***Research Question: What is the activation energy (kJmol-1) of the decomposition of hydrogen peroxide (H2O2) to oxygen (O2) and water (H2O) by catalase (0.1%), by measuring the time taken for 10cm3 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 (H2O2) in the blood. H2O2 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 H2O2. (GMO Compass, 2010).

2H2O2 (aq) → 2H2O (l) + O2 (g) (in the presence of catalase)

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 (EA). This led me to my research question; ***What is the activation energy (kJmol-1) of the decomposition of hydrogen peroxide (H2O2) to oxygen (O2) and water (H2O) by catalase (0.1%), by measuring the time taken for 10cm3 of oxygen gas to be evolved (s) at different temperatures (K)?***

# 2: Investigation

* 1. **: Reaction under study**

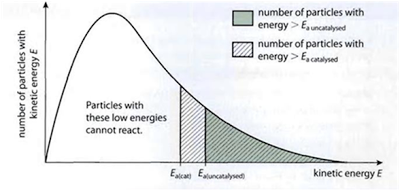
2H2O2 (aq) → 2H2O (l) + O2 (g) (in the presence of catalase) at 298.0K, 300.5K, 303.0K, 305.5K and 308K.

# : Background Information

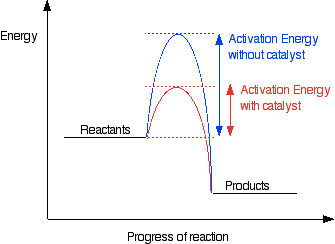
Previous research has shown that the rate expression for decomposition of H2O2 in the presence of catalase is 𝑟𝑎𝑡𝑒 = 𝑘 H2O2 catalase (Tao, 2009), where *k* is the rate constant. The rate refers to the rate of reaction, which is defined in this investigation as the change in the concentration of H2O2 per second (moldm-3s-1).

The EA 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 EA 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, H2O2 molecules move into the active site of catalase molecules. Therefore, H2O2 molecules are brought closer together by catalase. In the process, it provides an alternative reaction pathway with a lower EA. Hence catalase, acts as a biological catalyst, reducing EA, as illustrated on the Maxwell-Boltzmann Distribution and the enthalpy change diagram on the next page.

## 1



*Figure 1: A Maxwell-Boltzmann distribution displaying how there are an increased number of particles with energies more than or equal to the EA 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)*

# : Calculations

The EA of the decomposition was found through a clock reaction. A stopwatch was started when the reaction began and was stopped when 10cm3 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 m3, 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 (8.3145JK-1mol-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 H2O2 consumed was determined, using the molar ratio between O2 and H2O2, which is 1:2, as seen from the equation: 2H2O2 (aq) → 2H2O (l) + O2 (g). The number of moles of H2O2 consumed was subsequently divided by the time taken to evolve 10cm3 of oxygen gas as well as the volume of the solution in dm3 to produce a value for the rate of reaction in moldm-3s-1. Using the equation 𝑅 = 𝑘 H2O2 catalase , the rate of reaction was divided by the concentrations of H2O2 and catalase to produce a value for the rate constant (*k*) in dm3mol-1s-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 (8.3145JK-1 mol-1) and *T* is temperature in Kelvin. Please note that ln 𝑘 is simply the natural logarithm of *k* (log𝑒 𝑘).

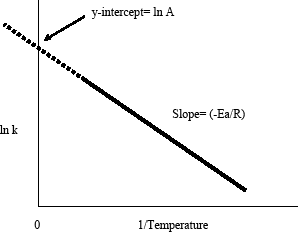
ln 𝑘 = ln 𝐴 − EA

𝑅𝑇

This equation was plotted using the *k* values found at each temperature, with ln 𝑘 on the *y*-axis and 1 on

𝑇

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 − 𝐸A*)*

*(John, 2013)* 2

𝑅𝑇

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 − EA, the coefficient of 1 in the Arrhenius

𝑅 𝑇

equation. The gradient was then multiplied by – 𝑅, such that the product of this multiplication was equal to EA.

# 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 1, whose gradient is used to find

𝑇

the value of EA. Therefore, the independent variable that was chosen was relevant to the investigation, whose purpose is to find the EA 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 1 and

𝑇

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 (410C) (Abuchowski, 1977). The conformation change results in the deterioration in the shape of the active site, such that fewer H2O2 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 10cm3 of oxygen gas to be evolved. This was selected as the dependent variable, since it allows for the quantification of the rate of reaction (moldm-3s-1). By using the equation 𝑅 = 𝑘 H2O2 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 EA, 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 moldm-3s-1 because the units for the rate constant for a second order rate expression (this is the order of the reaction under study) are dm3mol-1s-1. Therefore, to ensure the rate constant found is found in terms of dm3mol-1s-1, the units of the rate of reaction must be moldm-3s-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 (Su), 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 neglible 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 H2O2 and catalase in the experiment to determine the activation energy of the catalysed decomposition were kept at 0.01moldm-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 1.5moldm-3 H2O2 were diluted to reduce their concentration to 0.01moldm-3.
    3. **Volume of reactants**: The volume of H2O2, the pH 7 buffer and catalase were also kept constant at 10cm3, 5cm3 and 5cm3 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 ⋅ 107 molecules of H2O2 (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

* 1. **: Apparatus**
     1. Memmert Water Bath (±0.1K)
     2. 25 cm3 pipette (±0.06cm3)
     3. 250cm3 volumetric flask
     4. 50cm3 glass gas syringe (±0.1cm3)
     5. Retort stand
     6. 1 Test Tube
     7. 20.5cm rubber tubing

8. 6.67cm3 of 1.5 moldm-3 H2O2

1. 125cm3 of 0.1% catalase solution
2. 393.3cm3 of distilled water

# : 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 H2O2 (0.01moldm-3) in the test tube, connected to a glass gas syringe held by a retort stand*

# : Experimental Procedure

* + 1. Prepare a standard solution of H2O2 with a concentration of 0.01 moldm-3 by diluting a 1.5moldm-3 sample of H2O2 in a volumetric flash. Immediately, seal the volumetric flask to reduce the risk of H2O2 decomposing immediately. The concentration of H2O2 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 H2O2 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 (±0.1K).
    3. Use a graduated pipette (±0.06cm3) to measure exactly 5cm3 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 (±0.06cm3) to transfer 5cm3 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 50cm3 glass gas syringe (±0.1cm3).
    7. Use a graduated pipette (±0.06cm3) to transfer 10cm3 of the prepared H2O2 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 H2O2 that remains in the pipette and the solution, to ensure that exactly 10cm3 of H2O2 is added to the solution.
    8. Immediately, cover the test tube with the rubber bung connected to the 25-cm3 gas syringe and

start a digital stopwatch (±0.01s).

* + 1. Record the time taken (s) for 10cm3 of oxygen gas to be evolved using the stopwatch.
    2. 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.
    3. Repeat steps 3-10 at the following temperatures; 300.5K, 303.0K, 305.5K and 308.0K (±0.1K).

# : Risk Assessment

**Safety Considerations:** H2O2 (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 10cm3 of oxygen gas (s) at each temperature (K) for the catalysed decomposition of hydrogen peroxide (0.01moldm-3)

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Temperature (K) (**±**0.1K)** | **298.0** | **300.5** | **303.0** | **305.5** | **308.0** |
| **Replicate** | **Time taken to produce 10cm3 of oxygen gas (s) (**±**0.01s)** | **Time taken to produce 10cm3 of oxygen gas (s) (**±**0.01s)** | **Time taken to produce 10cm3 of oxygen gas (s) (**±**0.01s)** | **Time taken to produce 10cm3 of oxygen gas (s) (**±**0.01s)** | **Time taken to produce 10cm3 of oxygen gas (s) (**±**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 (s2)** | 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 (s2) of the sets of data they belong to decrease significantly following their removal.

# : Qualitative Observations

* + 1. As temperature increased, the vigour of the effervescence observed in the test tube visibly increased.
    2. The colour of the solution remained constant; a very light green colour.
    3. The gas syringe indicated 5cm3 of oxygen gas within 20 seconds, whereas more than 20 seconds was required to produce the remaining 5cm3.

# 6: Processed Data

The average time taken to evolve 10cm3 of oxygen gas was found by the following formula



time taken to evolve 10cm3of oxygen gas for each replicate

number of replicates

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

52.32 + 50.82 + 51.68 + 52.73 + 50.85 = 51.68s

5

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*

10

1,000,000

𝑃𝑉 = 𝑛𝑅𝑇 = 101,000

𝑛 = 1.01

8.31 ⋅ 298

= 𝑛 8.3145 298

= 4.08 ⋅ 10!4 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 ⋅ 10!4 = 8.16 ⋅ 10!4moles

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*

8.16 ⋅ 10!4

!4 !3 !1

(0.02 ⋅ 51.68) = 7.89 ⋅ 10

moldm s

### Table 2: A processed data table showing the average time taken to evolve 10cm3 of oxygen gas (s) (±0.01s), 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) (±0.1K)

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Temperature (K) (**±**0.1K)** | **Average time taken to evolve 10cm3 of oxygen gas (s) (**±**0.01s)** | **Number of moles of oxygen evolved (**𝟏𝟎!𝟒**mol)** | **Number of moles of H2O2 consumed (**𝟏𝟎!𝟒**mol)** | **Rate of reaction (moldm-3s-1)** |
| **298.00** | 51.68 | 4.08 | 8.16 | 7.89 ⋅ 10!4 |
| **300.50** | 49.99 | 4.04 | 8.08 | 8.08 ⋅ 10!4 |
| **303.00** | 49.14 | 4.01 | 8.02 | 8.16 ⋅ 10!4 |
| **305.50** | 47.33 | 3.98 | 7.96 | 8.41 ⋅ 10!4 |
| **308.00** | 44.74 | 3.94 | 7.88 | 8.81 ⋅ 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*

Temperature !1 = 1

298𝐾

= 3.36 ⋅ 10!3K!1

*k* (rate constant) values were found using the rate equation for the decomposition of hydrogen peroxide ( 𝑅 = 𝑘 H2O2 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 ⋅ 10!5 moldm-3). The concentration of H2O2 was assumed to remain constant at 0.01moldm-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).

𝑚𝑎𝑠𝑠 𝑜𝑓 𝑐𝑎𝑡𝑎𝑙𝑎𝑠𝑒 ÷ 𝑅𝑒𝑙𝑎𝑡𝑖𝑣𝑒 𝑚𝑜𝑙𝑒𝑐𝑢𝑙𝑎𝑟 𝑚𝑎𝑠𝑠 𝑜𝑓 𝑐𝑎𝑡𝑎𝑙𝑎𝑠𝑒

𝑚𝑎𝑠𝑠 𝑜𝑓 𝑠𝑜𝑙𝑢𝑡𝑖𝑜𝑛

𝑚𝑎𝑠𝑠 𝑜𝑓 𝑐𝑎𝑡𝑎𝑙𝑎𝑠𝑒

= 𝑟𝑒𝑙𝑎𝑡𝑖𝑣𝑒 𝑚𝑜𝑙𝑒𝑐𝑢𝑙𝑎𝑟 𝑚𝑎𝑠𝑠 𝑜𝑓 𝑐𝑎𝑡𝑎𝑙𝑎𝑠𝑒 = 𝑛𝑢𝑚𝑏𝑒𝑟 𝑜𝑓 𝑚𝑜𝑙𝑒𝑠 𝑜𝑓 𝑐𝑎𝑡𝑎𝑙𝑎𝑠𝑒

𝑚𝑎𝑠𝑠 𝑜𝑓 𝑠𝑜𝑙𝑢𝑡𝑖𝑜𝑛 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 (moldm-3) was found*

Assuming we have a sample weighing 100g

0.1𝑔 ÷ 33000𝑔𝑚𝑜𝑙!1 = 0.1𝑔 ÷ 33000𝑔mol!1 = 3.03 ⋅ 10!5𝑚𝑜𝑙𝑑𝑚!3

100𝑔 0.1dm3

*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*

7.89 ⋅ 10!4

𝑟𝑎𝑡𝑒

= 𝑘

H2O2 catalase

3

!1 !1

𝑘 = (0.01)(3.03 ⋅ 10!5) = 2603.96dm

mol s

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 = log𝑒 2603.96 = 7.86

**Table 3: A processed data table displaying the rate constants of the reaction (moldm-3s-1) at different temperature (K) (**±**0.1K) and** 𝐭𝐞𝐦𝐩𝐞𝐫𝐚𝐭𝐮𝐫𝐞 !𝟏 **(**𝐊!𝟏)

|  |  |  |  |
| --- | --- | --- | --- |
| **Temperature (K)**  **(**±**0.1K)** | 𝐓𝐞𝐦𝐩𝐞𝐫𝐚𝐭𝐮𝐫𝐞 !𝟏  **(10-3 K-1)** | ***k***  **(moldm-3s-1)** | 𝐥𝐧 𝒌 |
| 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 𝒌 |
| 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 !1on the x-axis and ln 𝑘 on the y-axis.

**Graph 1: An Arrhenius graph displaying how** 𝐥𝐧 𝒌 **varies with** 𝐓𝐞𝐦𝐩𝐞𝐫𝐚𝐭𝐮𝐫𝐞 !𝟏

### (10-3 K-1) with the equation of the line of regression and its R2 value indicated

**ln *k***

The gradient of the line is given by the coefficient of x on the line of regression, which is



7.98

7.96

7.94

7.92

7.9

7.88

7.86

7.84

3.24

y = -0.9746x + 11.126

R² = 0.88267

3.26

3.28

3.3

3.32

3.34 3.36

3.38

**10^(-3) \* 1/Temperature (1/K)**

-0.9746. The gradient for an Arrhenius graph, which is the type of graph shown above, is equal to − EA.

𝑅

Hence, the gradient was multiplied by – 𝑅 to provide a value for EA in Jmol-1.

EA = −0.9746 ⋅ −8.3145 = 8.10Jmol!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 EA was recalculated using the graph below.

**Graph 2: An Arrhenius graph displaying how** 𝐥𝐧 𝒌 **varies with** 𝐓𝐞𝐦𝐩𝐞𝐫𝐚𝐭𝐮𝐫𝐞 !𝟏

### (10-3 K-1) with the equation of the line of regression as well as its R2 value indicated and the anomalous data discarded



7.98

7.96

7.94

7.92

7.9

7.88

7.86

7.84

3.24

y = -0.6667x + 10.1 R² = 0.99999

3.26

3.28

3.3

3.32

3.34

3.36

3.38

**10^(-3) \* 1/Temperature (1/K)**

**ln *k***

EA = −0.6667 ⋅ −8.3145 = 5.54Jmol!1 = 0.00554kJmol!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 10cm3 of oxygen gas to be evolved across all replicates was found by finding the percentage uncertainty of each measurement for the time taken for 10cm3 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*

0.1 ⋅ 100% + ⋯ + 0.1 ⋅ 100%

52.32 50.85 = 0.194%

5

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

𝑢𝑛𝑐𝑒𝑟𝑡𝑎(𝑛𝑡𝑦 𝑜ƒ 𝑎𝑝𝑝𝑎𝑟𝑎𝑡𝑢𝑠

by the following formula;

𝑚𝑒𝑎𝑠𝑢𝑟𝑒𝑑 𝑣𝑎𝑙𝑢𝑒 𝑢𝑠i𝑛% 𝑡!𝑒 𝑎𝑝𝑝𝑎𝑟𝑎𝑡𝑢𝑠

⋅ 100%.

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

0.1 ⋅ 100% = 1%

10

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*

1.5276% + 1.5333% + 1.5371% + 1.5441% + 1.5568% = 1.5398%

5

The uncertainty of the EA value calculated was found by multiplying the random error of the experiment

(%) by the EA value calculated in the manner shown below.

1.5398% ⋅ 0.00554Jmol!1 = ±0.0000853kJmol-1

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

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **Temperature (K)**  **(**±**0.1K)** | **Average percentage uncertainty of the time taken for 10cm3 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

* 1. **: Conclusion**

The EA of the catalysed decomposition of H2O2 (0.01moldm-3) in the presence of catalase (0.1%) was successfully found in the course of this investigation to be 0.00554kJmol-1 ±0.0000853kJmol-1. This value is in general agreement with previous research conducted on the reaction under study. A literature value (0.00658kJmol-1) (Su) was used in order to calculate the experiment’s total error.

Total error = 𝑙𝑖𝑡𝑒𝑟𝑎𝑡𝑢𝑟𝑒 𝑣𝑎𝑙𝑢𝑒 − 𝑐𝑎𝑙𝑐𝑢𝑙𝑎𝑡𝑒𝑑 𝑣𝑎𝑙𝑢𝑒 ⋅ 100% = 0.00658 − 0.00554 ⋅ 100% = 19.9%

𝑐𝑎𝑙𝑐𝑢𝑙𝑎𝑡𝑒𝑑 𝑣𝑎𝑙𝑢𝑒 0.00658

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

𝑆𝑦𝑠𝑡𝑒𝑚𝑎𝑡𝑖𝑐 𝑒𝑟𝑟𝑜𝑟 = 𝑇𝑜𝑡𝑎𝑙 𝑒𝑟𝑟𝑜𝑟 − 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 is 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 5cm3 of oxygen gas was produced is supported by the work of P. George, who found that the rate of decomposition of H2O2 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.

# : 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 (s2) in the times taken for 10cm3 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 R2 value indicated in graph 1 (0.99999) demonstrates the accuracy of the processed data because this R2 value indicates a strong correlation between 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.

# : Weaknesses

However, the experiment had a number of weaknesses.

The data used to find the EA 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 H2O2 were only made once every day, hence increasing the possibility that, before being transferred to the test tube, a small amount of the H2O2 had possibly decomposed. This possibly reduced the concentration of H2O2, 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 EA that was calculated. This can be rectified by preparing standard solutions of H2O2 just before the addition of H2O2 to the test tube, allowing for more accurate rate constant values to be calculated. A more accurate value for EA could then be calculated.

Futhermore, 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 EA value calculated in this experiment. This can be rectified by inserting a thermometer into the gas jar following the collection of the 10cm3 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 EA value, as the gradient calculated may not have been a representation of the true gradient, consequently affecting the final EA 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 5cm3 of gas was evolved in less time than the remaining 5cm3 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 5cm3, 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 EA value produced, because it contained sodium hydroxide, whose dissacoiated sodium ions could have potentially caused in a conformation change in the catalase, due to its positive charge, hence the rate at which the H2O2 decomposed was possibly reduced. Conversely, research conducted by Eyster found that the presence of sodium ions had a neglible effect on the rate at which the catalysed decomposition of

H2O2 occurs in the presence of catalase (Eyster, 1953), hence this limitation had a minor effect on the investigation.

# : 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 H2O2. This knowledge will be helpful in maximising the efficiency of the preservation of egg products.

# : Limitations of the scope of the investigation

However, the investigation is limited because it does not calculate the EA of the uncatalysed decomposition of H2O2. 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 H2O2. This

limitation can be rectified by extending the investigation to conduct the same experiment in the absence of catalase, to find the EA of the uncatalysed decomposition of H2O2.

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