coefficient of $\frac{1}{T}$ in the Arrhenius equation. The gradient was then multiplied by $-R$, such that the product of this multiplication was equal to E_A .

3: Variables

Independent Variable: Temperature (K). This is because, use of the Arrhenius equation, requires values of k at different temperatures, in order to plot a graph of $\ln k$ against $\frac{1}{T}$, whose gradient is used to find the value of E_A . Therefore, the independent variable that was chosen was relevant to the investigation, whose purpose is to find the E_A of the catalysed decomposition of hydrogen peroxide. The temperature values that were tested were 298.0K, 300.5K, 303.0K, 305.5K and 308.0K. The use of 5 different temperature values increased the reliability of the results, because it increased the number of data points on the graph, allowing for a more accurate representation of the linear relationship between $\frac{1}{T}$ and $\ln k$. The values chosen also do not exceed 313K, because previous studies have shown that catalase starts to denature (undergo an irreversible conformation change) at temperatures exceeding 313K (41°C) (Abuchowski, 1977). The conformation change results in the deterioration in the shape of the active site, such that fewer H_2O_2 molecules can “lock” into it. Therefore, if the experiment were conducted at temperatures in excess of 313K, the investigation would yield inaccurate values for k , because the concentration of reacting catalase would be lower than the value used in data processing.

Dependent Variable: The time taken (s) for 10cm^3 of oxygen gas to be evolved. This was selected as the dependent variable, since it allows for the quantification of the rate of reaction ($\text{mol dm}^{-3}\text{s}^{-1}$). By using the equation $R = k[H_2O_2][\text{catalase}]$, we can find the value of the rate constant at each temperature, by dividing the rate of reaction found by the product of the concentrations of hydrogen peroxide and catalase. k is essential for use in the Arrhenius equation, where $\ln k$ is used to find E_A , hence the dependent variable chosen is fully relevant to the investigation. It is important to note that the units for the rate were chosen to be $\text{mol dm}^{-3}\text{s}^{-1}$ because the units for the rate constant for a second order rate expression (this is the order of the reaction under study) are $\text{dm}^3\text{mol}^{-1}\text{s}^{-1}$. Therefore, to ensure the rate constant found is found in terms of $\text{dm}^3\text{mol}^{-1}\text{s}^{-1}$, the units of the rate of reaction must be $\text{mol dm}^{-3}\text{s}^{-1}$.

Controlled Variables

1. **pH:** The pH was kept constant at pH 7 using a sodium hydroxide buffer solution. This was to ensure that the catalase in each experiment was operating at its optimum pH (S_u), allowing for an accurate basis for comparison in data processing. A buffer is a solution that resists changes in pH when small amounts of acid or base are added, therefore, it allowed the pH of the mixture to remain constant, with negligible changes. This also ensured that the pH was not a factor that affected the differences in the rate constant at different temperatures.
2. **Concentrations of reactants:** The concentration of H_2O_2 and catalase in the experiment to determine the activation energy of the catalysed decomposition were kept at 0.01mol dm^{-3} and 0.1% respectively to ensure that the changes in the rate of reaction when different temperatures were compared were only caused by temperature, and not concentration, which is another factor that affects the rate of reaction. To do so, samples of $1.5\text{mol dm}^{-3} H_2O_2$ were diluted to reduce their concentration to 0.01mol dm^{-3} .
3. **Volume of reactants:** The volume of H_2O_2 , the pH 7 buffer and catalase were also kept constant at 10cm^3 , 5cm^3 and 5cm^3 respectively. These quantities were measured and added using a graduated pipette. A low volume and concentration of catalase was chosen because each catalase molecule can react with approximately $4 \cdot 10^7$ molecules of H_2O_2 (RSC, 2007). Therefore, a low volume and low concentration of catalase was chosen, so the progress of the reaction would be easily observable.
4. **Pressure:** Data collected by Singapore’s National Environmental Agency (NEA) has shown that pressure in Singapore, both indoors and outdoors undergoes small fluctuations around 101kPa (NEA, 2015). Therefore, pressure can be considered a controlled variable, since previous statistics

and research have shown that pressure in Singapore stays relatively constant at 101kPa (NEA, 2015). This would have affected the calculations for the number of moles of oxygen gas evolved, because the value of P in the ideal gas equation would fluctuate.

4: Method

4.1: Apparatus

1. Memmert Water Bath ($\pm 0.1\text{K}$)
2. 25 cm^3 pipette ($\pm 0.06\text{cm}^3$)
3. 250cm^3 volumetric flask
4. 50cm^3 glass gas syringe ($\pm 0.1\text{cm}^3$)
5. Retort stand
6. 1 Test Tube
7. 20.5cm rubber tubing
8. 6.67cm^3 of $1.5\text{ moldm}^{-3}\text{ H}_2\text{O}_2$
9. 125cm^3 of 0.1% catalase solution
10. 393.3cm^3 of distilled water

4.2: Photograph of set-up



A photograph taken by myself using an iPhone 6, on 26/04/2016, that displays the experiment in progress in the Memmert water bath, with the mixture of catalase (0.1%) and H_2O_2 (0.01 moldm^{-3}) in the test tube, connected to a glass gas syringe held by a retort stand

4.3: Experimental Procedure

1. Prepare a standard solution of H_2O_2 with a concentration of 0.01 moldm^{-3} by diluting a 1.5 moldm^{-3} sample of H_2O_2 in a volumetric flask. Immediately, seal the volumetric flask to reduce the risk of H_2O_2 decomposing immediately. The concentration of H_2O_2 was kept constant at 0.01 moldm^{-3} because the decomposition of hydrogen peroxide with catalase present is a very fast reaction, hence a low concentration of H_2O_2 was chosen to allow for the progress of the reaction to be more easily observed, therefore reducing systematic error.
2. Set the Memmert water bath to 298.0K ($\pm 0.1\text{K}$).
3. Use a graduated pipette ($\pm 0.06\text{cm}^3$) to measure exactly 5cm^3 of the catalase solution and place it into a test tube. For this and all remaining measurements with the pipette, read the pipette at the meniscus to ensure that the volumes of solutions added to the test tube are accurate.
4. Use a graduated pipette ($\pm 0.06\text{cm}^3$) to transfer 5cm^3 of the pH 7 buffer to the test tube containing the catalase solution.
5. Place the test tube holding the catalase solution into the water bath for exactly 10 minutes, with the lid closed, to allow the temperature of the test tube and its contents to equalise.
6. Connect a 20.5 cm rubber tube to a 50cm^3 glass gas syringe ($\pm 0.1\text{cm}^3$).
7. Use a graduated pipette ($\pm 0.06\text{cm}^3$) to transfer 10cm^3 of the prepared H_2O_2 solution into a test tube and allow the tip of the pipette to touch the surface of the solution to allow for cohesion between any H_2O_2 that remains in the pipette and the solution, to ensure that exactly 10cm^3 of H_2O_2 is added to the solution.
8. Immediately, cover the test tube with the rubber bung connected to the 25-cm^3 gas syringe and start a digital stopwatch ($\pm 0.01\text{s}$).
9. Record the time taken (s) for 10cm^3 of oxygen gas to be evolved using the stopwatch.

- Repeat steps 4-9 for a total of 4 additional times to reduce the impact of random error on the results and allow for the collection of sufficient data.
- Repeat steps 3-10 at the following temperatures; 300.5K, 303.0K, 305.5K and 308.0K ($\pm 0.1K$).

4.4: Risk Assessment

Safety Considerations: H_2O_2 (aq) is a powerful bleaching agent and “causes skin irritation...discolouration, swelling and the formation of papules and vesicles (blisters).” (Fisher Scientific, 2000) Therefore, to ensure a high level of safety during the experiment, latex gloves and goggles were worn throughout the duration of the investigation.

Ethical Considerations: There were no ethical considerations to be taken into account.

Environmental Considerations: There were no environmental considerations to be taken into account.

5: Raw Data

Table 1: A raw data table showing the time taken by each replicate to produce $10cm^3$ of oxygen gas (s) at each temperature (K) for the catalysed decomposition of hydrogen peroxide ($0.01mol dm^{-3}$)

Temperature (K) ($\pm 0.1K$)	298.0	300.5	303.0	305.5	308.0
Replicate	Time taken to produce $10cm^3$ of oxygen gas (s) ($\pm 0.01s$)	Time taken to produce $10cm^3$ of oxygen gas (s) ($\pm 0.01s$)	Time taken to produce $10cm^3$ of oxygen gas (s) ($\pm 0.01s$)	Time taken to produce $10cm^3$ of oxygen gas (s) ($\pm 0.01s$)	Time taken to produce $10cm^3$ of oxygen gas (s) ($\pm 0.01s$)
1	52.32	51.32	49.05	48.32	45.04
2	50.82	49.96	47.78	47.56	42.39
3	51.68	50.23	50.09	47.32	46.45
4	52.73	49.45	50.00	45.10	41.92
5	50.85	48.99	48.78	43.03	42.73
Variance (s^2)	0.74	0.78	0.91	4.71	3.80
Standard Deviation (s)	0.86	0.88	0.95	2.17	1.95

The cancelled values (indicated by a line through the value) were excluded from further calculations as they are anomalous points, as substantiated by the fact that the standard deviation (s) and variance (s^2) of the sets of data they belong to decrease significantly following their removal.

5.1: Qualitative Observations

- As temperature increased, the vigour of the effervescence observed in the test tube visibly increased.
- The colour of the solution remained constant; a very light green colour.
- The gas syringe indicated $5cm^3$ of oxygen gas within 20 seconds, whereas more than 20 seconds was required to produce the remaining $5cm^3$.

6: Processed Data

The average time taken to evolve $10cm^3$ of oxygen gas was found by the following formula

$$\frac{\sum \text{time taken to evolve } 10cm^3 \text{ of oxygen gas for each replicate}}{\text{number of replicates}}$$

Example calculation displaying how the average time taken to evolve 10cm³ of oxygen gas was calculated for 298K

$$\frac{52.32 + 50.82 + 51.68 + 52.73 + 50.85}{5} = 51.68\text{s}$$

To calculate the number of moles of oxygen evolved, the ideal gas law equation ($PV = nRT$) was used.

Example Calculation displaying how the number of moles of oxygen evolved was calculated for 298K

$$PV = nRT = (101,000) \left(\frac{10}{1,000,000} \right) = n(8.3145)(298)$$

$$n = \frac{1.01}{8.31 \cdot 298} = 4.08 \cdot 10^{-4} \text{ moles}$$

The remaining values of n were found in a similar manner.

The number of moles of H₂O₂ consumed was found by multiplying the number of moles of oxygen evolved by 2, because according to the equation for the reaction, the molar ratio of O₂ to H₂O₂ is 1:2.

Example Calculation displaying how the number of moles of H₂O₂ consumed was calculated for 298K

$$2(4.08 \cdot 10^{-4}) = 8.16 \cdot 10^{-4} \text{ moles}$$

The rate of reaction was then calculated by dividing the number of moles of H₂O₂ consumed by the volume of the solution (0.02dm³), and subsequently by the average time taken for 10cm³ of oxygen gas to be evolved.

Example Calculation displaying how the rate of reaction was calculated for 298K

$$\frac{8.16 \cdot 10^{-4}}{(0.02 \cdot 51.68)} = (7.89 \cdot 10^{-4}) \text{ moldm}^{-3}\text{s}^{-1}$$

Table 2: A processed data table showing the average time taken to evolve 10cm³ of oxygen gas (s) (± 0.01 s), number of moles of oxygen evolved (mol), the number of moles of H₂O₂ consumed (mol) and the rate of reaction (moldm⁻³s⁻¹) for each temperature (K) (± 0.1 K)

Temperature (K) (± 0.1 K)	Average time taken to evolve 10cm ³ of oxygen gas (s) (± 0.01 s)	Number of moles of oxygen evolved (10 ⁻⁴ mol)	Number of moles of H ₂ O ₂ consumed (10 ⁻⁴ mol)	Rate of reaction (moldm ⁻³ s ⁻¹)
298.00	51.68	4.08	8.16	$7.89 \cdot 10^{-4}$
300.50	49.99	4.04	8.08	$8.08 \cdot 10^{-4}$
303.00	49.14	4.01	8.02	$8.16 \cdot 10^{-4}$
305.50	47.33	3.98	7.96	$8.41 \cdot 10^{-4}$
308.00	44.74	3.94	7.88	$8.81 \cdot 10^{-4}$

Please note that full values were used in calculations, but the displayed values are shown in a manner that is consistent with the uncertainty of the apparatus used.

(Temperature)⁻¹ values were established by calculating the reciprocal of each temperature value that was tested.

Example calculation displaying how (Temperature)⁻¹ was calculated for 298K

$$(\text{Temperature})^{-1} = \frac{1}{298\text{K}} = 3.36 \cdot 10^{-3}\text{K}^{-1}$$

k (rate constant) values were found using the rate equation for the decomposition of hydrogen peroxide ($R = k[\text{H}_2\text{O}_2][\text{catalase}]$). The rate of reaction at each temperature was found divided by the concentration of H₂O₂ and subsequently by the concentration of catalase ($3.03 \cdot 10^{-5} \text{ moldm}^{-3}$). The concentration of H₂O₂ was assumed to remain constant at 0.01 moldm^{-3} , due to the small number of moles of H₂O₂ consumed in the clock reactions (please see table 2).

The units for the concentration of catalase were converted to moldm⁻³ from % to generate the rate constant. In the calculation of this concentration, a critical assumption was made; that 100g of water

(H₂O) has a volume of 0.1dm³. Hence, the concentration of catalase (percentage by mass) was divided by the molecular mass of catalase (33,000) (RSC).

$$\frac{\text{mass of catalase}}{\text{mass of solution}} \div (\text{Relative molecular mass of catalase})$$

$$= \left(\frac{\text{mass of catalase}}{\text{relative molecular mass of catalase}} \right) \div \text{mass of solution} = \frac{\text{number of moles of catalase}}{\text{mass of solution}}$$

The concentration of catalase remains unchanged, because as a catalyst, it is simultaneously regenerated as it is being used to provide an alternative reaction pathway. This is substantiated by my observation that the colour of the solution (a light green, because of the catalase) remained constant throughout the reaction.

Example calculation displaying how the concentration of catalase (mol dm⁻³) was found

Assuming we have a sample weighing 100g

$$\frac{0.1g}{100g} \div (33000gmol^{-1}) = \frac{0.1g}{0.1dm^3} \div (33000gmol^{-1}) = 3.03 \cdot 10^{-5}mol dm^{-3}$$

k at each temperature was then found by dividing the rate of reaction (*R*) at each temperature by the product of the concentrations of H₂O₂ and catalase.

Example calculation displaying how *k* at 298K was found

$$\frac{\text{rate}}{[\text{H}_2\text{O}_2][\text{catalase}]} = k$$

$$k = \frac{7.89 \cdot 10^{-4}}{(0.01)(3.03 \cdot 10^{-5})} = 2603.96dm^3mol^{-1}s^{-1}$$

ln *k* was calculated by taking the natural logarithm of the calculated *k* values.

Example calculation displaying how ln *k* at 298K was found

$$\ln k = \log_e 2603.96 = 7.86$$

Table 3: A processed data table displaying the rate constants of the reaction (mol dm⁻³s⁻¹) at different temperature (K) (±0.1K) and (temperature)⁻¹ (K⁻¹)

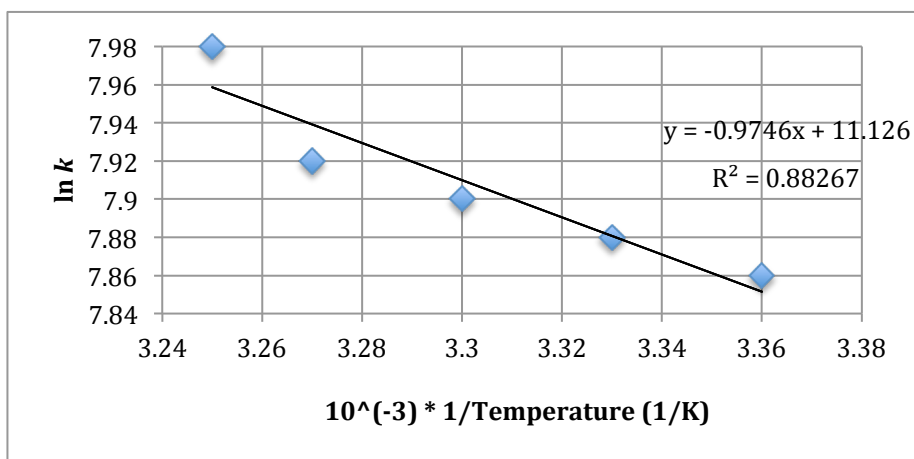
Temperature (K) (±0.1K)	(Temperature) ⁻¹ (10 ⁻³ K ⁻¹)	<i>k</i> (mol dm ⁻³ s ⁻¹)	ln <i>k</i>
298.0	3.36	2600	7.86
300.5	3.33	2670	7.88
303.0	3.30	2690	7.90
305.5	3.27	2780	7.92
308.0	3.25	2910	7.98

Table 4: A processed data table displaying how ln *k* varies with (temperature)⁻¹ (K⁻¹)

(Temperature) ⁻¹ (10 ⁻³ K ⁻¹)	ln <i>k</i>
3.36	7.86
3.33	7.88
3.30	7.90
3.27	7.92
3.25	7.98

An Arrhenius graph was then plotted, with (Temperature)⁻¹ on the x-axis and ln *k* on the y-axis.

Graph 1: An Arrhenius graph displaying how $\ln k$ varies with $(\text{Temperature})^{-1}$ (10^{-3} K^{-1}) with the equation of the line of regression and its R^2 value indicated

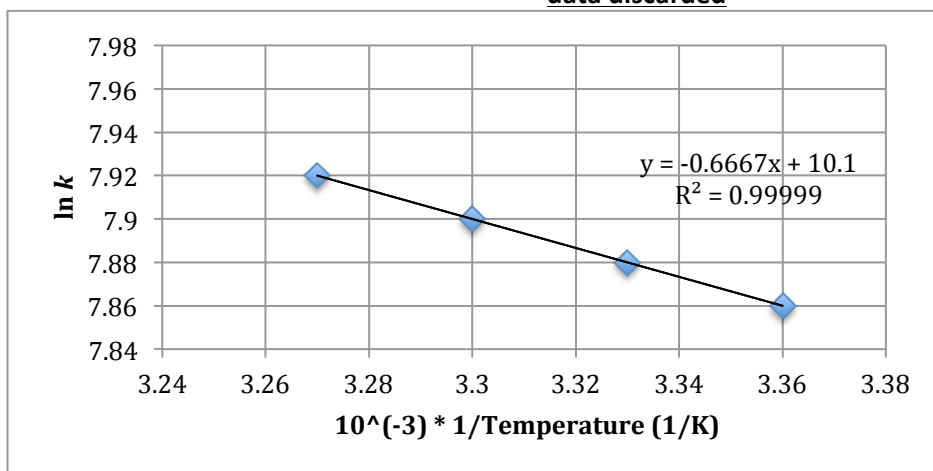


The gradient of the line is given by the coefficient of x on the line of regression, which is -0.9746 . The gradient for an Arrhenius graph, which is the type of graph shown above, is equal to $-\frac{E_A}{R}$. Hence, the gradient was multiplied by $-R$ to provide a value for E_A in Jmol^{-1} .

$$E_A = -0.9746 \cdot (-8.3145) = 8.10 \text{ Jmol}^{-1}$$

The point $(3.25, 7.98)$ does not follow the line of regression as closely as the remaining points, hence it was discarded as an anomalous data point, and E_A was recalculated using the graph below.

Graph 2: An Arrhenius graph displaying how $\ln k$ varies with $(\text{Temperature})^{-1}$ (10^{-3} K^{-1}) with the equation of the line of regression as well as its R^2 value indicated and the anomalous data discarded



$$E_A = -0.6667 \cdot (-8.3145) = 5.54 \text{ Jmol}^{-1} = 0.00554 \text{ kJmol}^{-1}$$

The systematic error of the experiment is another factor that must be taken into account, hence it was calculated in the section below.

7: Calculation of Random Error

The average percentage uncertainty of the time taken for 10 cm^3 of oxygen gas to be evolved across all replicates was found by finding the percentage uncertainty of each measurement for the time taken for 10 cm^3 of oxygen gas to be evolved for each replicate and subsequently dividing this value by 5 (as there were 5 replicates for each temperature).

Example calculation displaying how the average percentage uncertainty of the time taken for 10cm³ of oxygen gas to be evolved across all replicates at 298K was calculated

$$\frac{\left(\frac{0.1}{52.32} \cdot 100\%\right) + \dots + \left(\frac{0.1}{50.85} \cdot 100\%\right)}{5} = 0.194\%$$

The percentage uncertainty of the volume of oxygen evolved, the total volume of solution added to the test tube for each replicate and the temperature the water bath was set to for each replicate were found by the following formula; $\frac{\text{uncertainty of apparatus}}{\text{measured value using the apparatus}} \cdot 100\%$.

Example calculation displaying how the average percentage uncertainty of the volume of oxygen evolved across all replicates was calculated

$$\frac{0.1}{10} \cdot 100\% = 1\%$$

The total uncertainty for each temperature was ascertained by performing the sum of the average percentage uncertainty of the time taken for 10cm³ of oxygen gas to be evolved across all replicates, the percentage uncertainty of the volume of oxygen evolved, the total volume of solution added to the test tube for each replicate and the temperature the water bath was set to for each replicate.

Example calculation displaying how the total uncertainty at 298K was calculated

$$0.1940\% + 1.0000\% + 0.3000\% + 0.0336\% = 1.5276\%$$

Next, the random error of the investigation was calculated, by adding the total uncertainties at each temperature and dividing the sum by 5.

Example calculation displaying how the random error of the investigation was calculated

$$\frac{1.5276\% + 1.5333\% + 1.5371\% + 1.5441\% + 1.5568\%}{5} = 1.5398\%$$

The uncertainty of the E_A value calculated was found by multiplying the random error of the experiment (%) by the E_A value calculated in the manner shown below.

$$1.5398\% \cdot 0.00554\text{Jmol}^{-1} = \pm 0.000853\text{kJmol}^{-1}$$

Table 5: A table showing the error from each apparatus and the total random error for each replicate at each temperature

Temperature (K) ($\pm 0.1\text{K}$)	Average percentage uncertainty of the time taken for 10cm ³ of oxygen gas to be evolved across all replicates (%)	Percentage uncertainty of the volume of oxygen evolved (%)	Percentage uncertainty of the total volume of solution added to the test tube for each replicate (%)	Percentage uncertainty of the temperature the water bath was set to for each replicate (%)	Total uncertainty (%)	Random error of the investigation (%)
298.0	0.1940	1.0000	0.3000	0.0336	1.5276	1.5398
300.5	0.2000	1.0000	0.3000	0.0333	1.5333	
303.0	0.2041	1.0000	0.3000	0.0330	1.5371	
305.5	0.2114	1.0000	0.3000	0.0327	1.5441	
308.0	0.2243	1.0000	0.3000	0.0325	1.5568	

8: Evaluation

8.1: Conclusion

The E_A of the catalysed decomposition of H_2O_2 (0.01mol dm^{-3}) in the presence of catalase (0.1%) was successfully found in the course of this investigation to be $0.00554\text{kJ mol}^{-1} \pm 0.0000853\text{kJ mol}^{-1}$. This value is in general agreement with previous research conducted on the reaction under study. A literature value ($0.00658\text{kJ mol}^{-1}$) (Su) was used in order to calculate the experiment's total error.

$$\text{Total error} = \frac{\text{literature value} - \text{calculated value}}{\text{calculated value}} \cdot 100\% = \frac{0.00658 - 0.00554}{0.00658} \cdot 100\% = 19.9\%$$

The systematic error can now be found by subtracting the random error of the experiment from the total error of the experiment.

$$\text{Systematic error} = \text{Total error} - \text{random error} = 19.9\% - 1.5398\% = 18.3602\%$$

Despite the relatively high systematic error, I have high confidence in my results due to their precision (as can be seen by the low standard deviation and variance amongst replicates) as well as the low total error of the experiment. The experiment is also in agreement with the current scientific consensus, such as the increase in the rate constant as temperature increases, which is substantiated by my observation that at higher temperatures, the effervescence observed in the reaction was more vigorous. This indicates an increase in the rate of reaction as temperature increased, which stemmed from an increase in the rate constant. My observation that the rate of reaction started to decline after 5cm^3 of oxygen gas was produced is supported by the work of P. George, who found that the rate of decomposition of H_2O_2 slowly declined over the course of the reaction, but only marginally (George, 1947). Despite the age of this study, it is reliable because of the stature of the author, a Professor at the University of Cambridge. Because of his position, he had access to extremely precise apparatus and conducted multiple repeats; hence, the conclusions he drew were of low uncertainty and are therefore reliable.

8.2: Strengths

The experiment had low random error (1.5398%) due to low uncertainty of the apparatus used, increasing the certainty of the conclusion drawn in the section above. In addition, the low standard deviation (s) and variance (s^2) in the times taken for 10cm^3 of oxygen to be evolved at each temperature across all replicates were very low, after anomalous points had been removed. This delineates the fact that my results are extremely precise. Furthermore, the high R^2 value indicated in graph 1 (0.99999) demonstrates the accuracy of the processed data because this R^2 value indicates a strong correlation between $(\text{Temperature})^{-1}$ and $\ln k$, which is the ideal description of an Arrhenius graph. My processed data is thus consistent with established scientific theories. The use of a water bath was also a strength of the experiment because it allowed for the uniform distribution of thermal energy in the solution. Hence temperature, as an independent variable, was effectively controlled.

8.3: Weaknesses

However, the experiment had a number of weaknesses.

The data used to find the E_A is limited because of the exclusion of the value of the rate constant at 298.0K, decreasing the number of data points on the Arrhenius graph (Graph 2). This had the effect of increasing the potential impact of random error on the investigation, as substantiated by the relatively high systematic error of 18.3602%. Therefore, the investigation is limited because of the use of only 4 data points for the Arrhenius graph, decreasing the certainty of the conclusion drawn. This can be rectified by repeating the experiment at 298.0K and 295.5K in order to increase the number of data points on the graph, which would decrease the impact of random error on the results.

New solutions of H_2O_2 were only made once every day, hence increasing the possibility that, before being transferred to the test tube, a small amount of the H_2O_2 had possibly decomposed. This possibly reduced the concentration of H_2O_2 , a change that was not accounted for in the calculations, hence

resulting in the values of the rate constants calculated not being accurate representations of the true rate constants. This could have potentially affected the accuracy of the value of E_A that was calculated. This can be rectified by preparing standard solutions of H_2O_2 just before the addition of H_2O_2 to the test tube, allowing for more accurate rate constant values to be calculated. A more accurate value for E_A could then be calculated.

Furthermore, the temperature used in the calculation of the number of moles of oxygen evolved may not have been a representation of the oxygen's true temperature, since its temperature was assumed to be the same as that of the water bath following equalisation. Therefore, it is possible that value for the number of moles of oxygen used in the calculation of $\ln k$ was unreliable, impeding the reliability of the E_A value calculated in this experiment. This can be rectified by inserting a thermometer into the gas jar following the collection of the 10cm^3 of oxygen, such that the actual temperature of the oxygen produced can be measured.

The small range of the independent variable was also a weakness, because it limits the accuracy of the gradient value calculated by Microsoft Excel, as a result of fewer coordinates on the graph. This weakness possibly had an effect on the final E_A value, as the gradient calculated may not have been a representation of the true gradient, consequently affecting the final E_A value calculated. To reduce the impact of this limitation, the experiment could have been repeated at 5 additional temperatures, all lower than 298.0K and none higher than 308.0K , as catalase would denature at temperatures higher than 308.0K .

Furthermore, it is possible that the rate calculated does not reflect the initial rate, because the first 5cm^3 of gas was evolved in less time than the remaining 5cm^3 in all the experiments conducted, indicating that the rate calculated was the average rate, hence the concentration values employed in the rate equation to calculate the different values of k were likely not reflections of the true values. However, if the stopwatch were stopped at 5cm^3 , such that the rate calculated would be the initial rate, the random error of the investigation would increase, as the percentage uncertainty of the volume of gas measured increases from 1% to 2%, thereby increasing the random error of the investigation. Considering this possible increase in uncertainty, it is pellucid that the calculation of the average rate was accurate, as it significantly reduced random error relative to if the investigation calculated the initial rate of reaction.

In addition, the buffer solution used could have potentially affected the accuracy of the final E_A value produced, because it contained sodium hydroxide, whose dissociated sodium ions could have potentially caused in a conformation change in the catalase, due to its positive charge, hence the rate at which the H_2O_2 decomposed was possibly reduced. Conversely, research conducted by Eyster found that the presence of sodium ions had a negligible effect on the rate at which the catalysed decomposition of H_2O_2 occurs in the presence of catalase (Eyster, 1953), hence this limitation had a minor effect on the investigation.

8.4: Extensions

A possible extension to this investigation would be to deduce the difference in the activation energy of the catalysed reactions, in the presence of different catalysts, such as transition metal ions and iodide ions, to find which catalyst can reduce the activation energy of the reaction to the greatest extent. This would uncover the catalyst would best suited in the preservation of egg products. Another investigation could also be carried out to assess if other chemical reactions can produce more oxygen per unit time, relative to the catalysed decomposition of H_2O_2 . This knowledge will be helpful in maximising the efficiency of the preservation of egg products.

8.5: Limitations of the scope of the investigation

However, the investigation is limited because it does not calculate the E_A of the uncatalysed decomposition of H_2O_2 . This limits the extent to which the investigation examines the magnitude of the difference between the activation energy of the uncatalysed and catalysed decomposition of H_2O_2 . This

limitation can be rectified by extending the investigation to conduct the same experiment in the absence of catalase, to find the E_A of the uncatalysed decomposition of H_2O_2 .

9: Bibliography

1. Tucker, G.A. and Woods, L.F.J., 1995. *Enzymes in food processing*. Springer Science & Business Media, Retrieved from <https://books.google.com.sg/books?id=KKDaBwAAQBAJ&pg=PA29&lpg=PA29&dq=glucose+and+gluconic+acid+eggs&source=bl&ots=FqzSaVLC7x&sig=fmmGOPTvNEaYZUtjFibYXR02GTs&hl=en&sa=X&ved=0ahUKEwjz05WMz6TMAhXHHY4KHbKABkIQ6AEIMDAE#v=onepage&q=glucose%20and%20gluconic%20acid%20eggs&f=false>, date of access 14/4/2016, 10:14pm
2. Tao, Z., Raffel, R.A., Soud, A.K. and Goodisman, J., 2009. Kinetic studies on enzyme-catalyzed reactions: oxidation of glucose, decomposition of hydrogen peroxide and their combination. *Biophysical journal*, 96(7), pp.2977-2988.
3. Avogadro Chemistry., Material Safety Data Sheet – Hydrogen Peroxide. ACC#11189 Sections 1-3. Retrieved from http://avogadro.chem.iastate.edu/MSDS/H2O2_30pct.htm, date of access 20/4/2016, 5:30pm
4. GMO Compass., 2010. Catalase. GMO. Retrieved from <http://www.gmo-compass.org/eng/database/enzymes/89.catalase.html>, date of access 5/11/15, 3:23pm
5. Abuchowski, A., McCoy, J.R., Palczuk, N.C., van Es, T. and Davis, F.F., 1977. Effect of covalent attachment of polyethylene glycol on immunogenicity and circulating life of bovine liver catalase. *Journal of Biological Chemistry*, 252(11), pp.3582-3586.
6. Eyster, H., 1953. Effects of certain inorganic ions on corn leaf catalase. *The Ohio Journal of Science*, 53(2), pp.102-104.
7. Su., *Catalase Kinetics*. MIT
8. Gems., 2011. 6.25-6.27. Gems Chemistry Blog . Retrieved from <http://gemchemistry12.blogspot.sg/2011/05/625-627.html>, date of access 27/4/2016, 12:27pm
9. John, D., 2013. *The Arrhenius law: Arrhenius Plots*. Chemwiki. Chemwiki.ucdavis.edu. Retrieved from http://chemwiki.ucdavis.edu/Core/Physical_Chemistry/Kinetics/Modeling_Reaction_Kinetics/Temperature_Dependence_of_Reaction_Rates/The_Arrhenius_Law/The_Arrhenius_Law%3A_Arrhenius_Plots, date of access 27/4/2016, 6:15am
10. National Environment Agency (NEA)., 2015. *Local Climatology*. NEA Weather and Climate. Retrieved from <http://www.nea.gov.sg/weather-climate/climate-information/local-climatology>, date of access 20/4/2016, 5:21pm
11. George, P., 1947. *Reaction between catalase and hydrogen peroxide*. *Nature*, 160, pp.41-43.
12. The Royal Chemistry Society (RSC)., *Catalase*. RSC.
13. RSC., 2007. *Investigating Activation Energies*. RSC. Retrieved from <http://www.rsc.org/Education/EiC/issues/2007Jan/InvestigatingActivationEnergies.asp>, date of access 1/5/2016, 8:43am