

## Experiment 5

### **Experiment Title:**

“A KINETIC STUDY: CATALASE”

### **Author(s) Name(s):**

Adrian Antoniw 300125175

Ahnaf Tahmid 300127653

Balikca Visvalingam 301127783

Bradley Liu 300125419

Cameron Jackson 300114178

Chunyang Lu 300100828

Erin Burnside 300110379

Howard Tseng 300108234

Humza Muhammad 300135666

Jade Mackenzie 300115113

Kirsten Fisher 8832310

Kristyn VanHoof 300111249

Liam Leahy

Maya Williams 300117752

Ross McNamara 300119731

Ryan Cho 300138574

### **Submitting Author's Partner:**

**TA (Demonstrator)'s Name:**

Suttinee Poolsup

**Date Experiment Performed:** November 19, 2019

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**Attach here (if required, indicate the appropriate document(s)):**

**Medical or other Acceptable Document:** \_\_\_\_\_

**Change of Lab Day Form:** \_\_\_\_\_

**Change of Lab Section Form:** \_\_\_\_\_

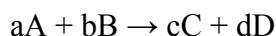
**Late Pass:** \_\_\_\_\_

## Introduction

When humans digest fats and starches, poisonous hydrogen peroxide is produced. Hydrogen peroxide harms humans DNA and cell layers, which causes aging. The enzyme called catalase in humans cells quickens the decomposition of hydrogen peroxide into oxygen and water. Catalase is a heme protein commonly found in the liver, red blood cells, and kidney [1]. Catalase has an optimal pH of 6.8-7.5, when a solution has passed this pH the enzyme will not function as it denatures.

Enzymes are long chains of amino acids that fold into a unique shape that allows it to carry out a specific chemical reaction. Enzymes have a specific active sites built into their structures that allow them to bond to substrates. While bound the enzyme changes the chemical composition of the substrate. In this reaction the  $H_2O_2$  is the substrate and binds to the active site in the catalase. The catalase changes the chemical bonds within  $H_2O_2$  and releases it as oxygen and hydrogen gas.

A chemical reaction can occur at various speeds, either being a fast or slow reaction, and the speed of the reaction is determined through the rate of the reaction. The rate of a reaction is proportional to the result of the fixation of the reactants, each raised to a power. Rate is an important factor when it comes to reaction because a faster rate would give a wanted yield of product in a short period of time, whereas a slow reaction would take a longer period of time to yield the desired amount of product. The order of reaction is necessary to understand how each species in the reaction affect the speed of the reaction in overall. For example, in a general equation:



The rate of the reaction can be expressed as:

$$\text{rate} = k [A]^n [B]^m$$

The species inside the square bracket represents the concentration, which is measured in mol/L. 'n' and 'm' are the order of the reaction which can be determined experimentally using the initial rate at a given temperature and not from the coefficient 'a' and 'b'. The sum of n and m gives the total order of reaction. 'k' represents the rate constant which is unique to different reactions and is dependent on the temperature. It can be found graphically, the initial rate is found from the slope of the start of the graph of time (independent) vs concentration (dependent) [1].

Kinetics are important when looking at catalase. A catalyst lowers the activation energy with an alternate path, therefore increasing the speed of a reaction. Activation energy is the minimum amount of energy needed for a reaction to occur and is measured in kJ/mol. The activation

energy of different catalysts helps determine the most efficient catalyst. In this experiment, the catalyst used were catalase (obtained from lettuce) and potassium iodide.

The rate of reaction is greatly affected by temperature. As temperatures decrease a catalysts rate of reaction will decrease. Molecules in the reaction move slower, substrates and active sites collide less thus reacting less. As temperatures increase the rate of reaction increases. Molecules are agitated and move faster, more substrates collide more frequently with active sites resulting in more products being formed in a shorter amount of time. Thus the rate of reaction increases. However, if the temperature increases too much and the enzyme denatures rendering it functionless.

The Arrhenius equation  $k = Ae^{(-EA/RT)}$  reflects the rate constant to the temperature of a reaction. This equation is useful for finding the activation energy of a reaction. Rearranging this equation gives a function where the rate constant(dependent variable) is dependent on the temperature(independent variable).

$$\ln(k) = - E_a/R(1/T) + \ln A$$

If a molecule does not meet the activation energy, the reaction will not occur. The activation energy can be calculated by performing a reaction at different temperatures, and measuring the rate constant at each temperature. Plotting the graph of  $\ln k$  vs temperature inverse will generate a linear function, with the slope of this graph being  $-E_a/R$ . The activation energy then can be found through isolating the variable from the slope of the graph.

## **Procedure:**

### **Part 1: Obtaining the Catalase**

1. Wear appropriate safety equipment, safety goggles and lab coat.
2. Obtain catalase from the lettuce by blending it. Add 500mL of distilled water per every 50g of lettuce and blend the mixture until all the lettuce leaves have been crushed. Filter the resulting blend using a cloth to obtain the catalase solution.
3. Place the obtained catalase in an ice water bath to preserve the enzyme for the duration of the lab.

### **Part 2: Determining the Order of Reaction**

1. Connect the gas pressure sensor to the LabQuest 2.

2. Determining the partial order of the reaction with respect to  $\text{H}_2\text{O}_2(x)$ .

(For this part of the experiment we determined  $x$  by using three different volumes of  $\text{H}_2\text{O}_2$  keeping all the other variables such as the concentration and volume of the catalase and the temperature at which the experiments are conducted constant.)

3. Start the first trial by measuring 2 mL of  $\text{H}_2\text{O}_2$  using a 10 mL graduated cylinder and add the  $\text{H}_2\text{O}_2$  to a vial.

4. Using another graduated cylinder, measure 10 mL of catalase and add that to a 250 mL conical flask.

5. Use a pair of tweezers to slowly insert the vial into the conical flask making sure that the vial does not tip over so the solutions do not react yet.

6. Connect the gas pressure sensor to a specific bung and use the bung to close the opening of the conical flask.

7. Start collecting data and immediately shake the conical flask making sure that the solutions are well mixed. Continue mixing and collected data for 2 minutes.

8. A graph of pressure( KPa) against time(s) is observed.

9. Redo steps 3-8 for two more trials using the same volume of catalase but increasing the volume of hydrogen peroxide

Half-dilution of the  $\text{H}_2\text{O}_2$ : To 1mL of  $\text{H}_2\text{O}_2$ , add 1 mL of distilled water. This addition of distilled water makes the concentration of the same volume of  $\text{H}_2\text{O}_2$ ( total 2 mL) to be halved.

8. Repeat steps 1-7 for 3 new concentration of  $\text{H}_2\text{O}_2$  (0.75 M, 0.50M, 0.25M)

9.  $x$  can now be calculated.

### **Part 3: Determining the Activation Energy for the Catalase-catalysed Decomposition of Hydrogen Peroxide**

1. Measure out 3 mL of 0.5M  $\text{H}_2\text{O}_2$  into the glass vial using a 10 mL graduated cylinder.

2. Measure out 20 mL of catalase using another 50 mL graduated cylinder and pour the 20 mL of catalase into a 250 mL Erlenmeyer Flask.

3. Use a pair of long steel tweezers to slowly insert the vial into the erlenmeyer flask, ensuring that the vial does not tip over.

4. Fill an 800 mL beaker with water and place the filled beaker on a hot plate to increase the temperature of the water to 30 degrees celsius. Check the temperature of the water in the beaker with a thermometer.
5. Connect the gas pressure sensor to a bung and use the bung to seal off the top of the Erlenmeyer flask.
6. Carefully place the erlenmeyer flask inside the 800mL beaker.
7. Start collecting data using the labquest and move the erlenmeyer flask in a circular motion to tip over the vial inside the erlenmeyer flask (all inside the water bath). Continue mixing and collected data for 60 seconds.
8. A graph of pressure (kPa) against time(s) is observed.
9. Remove the flask from the bath and dispose of the contents inside the flask in the organic waste bin. Rinse the vial and the flask with distilled water.
10. Repeat steps 1-10 using 3 different different temperatures of water (0 °C, 21.9 °C, 30 °C).
11. Plot the data collected in labquest and find the slope (delta pressure/delta time). This value is the initial rate.
12. Once the initial rate has been found, calculate the activation energy for the catalase-catalysed decomposition of  $\text{H}_2\text{O}_2$ .

#### **Part 4: Determining the Activation Energy for the KI-catalysed Decomposition of Hydrogen Peroxide**

1. Measure out 3 mL of 0.5M  $\text{H}_2\text{O}_2$  into the glass vial using a 10 mL graduated cylinder.
2. Measure out 30 mL of KI using a 50 mL graduated cylinder and pour 30 mL of KI into a 250 mL Erlenmeyer Flask.
3. Use a pair of long steel tweezers to slowly insert the vial into the erlenmeyer flask, ensuring that the vial does not tip over.
4. Fill an 800 mL beaker with water and place the filled beaker on a hot plate to increase the temperature of the water to 40 degrees celsius. Check the temperature of the water in the beaker with a thermometer.

6. Connect the gas pressure sensor to a bung and use the bung to seal off the top of the Erlenmeyer flask.
6. Carefully place the erlenmeyer flask inside 800mL beaker.
7. Start collecting data using the labquest and move the erlenmeyer flask in a circular motion to tip over the vial inside the erlenmeyer flask (all inside the water bath). Continue mixing and collected data for 60 seconds.
8. A graph of pressure (kPa) against time (s) is observed.
9. Remove the flask from the bath and dispose of the contents inside the flask in the inorganic waste bin. Rinse the vial and the flask with distilled water.
10. Repeat steps 1-10 for 3 different temperatures (21.9 °C, 30 °C, 60 °C)
11. Plot the data collected in labquest and find the slope (delta pressure/delta time). This value is the initial rate.
12. Once the initial rate has been found, calculate the activation energy for the KI-catalysed decomposition of  $\text{H}_2\text{O}_2$ .

### **Discussion:**

In this lab, various reactions were performed with different catalyst to determine and study the partial pressure and activation energy in each reaction. In part 2, the partial order of hydrogen peroxide was determined in it's decomposition with catalase which was obtained from blending lettuce in part 1 of the lab. The reaction yielded the products, oxygen and water, and the pressure of oxygen was measured as a function of time throughout all trials. In the table below, it shows the three different concentration of hydrogen peroxide which were used during the reaction and three trials were performed for each concentration in order to find the initial rate.

[ $\text{H}_2\text{O}_2$ ]	0.75%	0.5%	0.25%
Number of trails	Rate(KPa/s)		
1	0.67	0.195	0.0418
2	0.3275	0.162	0.0398
3	0.435	0.192	0.0462

Average rate	0.4775	0.183	0.0426
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Table 1. The concentration of  $H_2O_2$  versus initial rate of reaction with the presence of 20mL catalase solution.

We use the general expression of reaction rate,  $k=[H_2O_2]^x[catalase]^y$ , to calculate the partial order of reaction with respect to  $[H_2O_2]$ . We use  $[H_2O_2]=0.75\%$  and its corresponding rate 0.4775 as the first set of data;  $[H_2O_2]=0.5\%$  and  $k=0.183$  as the second set of data;  $[H_2O_2]=0.25\%$  and  $k=0.0426$  as the third set of data. To calculate the partial order, we need only 2 sets of data, so we choose 2 sets of data out of 3 each time, the results are listed in table 2.

	Partial order
$\text{Log}_{[H_2O_2]_1/[H_2O_2]_2} \text{Rate}_1/\text{Rate}_2$	2.365
$\text{Log}_{[H_2O_2]_1/[H_2O_2]_3} \text{Rate}_1/\text{Rate}_3$	2.200
$\text{Log}_{[H_2O_2]_2/[H_2O_2]_3} \text{Rate}_2/\text{Rate}_3$	2.102

Table 2. Partial order with respect to  $[H_2O_2]$  calculated using the first and second sets of data shown in row 2; the first and third sets of data shown in row 3; the second and the third sets of data shown in row 4.

We rounded the results in table 2 to their nearest integer, which is 2. Because the partial order can only be an integer. These results from calculations are slightly higher than 2. The possible sources of error come from trial 1 with  $[H_2O_2]=0.75\%$ , as observed in the first column of table 1, the rate obtained from that trial is 0.67. 0.67 is higher than average by 40.3%, so it is a major deviation from results of trial 2 and trial 3. Thus, the average rate corresponding to  $[H_2O_2]=0.75\%$  is higher than what it is supposed to be. Since  $\text{rate}_1$  is the antilogarithm in the logarithmic expression used to calculate partial rate in table 2 and because  $[H_2O_2]_1 > [H_2O_2]_2 > [H_2O_2]_3$ , so the base of all the logarithmic expressions in table 2 is greater than one, then this expression is strictly increasing, the increase of  $\text{rate}_1$  will cause the increase of partial order.

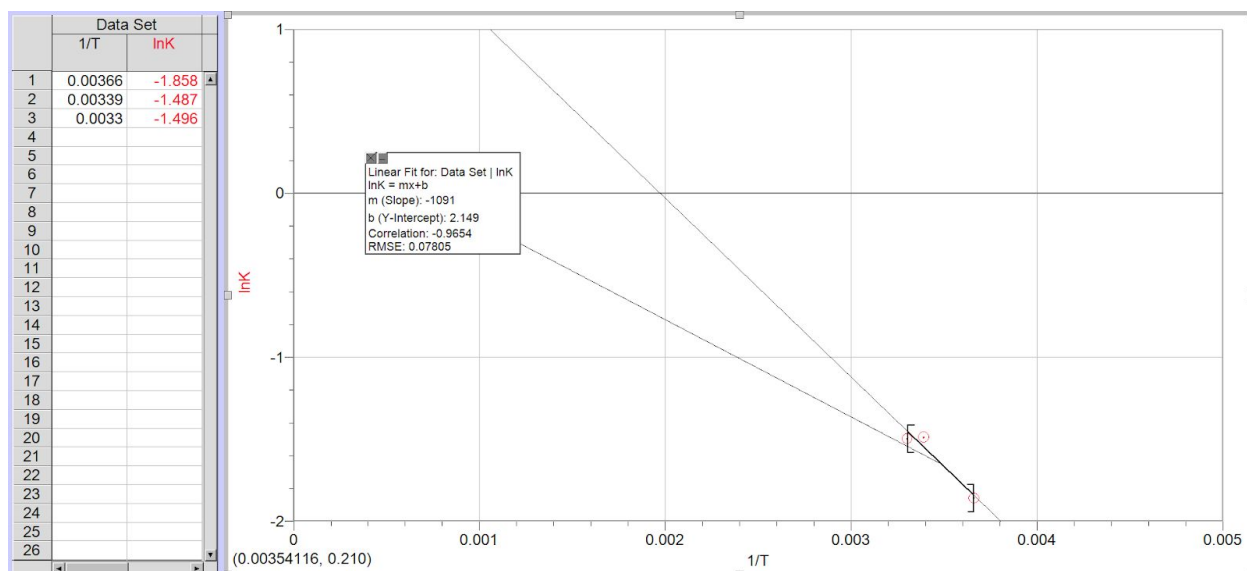
In part 3 of the lab, the purpose was to determine the activation energy by varying the temperature on the rate of reaction. The concentration of hydrogen peroxide and catalase was kept constant throughout the reaction. In the table below, it shows the three different temperatures in which the reactions occurred and three trials were performed for each temperature. Using the average of rates, it was converted to  $\ln(k)$  and the temperature was converted to  $1/T$ , which allowed for it to give a linear line in the graph plotted.

Temperature	273.15K	295.05K	303.15K
Number of trails	Rate(KPa/s)		
1	0.150	0.193	0.2097
2	0.150	0.294	0.234
3	0.168	0.192	0.2279
Average rate	0.156	0.226	0.224
ln(K)	-1.858	-1.487	-1.496
1/T	0.00366	0.00339	0.00330

Table 3. The initial rate of decomposition of H<sub>2</sub>O<sub>2</sub>, and its logarithm ln(K), with the presence of catalase in 3 different temperatures.

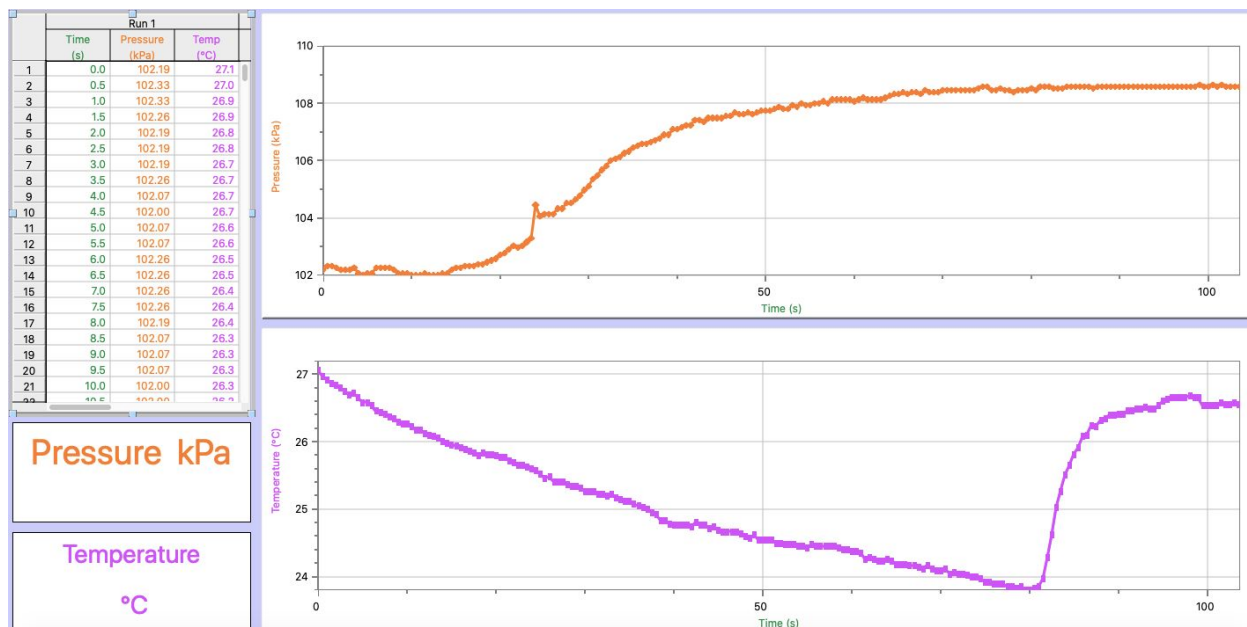
X coordinates (1/T)	Y coordinates (lnK)
0.00366	-1.858
0.00339	-1.487
0.00330	-1.496

Table 4. x and y coordinates for graph 1.



Graph 1. The linear fit line for 3 values of  $1/T$  in x coordinates and  $\ln K$  in y coordinates.

According to the Arrhenius equation (expressed in logarithmic form)<sup>[ref]</sup>, the slope of the line in graph 1 stands for  $-E_A/R$ . As shown in the box, the slope is  $-1091$ . Thus,  $E_A=131.224\text{J/mol}$ . From this graph, it can be seen that there is a negative, linear correlation between the inverse of temperature and the natural logarithm of the initial rate. This indicates that as temperature increases, the rate of the reaction would also increase because more molecules would gain the energy to surpass the activation energy.



Graph 2. data obtained at  $29^\circ\text{C}$  of catalase-catalyzed decomposition with hydrogen peroxide

From this graph, it can be seen that there is an exponential curve as the pressure increased with time progression. Furthermore, it was also noted that as the temperature increased, the rate of the reaction also increased which is indicated by a steeper curve in the graph. The relationship between the increased temperature, pressure, and the rate of reaction is described by the increase in the kinetic energy of the reaction. The result of the temperature change is an increased reaction energy, and a faster overall rate of reaction.

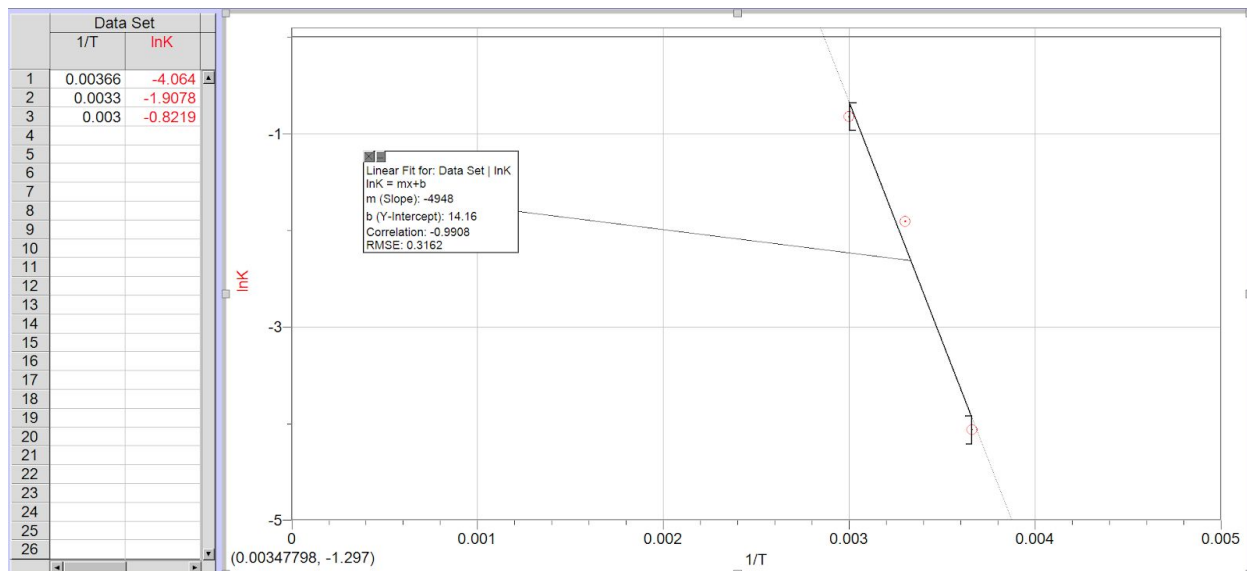
In graph 2, the slope calculated was 0.193, and was obtained from the first trial of the catalase-catalysed decomposition of hydrogen peroxide at 29°C. The calculated slope is representative of the rate of reaction outlined above and was used to identify the activation energy of the reaction.

The independent variable for this reaction was the temperature, which was around 29 °C throughout the reaction and the dependent variable was the pressure.

In part 4 of the lab, the purpose was same as part 3 where the activation energy was determined by varying the temperature on the rate of reaction. Only thing which was changed in this reaction was the catalyst used was potassium iodide whose concentration, along with the concentration of hydrogen peroxide kept constant. In the table below, it shows the three different temperatures in which the reactions occurred and three trials were performed for each temperature. Using the average of rates, it was converted to  $\ln(k)$  and the temperature was converted to  $1/T$ , which allowed for it to give a linear line in the graph plotted.

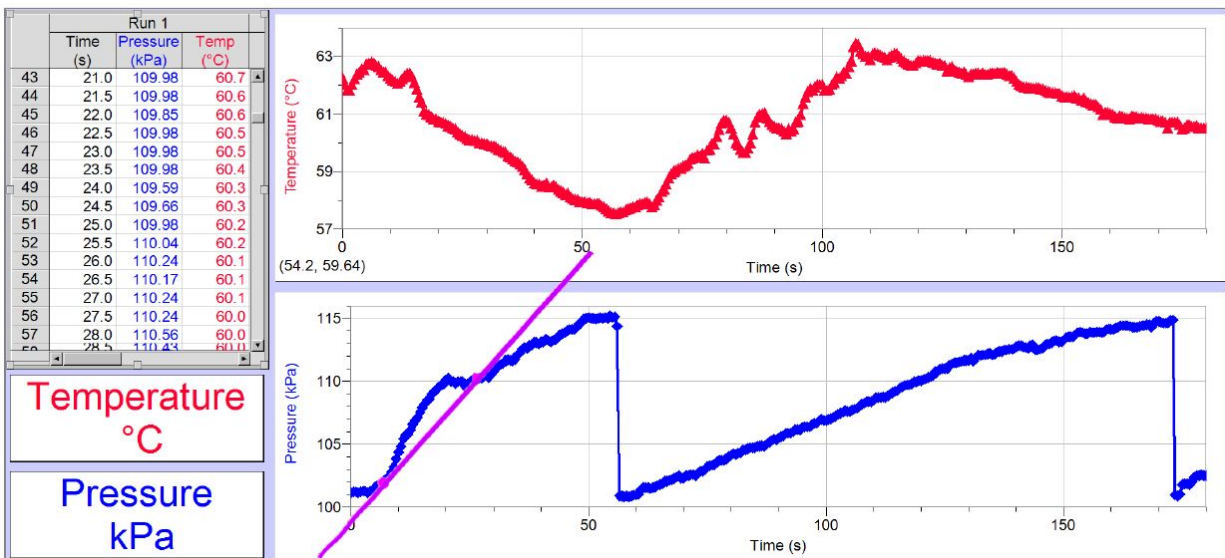
Temperature	273.15K	303.15K	333.15K
Number of trails	Rate(KPa/s)		
1	0.01183	0.1867	0.4359
2	0.0191	0.1114	0.4941
3	0.0206	0.1471	0.389
Average rate	0.01718	0.1484	0.4396
$\ln(K)$	-4.0640	-1.9078	-0.8219
$1/T$	0.00366	0.00330	0.00300

Table 5. The initial rate of decomposing  $\text{H}_2\text{O}_2$  and its logarithm with the presence of KI in 3 different temperatures.



Graph 2. The linear fit line for 3 values of  $1/T$  (x) and  $\ln K$  (y) from table 5.

Since the slope of the linear fit line is  $-4948$ ,  $E_A = 595.141 \text{ J/mol}$ .



Graph 3. Data obtained from calculating the slope of the points,  $(t=6.5\text{s}, P=101.74)$  and  $(t=26\text{s}, P=110.24)$ , for trial 1 of the  $60^\circ\text{C}$  KI-catalyzed decomposition.

The time in which the reaction takes to initiate an increase in pressure decreases as the temperature of the system is increased. This was seen dramatically when the temperature of the reaction was increased from 30 °C to 60 °C, as the pressure increase was seen almost immediately with the large increase in temperature. The relationship between the increased temperature, pressure, and the rate of reaction is described by the increase in the kinetic energy of the reaction. The result of the temperature change is an increased reaction energy, and a faster overall rate of reaction.

The slope of the increase in pressure, which is represented in graph 3, is 0.4359, and was obtained from the first trial of the KI-catalysed decomposition of hydrogen peroxide at 60 °C. The calculated slope is representative of the rate of reaction outlined above and was used to identify the activation energy of the reaction.

Compared to the activation energy without catalyst, which is 75000J/mol, the catalase lowered the  $E_a$  significantly. This was expected as catalyst increase the rate of reaction and lower the activation energy.

The independent variable for this reaction is temperature, which stayed around 60 °C throughout the reaction. The dependent variable is pressure, which changed. The temperature and pressure are measured with a thermometer and gas pressure sensor respectively. A steady set of measurements through 100 seconds will be enough, for 3 trials at each temperature. The measurements were distributed among the group by temperature, 20, 30 and 60 °C. To determine the activation energy of the reaction, the temperature and rate constant are required.

The design of the lab yielded results that correlated well with expectations. The effect of temperature on the rate of reaction was clear as well as the difference between the two catalysts. Due to the reaction occurring in a closed environment within the erlenmeyer flask the amount of  $O_2$  and  $H_2$  released can be tabulated by the gas pressure. The temperature of the reaction was controlled by the water bath. This ensured that the temperature of the reaction could be monitored throughout.

A possible source of error was the dilution of  $H_2O_2$ , everytime the  $H_2O_2$  was diluted at the beginning of each trial there was a margin of error in measurement. This could have greatly affected the rate of reaction. If this experiment were to be done again it would be important to have a large quantity of 0.5M  $H_2O_2$ . In addition, performing the experiment at precise temperatures was difficult with the equipment available. It was good enough to show the relationship between rate and temperature but the trials were inconsistent and likely affected the calculations of activation energy.

Some limitations which were present in this lab were that the beaker in which the reaction occurred was not completely shut off as some air was able to escape from opening in the lid and this would have affected the overall pressure which was obtained from the reaction.

### **Conclusion:**

Using the average rates from the 3 concentrations of  $\text{H}_2\text{O}_2$  the partial order was calculated using the general expression for reaction rate. The partial order with respect to the concentration of  $\text{H}_2\text{O}_2$  is 2. With catalase,  $E_A=131.224\text{J/mol}$ . With KI catalyst,  $E_A=595.141\text{J/mol}$ .

Comparing the rates of reaction from the presence of potassium iodide or spinach catalyst, it seems that the spinach catalyst gives off better yields in terms of reaction rates. When compared by the reaction rates calculated with the use of potassium iodide catalyst, there was a great difference in between them. In addition, the spinach catalyst results indicate that as temperature increases, the rate of the reaction would also increase because more molecules would gain the energy to surpass the activation energy which is sufficiently lower than normal, much greater than the reactions that used the potassium iodide catalyst. Furthermore, since the activation energy without the presence of catalyst is  $75,000\text{J/mol}$  and the activation energy for spinach catalyst is  $131.224\text{J/mol}$ , there is a large difference between the values. Moreover, the activation energy for spinach catalyst is less than the activation energy for potassium iodide catalyst which is  $595.141\text{J/mol}$ , meaning that less energy is required to initiate the reaction and therefore allow for a faster reaction compared to the potassium iodide catalyst trials.

### **Reference(s):**

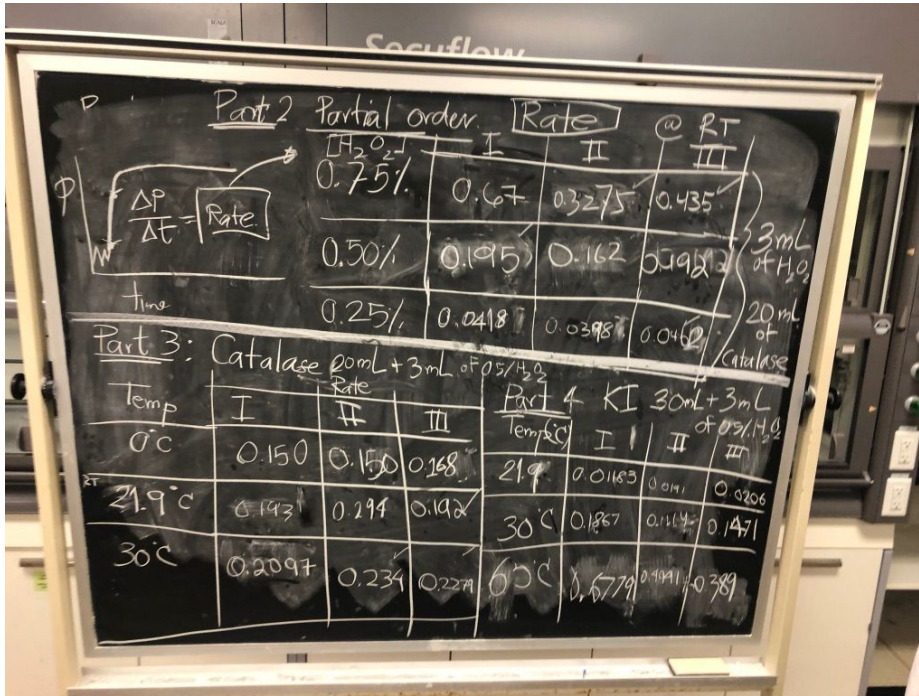
“A KINETIC STUDY: CATALASE”, General Chemistry Laboratory Manual, Dr. Rashmi Venkateswaran, 2019.

[ref for log expression of Arrhenius equation in discussion by Chunyang Lu

Bohn MA. The connection between the parameters of WLF equation and of Arrhenius equation. *Propellants Explos Pyrotech.* 2019;44:697.

**Appendix:**

**Raw Data:**



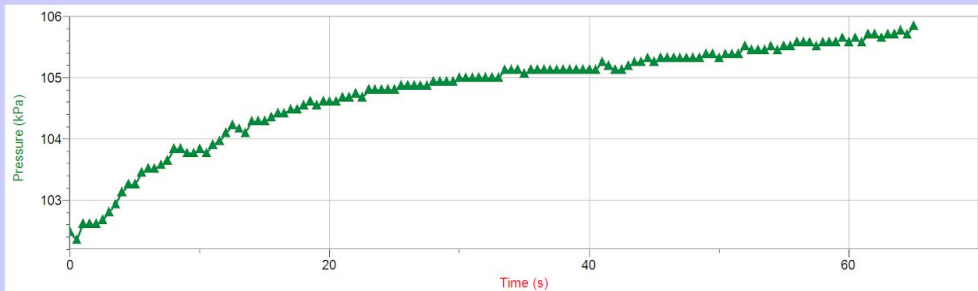
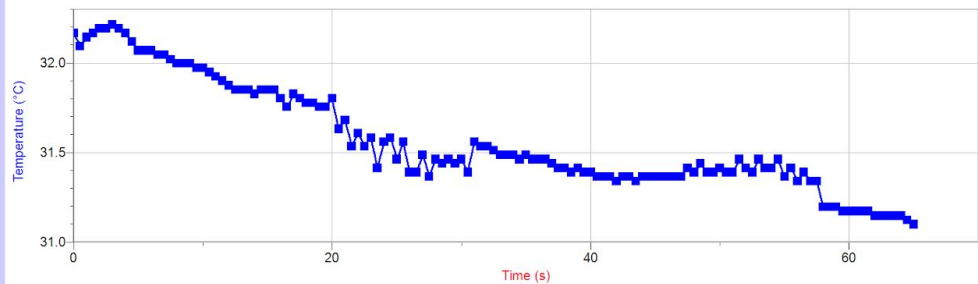
**Additional Graphs:**

Three trials for: KI Catalyst Temperature vs Pressure at 30°C

Run 1			
	Time (s)	Temp (°C)	Pressure (kPa)
1	0.0	32.2	102.47
2	0.5	32.1	102.35
3	1.0	32.1	102.60
4	1.5	32.2	102.60
5	2.0	32.2	102.60
6	2.5	32.2	102.67
7	3.0	32.2	102.80
8	3.5	32.2	102.93
9	4.0	32.2	103.12
10	4.5	32.1	103.25
11	5.0	32.1	103.25
12	5.5	32.1	103.44
13	6.0	32.1	103.51
14	6.5	32.0	103.51
15	7.0	32.0	103.57
16	7.5	32.0	103.64
17	8.0	32.0	103.83
18	8.5	32.0	103.83
19	9.0	32.0	103.76
20	9.5	32.0	103.76

Temperature °C

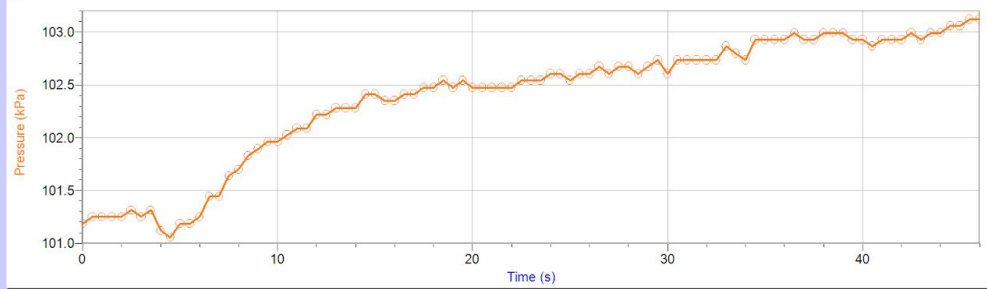
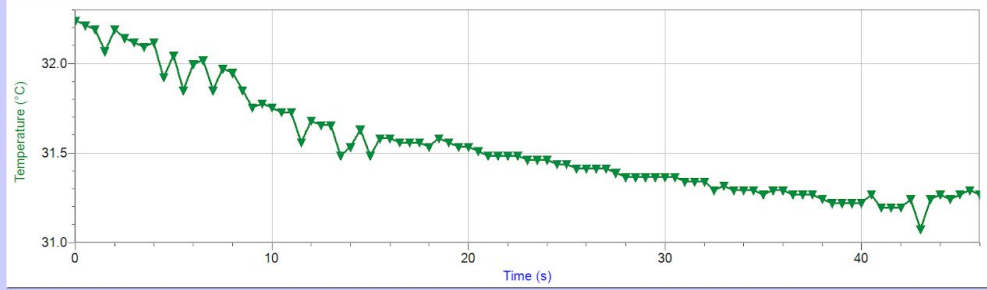
Pressure kPa



Run 1			
	Time (s)	Temp (°C)	Pressure (kPa)
1	0.0	32.2	101.19
2	0.5	32.2	101.25
3	1.0	32.2	101.25
4	1.5	32.1	101.25
5	2.0	32.2	101.25
6	2.5	32.1	101.31
7	3.0	32.1	101.25
8	3.5	32.1	101.31
9	4.0	32.1	101.12
10	4.5	31.9	101.06
11	5.0	32.0	101.19
12	5.5	31.9	101.19
13	6.0	32.0	101.25
14	6.5	32.0	101.44
15	7.0	31.9	101.44
16	7.5	32.0	101.64
17	8.0	31.9	101.70
18	8.5	31.9	101.83
19	9.0	31.8	101.89
20	9.5	31.8	101.96

Temperature  
°C

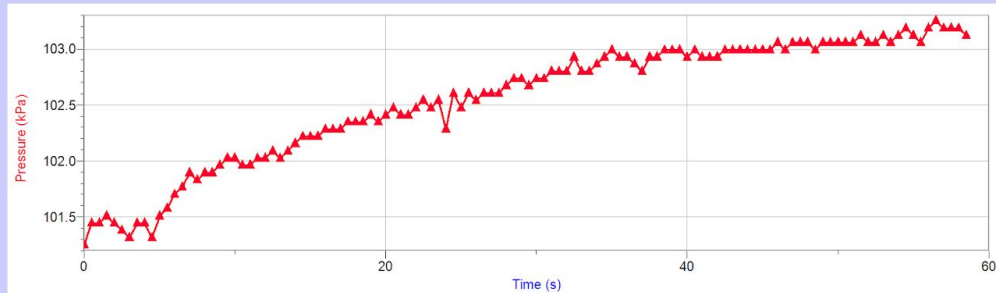
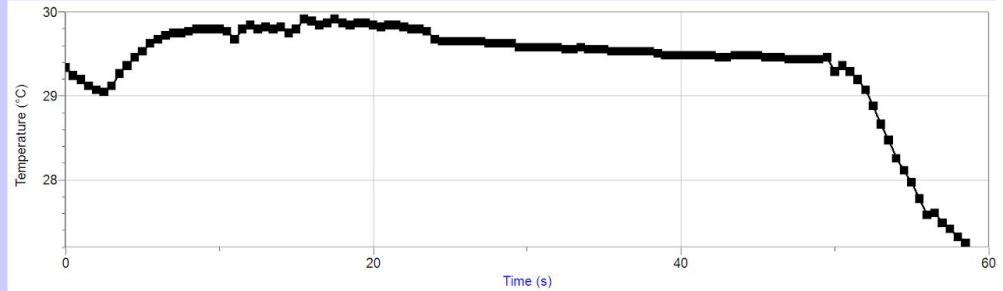
Pressure  
kPa



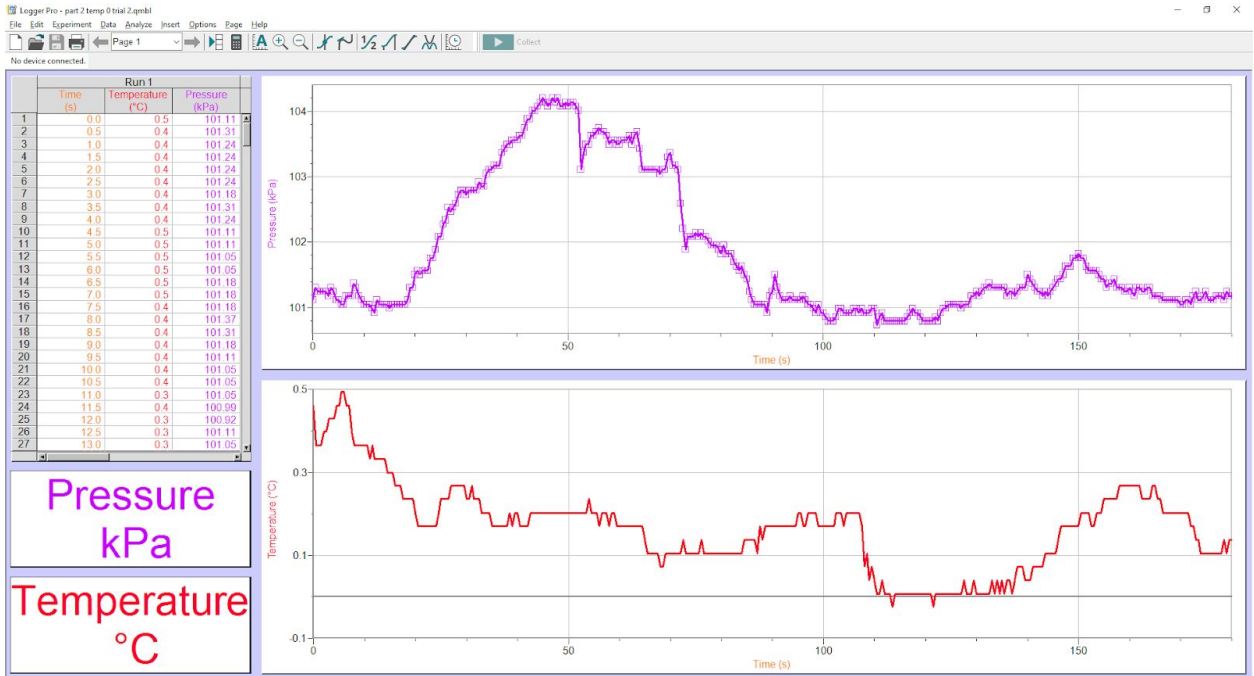
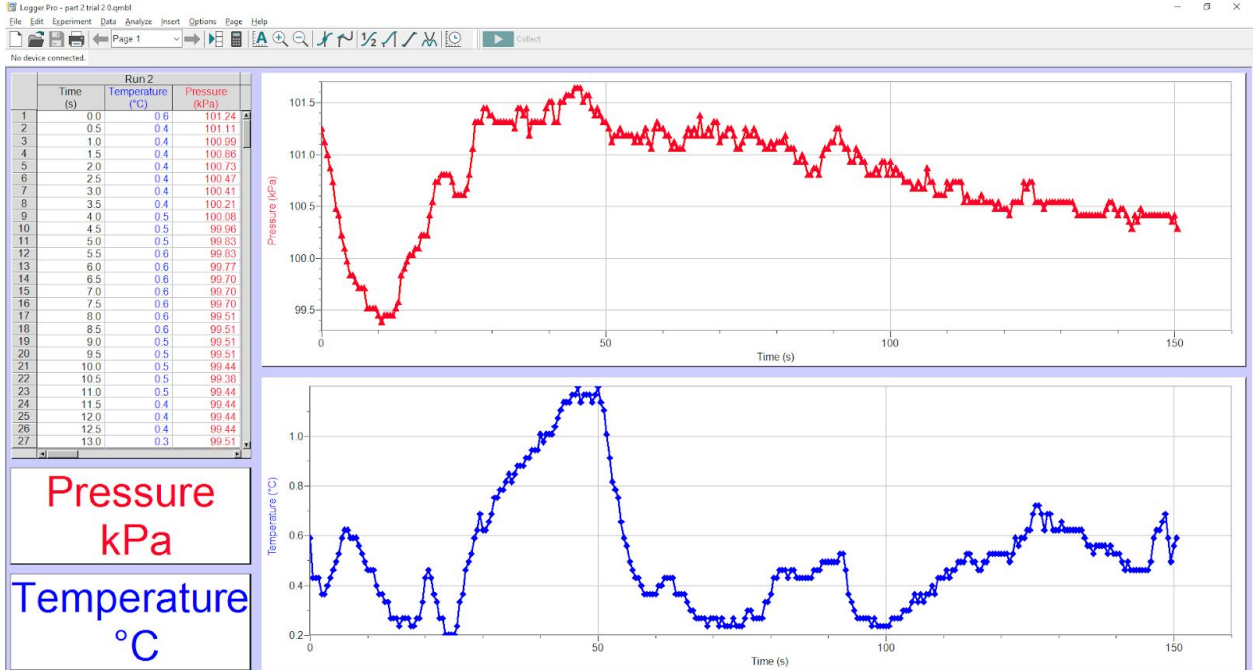
Run 1			
	Time (s)	Temp (°C)	Pressure (kPa)
1	0.0	29.3	101.25
2	0.5	29.2	101.44
3	1.0	29.2	101.44
4	1.5	29.1	101.51
5	2.0	29.1	101.44
6	2.5	29.0	101.38
7	3.0	29.1	101.31
8	3.5	29.3	101.44
9	4.0	29.4	101.44
10	4.5	29.5	101.31
11	5.0	29.5	101.51
12	5.5	29.6	101.57
13	6.0	29.7	101.70
14	6.5	29.7	101.77
15	7.0	29.7	101.89
16	7.5	29.7	101.83
17	8.0	29.8	101.89
18	8.5	29.8	101.89
19	9.0	29.8	101.96
20	9.5	29.8	102.02

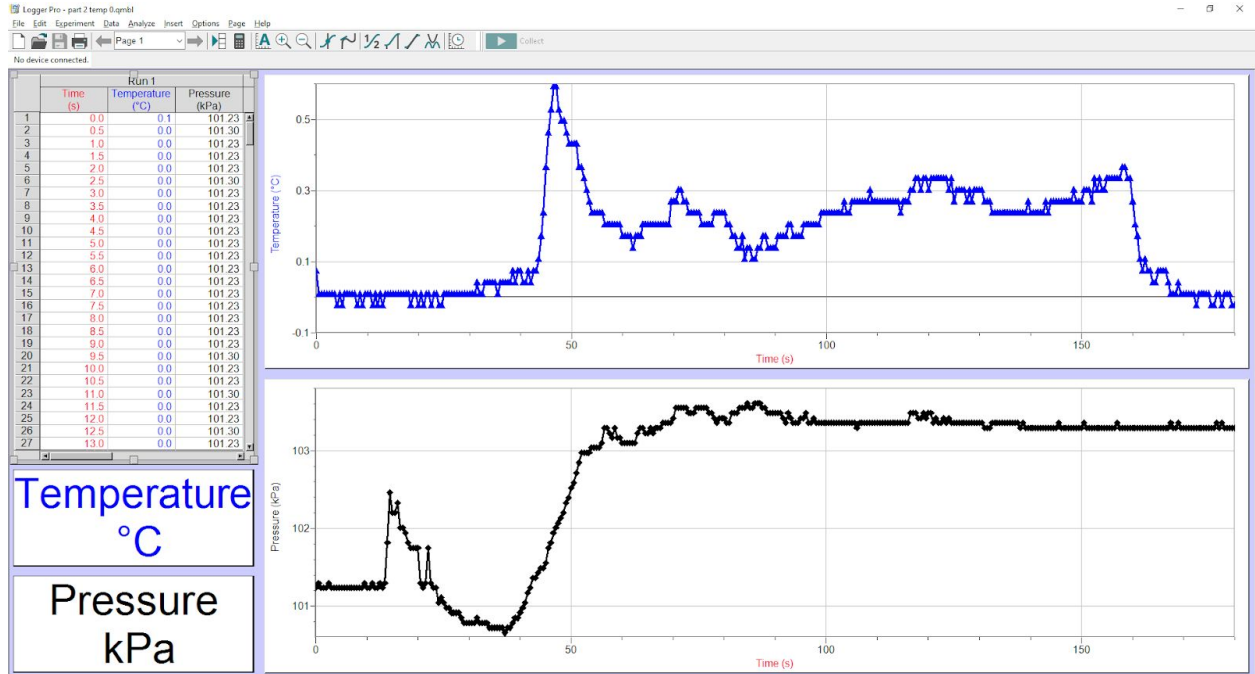
Temperature  
°C

Pressure  
kPa

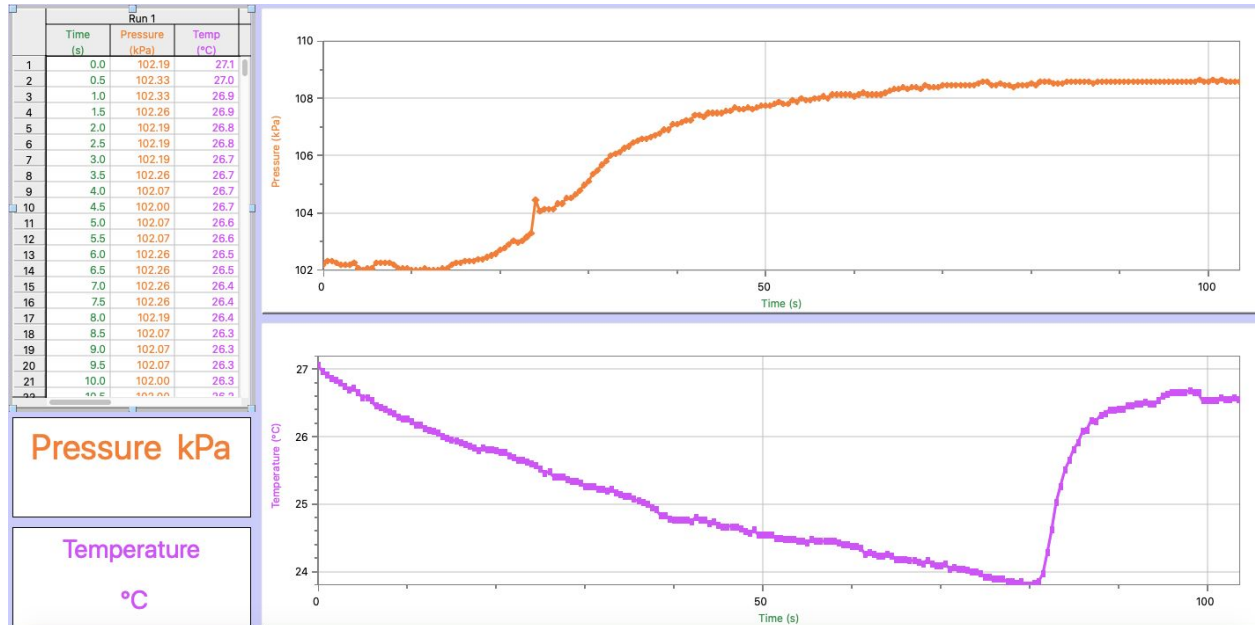


Three trials for: Spinach Catalyst Temperature vs Pressure at 0.0°C

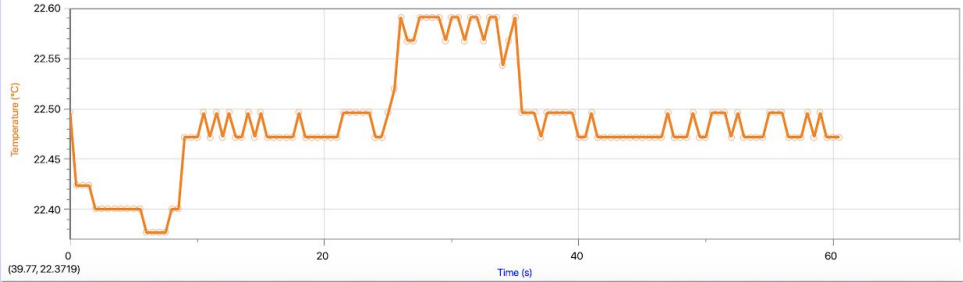
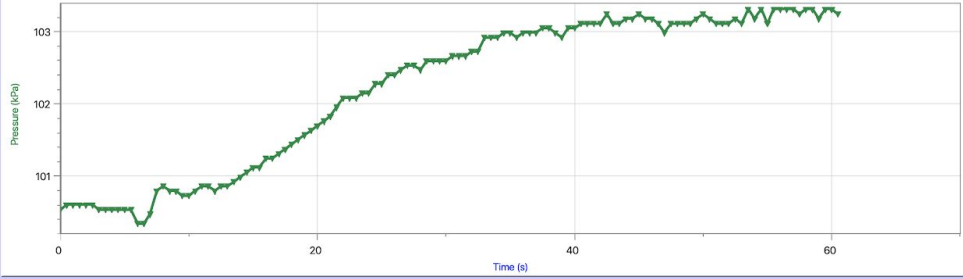




Three trials for: Spinach Catalyst Temperature vs Pressure at 21.9°C



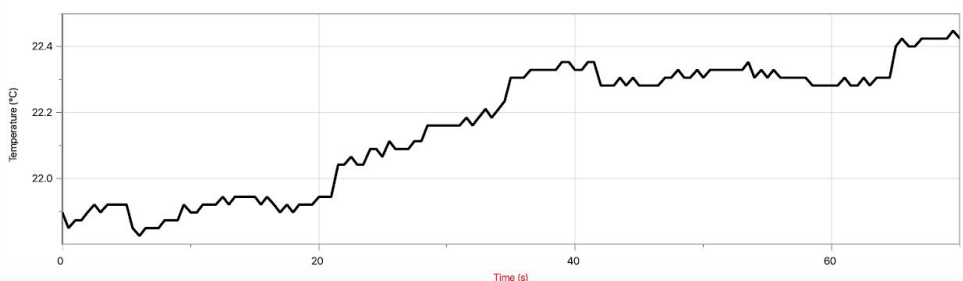
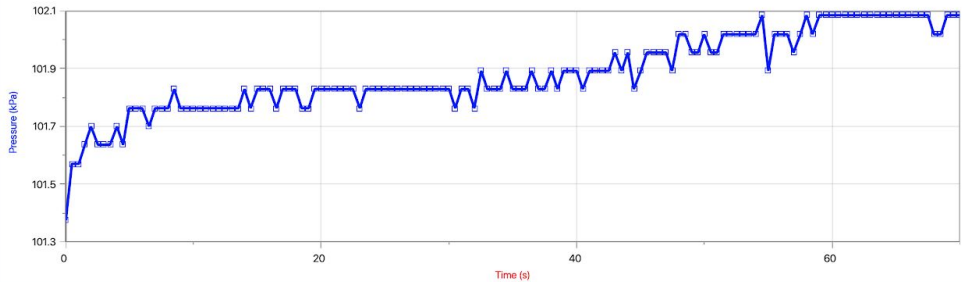
Run 5			
	Time (s)	Pressure (kPa)	Temperature (°C)
1	0.0	100.54	22.5
2	0.5	100.60	22.4
3	1.0	100.60	22.4
4	1.5	100.60	22.4
5	2.0	100.60	22.4
6	2.5	100.60	22.4
7	3.0	100.54	22.4
8	3.5	100.54	22.4
9	4.0	100.54	22.4
10	4.5	100.54	22.4
11	5.0	100.54	22.4
12	5.5	100.54	22.4
13	6.0	100.35	22.4
14	6.5	100.35	22.4
15	7.0	100.48	22.4
16	7.5	100.80	22.4
17	8.0	100.86	22.4
18	8.5	100.80	22.4
19	9.0	100.80	22.5
20	9.5	100.73	22.5
21	10.0	100.73	22.5
22	10.5	100.80	22.5



Pressure kPa

Temperature  
°C

Run 3			
	Time (s)	Pressure (kPa)	Temperature (°C)
1	0.0	101.38	21.9
2	0.5	101.57	21.8
3	1.0	101.57	21.9
4	1.5	101.64	21.9
5	2.0	101.70	21.9
6	2.5	101.64	21.9
7	3.0	101.64	21.9
8	3.5	101.64	21.9
9	4.0	101.70	21.9
10	4.5	101.64	21.9
11	5.0	101.76	21.9
12	5.5	101.76	21.8
13	6.0	101.76	21.8
14	6.5	101.70	21.8
15	7.0	101.76	21.8
16	7.5	101.76	21.8
17	8.0	101.76	21.9
18	8.5	101.83	21.9
19	9.0	101.76	21.9
20	9.5	101.76	21.9
21	10.0	101.76	21.9
22	10.5	101.76	21.9



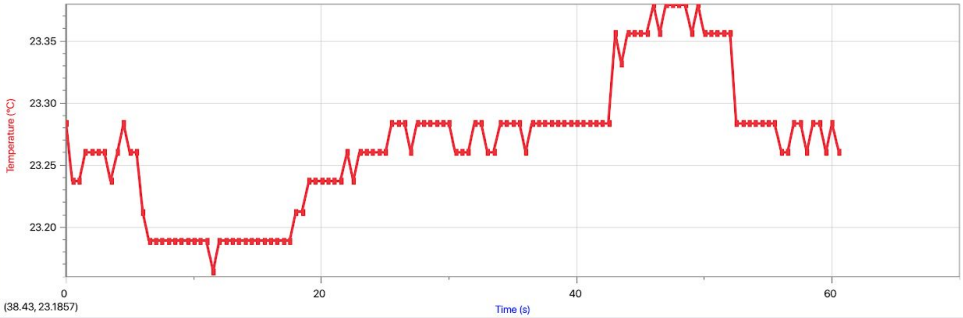
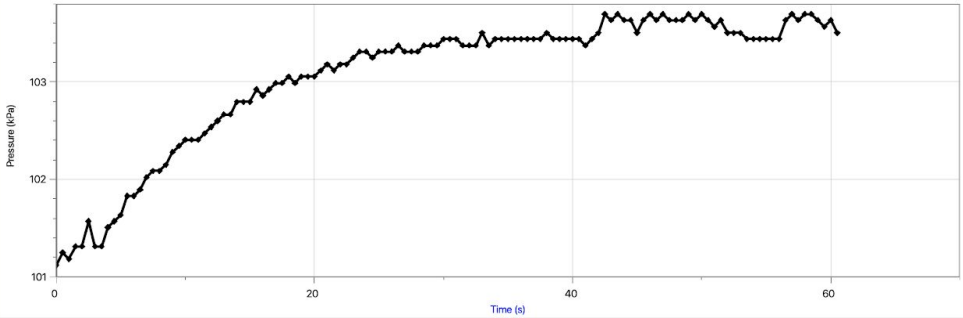
Pressure kPa

Temperature  
°C

Run 4			
	Time (s)	Pressure (kPa)	Temperature (°C)
1	0.0	10112	23.3
2	0.5	10125	23.2
3	1.0	10118	23.2
4	1.5	10131	23.3
5	2.0	10131	23.3
6	2.5	10157	23.3
7	3.0	10131	23.3
8	3.5	10131	23.2
9	4.0	10151	23.3
10	4.5	10157	23.3
11	5.0	10164	23.3
12	5.5	10183	23.3
13	6.0	10183	23.2
14	6.5	10189	23.2
15	7.0	10202	23.2
16	7.5	10209	23.2
17	8.0	10209	23.2
18	8.5	10215	23.2
19	9.0	10228	23.2
20	9.5	10234	23.2
21	10.0	10241	23.2
22	10.5	10241	23.2
23	11.0	10241	23.2
24	11.5	10247	23.2
25	12.0	10254	23.2

Pressure kPa

Temperature  
°C

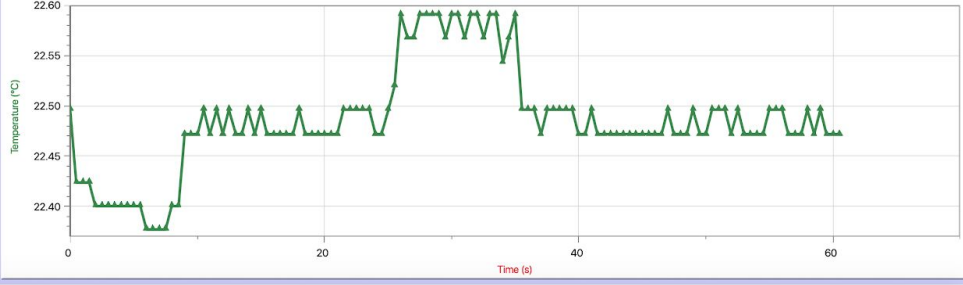
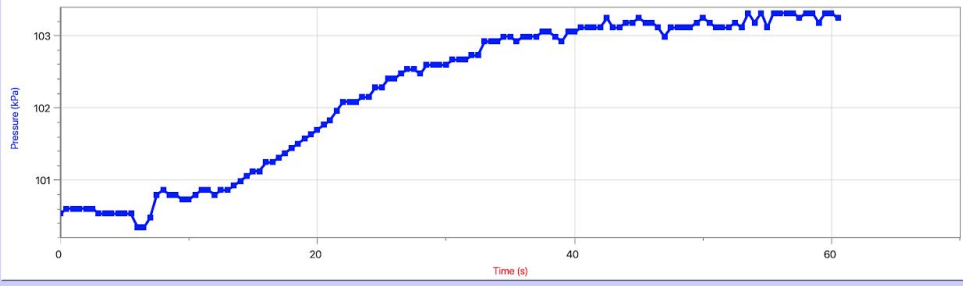


(38.43, 23.1857)

Run 5			
	Time (s)	Pressure (kPa)	Temperature (°C)
1	0.0	100.54	22.5
2	0.5	100.60	22.4
3	1.0	100.60	22.4
4	1.5	100.60	22.4
5	2.0	100.60	22.4
6	2.5	100.60	22.4
7	3.0	100.54	22.4
8	3.5	100.54	22.4
9	4.0	100.54	22.4
10	4.5	100.54	22.4
11	5.0	100.54	22.4
12	5.5	100.54	22.4
13	6.0	100.35	22.4
14	6.5	100.35	22.4
15	7.0	100.48	22.4
16	7.5	100.80	22.4
17	8.0	100.86	22.4
18	8.5	100.80	22.4
19	9.0	100.80	22.5
20	9.5	100.73	22.5
21	10.0	100.73	22.5
22	10.5	100.80	22.5

Pressure kPa

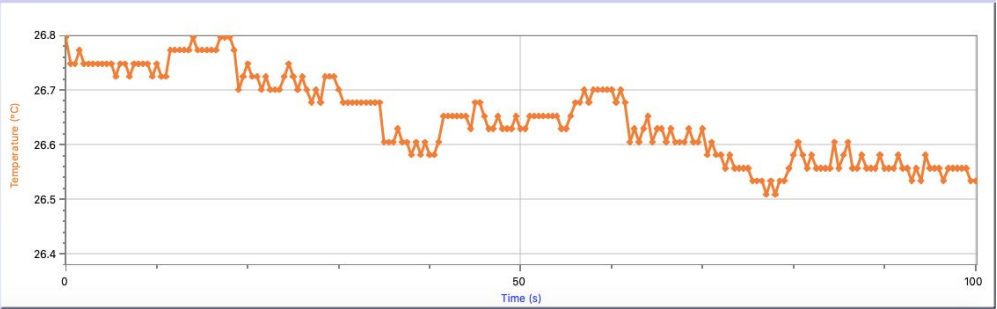
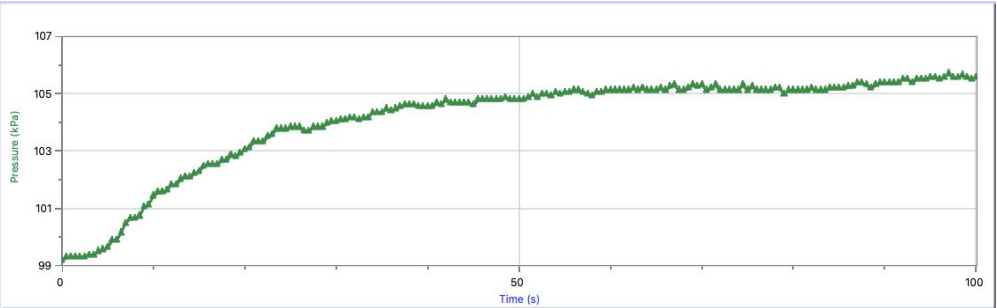
Temperature  
°C



Run 1			
Time (s)	Pressure (kPa)	Temp (°C)	
1	0.0	99.16	26.8
2	0.5	99.30	26.7
3	1.0	99.30	26.7
4	1.5	99.30	26.8
5	2.0	99.30	26.7
6	2.5	99.30	26.7
7	3.0	99.36	26.7
8	3.5	99.36	26.7
9	4.0	99.49	26.7
10	4.5	99.55	26.7
11	5.0	99.62	26.7
12	5.5	99.88	26.7
13	6.0	99.88	26.7
14	6.5	100.13	26.7
15	7.0	100.45	26.7
16	7.5	100.65	26.7
17	8.0	100.65	26.7
18	8.5	100.71	26.7
19	9.0	101.03	26.7
20	9.5	101.10	26.7
21	10.0	101.42	26.7

Pressure kPa

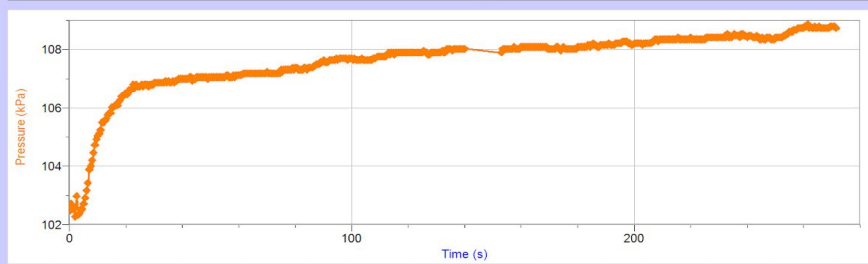
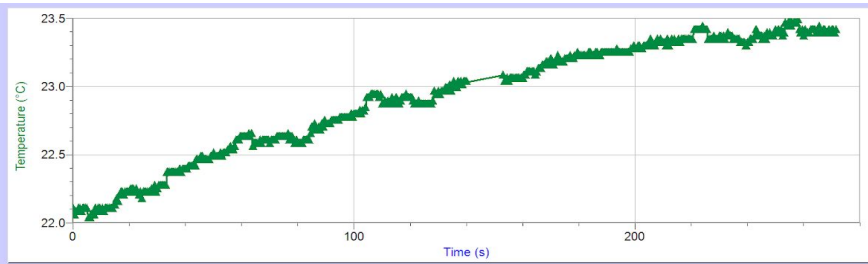
Temperature  
°C



Run 1			
Time (s)	Temp (°C)	Pressure (kPa)	
1	0.0	22.1	102.46
2	0.5	22.1	102.72
3	1.0	22.1	102.66
4	1.5	22.1	102.53
5	2.0	22.1	102.27
6	2.5	22.1	102.98
7	3.0	22.1	102.33
8	3.5	22.1	102.40
9	4.0	22.1	102.53
10	4.5	22.1	102.53
11	5.0	22.1	102.72
12	5.5	22.0	102.92
13	6.0	22.0	103.17
14	6.5	22.1	103.43
15	7.0	22.1	103.89
16	7.5	22.1	104.01
17	8.0	22.1	104.21
18	8.5	22.1	104.47
19	9.0	22.1	104.73
20	9.5	22.1	104.92

Temperature  
°C

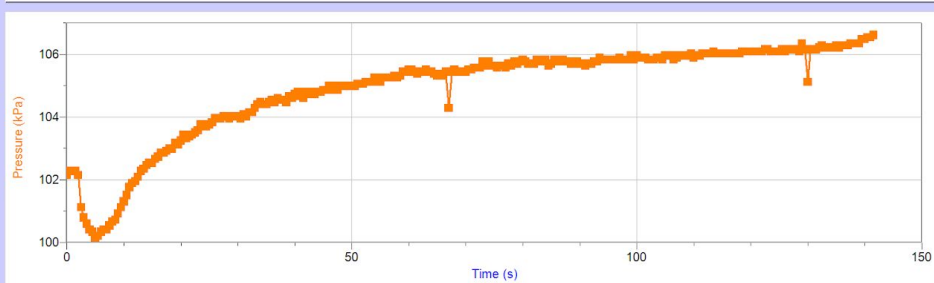
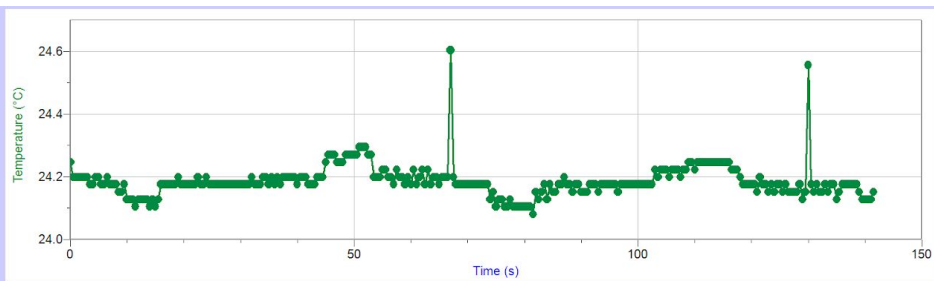
Pressure  
kPa



Run 1			
Time (s)	Temp (°C)	Pressure (kPa)	
1	0.0	24.2	102.14
2	0.5	24.2	102.27
3	1.0	24.2	102.27
4	1.5	24.2	102.27
5	2.0	24.2	102.14
6	2.5	24.2	101.11
7	3.0	24.2	100.78
8	3.5	24.2	100.59
9	4.0	24.2	100.39
10	4.5	24.2	100.33
11	5.0	24.2	100.14
12	5.5	24.2	100.20
13	6.0	24.2	100.33
14	6.5	24.2	100.39
15	7.0	24.2	100.39
16	7.5	24.2	100.53
17	8.0	24.2	100.66
18	8.5	24.2	100.72
19	9.0	24.2	100.91
20	9.5	24.2	101.11

Temperature  
°C

Pressure  
kPa

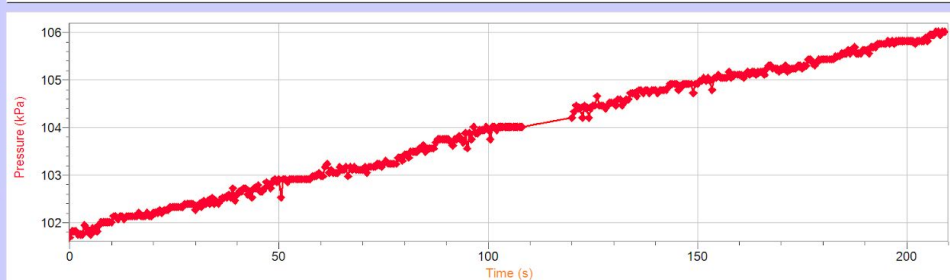
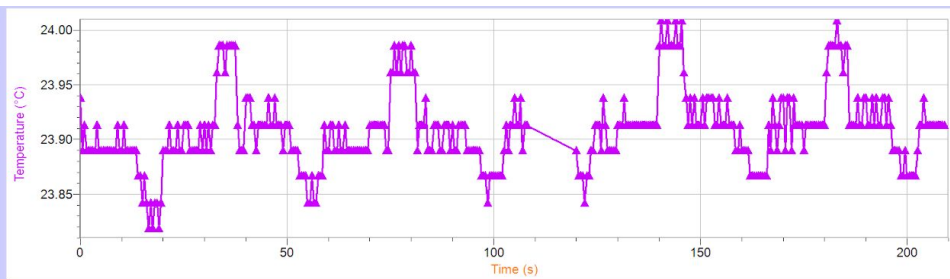


**Part 4 Trial 2 21.9°C**

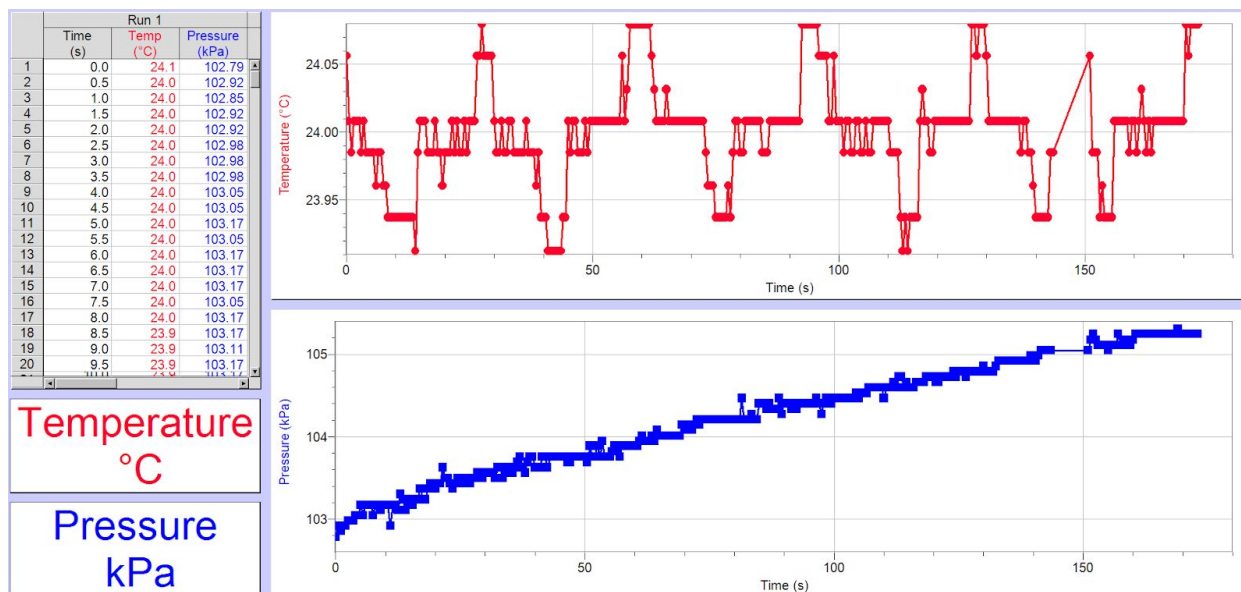
Run 1			
Time (s)	Temp (°C)	Pressure (kPa)	
1	0.0	23.9	101.69
2	0.5	23.9	101.82
3	1.0	23.9	101.82
4	1.5	23.9	101.82
5	2.0	23.9	101.75
6	2.5	23.9	101.75
7	3.0	23.9	101.75
8	3.5	23.9	101.95
9	4.0	23.9	101.88
10	4.5	23.9	101.82
11	5.0	23.9	101.75
12	5.5	23.9	101.86
13	6.0	23.9	101.88
14	6.5	23.9	101.82
15	7.0	23.9	101.95
16	7.5	23.9	102.01
17	8.0	23.9	102.01
18	8.5	23.9	102.01
19	9.0	23.9	102.01
20	9.5	23.9	102.01

Temperature  
°C

Pressure  
kPa



**Part 4 Trial 3 21.9°C**



**Sample Calculations:**

**Part 2 [H2O2]% is 0.75%**

**Trial 1 - Erin and Adrian**

**P1 = 102.83 V1 = 8**

**P2 = 106.18 V2 = 13**

**Rate = P2 - P1/ V2 - V1**

**Rate = 106.18 - 102.83 / 13-8**

**Rate = 0.67**

**Trial 2**

**Trial 3**

**Part 2 [H2O2]% is 0.50%**

**Trial 1 - Erin and Adrian**

**P1 = 102.50 V1 = 6**

**P2 = 104.15 V2 = 10**

**Rate = P2 - P1/ V2 - V1**

**Rate = 104.15 - 102.50 / 10-6**

**Rate = 0.3875**

**Trial 2 - Humza and Ryan**

**P1 = 101.88**

**T1 = 103.76**

**P2 = 11.5**

**T2 = 23.5**

**Rate = P1 - P2/T2 - T1**

**Rate = 1.88/12**

**Rate = 0.162**

**Trial 3**

**Part 2 [H202] % is 0.25%**

**Trial 1-**

**P1=101.12**

**T1=0**

**P2=103.63**

**T2= 60**

**Rate=P1-P2/T2-T1**

**Rate=2.51/60**

**Rate=0.04183**

**Trial 2-**

**P1=101.31**

**T1=0**

**P2=103.70**

**T2=60**

**Rate=2.39/60**

**Rate=0.03983**

**Trial 3-**

**P1=100.54**

**T1=0**

**P2=103.31**

**T2=60**

Rate=2.77/60

Rate=0.04616

Part 3 Ea of catalase Temp = 0°C

Trial 1 - Erin and Adrian

P1 = 100.72 V1 = 38

P2 = 102.97 V2 = 53

Rate = P2 - P1 / V2 - V1

Rate = 2.25 / 15

Rate = 0.15

Trial 2

Trial 3

Part 3 Ea of catalase Temp = Room temp 21.9°C

Trial 1 - Humza and ryan

PF = 106.6

PI = 102.46

T - 21.9°C

Rate = P1 - P2 / T2 - T1

Rate = 106.6 - 102.46 / 21.9

Rate = 0.193

Trial 2 - Humza And Ryan

PF = 104.98

PI = 102.33

T - 21.9°C

Rate = P1 - P2 / T2 - T1

Rate = 104.98 - 102.33 / 21.9

Rate = 0.294

Trial 3-

PF= 102.09

PI=101.38

T=21.9°C

Rate= 0.71/21.9

Rate=0.0324

**Part 3 Ea of catalase Temp = 30°C**

**Trial 1**

**P1 = 106.69 V1 = 20**

**P2 = 112.88 V2 = 47.5**

**Rate = P2 - P1 / V2 - V1**

**Rate = 6.19 / 27.15**

**Rate = 0.2279**

**Trial 2**

**Trial 3**

**Part 4 Ea of KI Temp = 21.9°C**

**Trial 1**

**Trial 2 - Humza and Ryan**

**P1 = 102.21**

**Time 1 = 14.5s**

**P2 = 105.82**

**Time 2 = 204s**

**Rate = P1 - P2 / T2 - T1**

**Rate = 3.61 / 189.5**

**Rate = 0.0191**

**Trial 3 - Humza and Ryan**

**P1 = 103.11**

**Time 1 = 12.05s**

**P2 = 103.76**

**Time 2 = 43.5s**

**Rate = P1 - P2 / T2 - T1**

**Rate = 0.65 / 31.5**

**Rate = 0.0206**

**Part 4 Ea of KI Temp = 30°C (Ross, Bradley)**

**Trial 1: rate=  $P1 - P2/T1-T2 = (103.44 - 102.6)/(5.5-1) = 0.18670$**

**Trial 2: rate=  $(102.22 - 101.19) / (12 - 5) = 0.14710$**

**Trial 3: rate=  $(101.88 - 101.44)/(7.5 - 4) = 0.11143$**

**Part 4 Ea of KI Temp = 60°C**

**Trial 1  $(110.24-101.74)/(26-6.5)=0.4359$**

**Trial 2**

**Trial 3**

**Assessment Criteria:**

**Assessment Criteria for Planning A Kinetic Study: Catalase**  
(print and paste in your lab notebook before coming to lab)

TA Name:		Names of Students in Group:	Cameron Jackson Maya Williams Hamza Mohammed Samiyoun (Elyan) Chen
		Date:	
Criteria:	Marks Possible	Assessment	
		Self	TA
1. Identify the problem and state it clearly in a way that can be tested.	1	1	0.5
2. Use proper apparatus, techniques and safety precautions.	0.5	0.5	0.5
3. Plan to vary only one independent variable at a time.	1	1	0.5
4. Controls on other variables are clearly stated.	0.5	0.5	0
5. Measurement errors are minimized by appropriate procedures or apparatus.	0.5	0.5	0.5
6. No invalid assumptions are made.	0.5	0.5	0.25
7. Reagents that need accurate measurement are identified.	0.5	0.5	0.25
8. Lab trials and repeats are clearly stated.	0.5	0.5	0.5
<b>TOTAL:</b>	<b>5</b>	<b>5</b>	<b>3.5</b>

Note: This grade will count towards your prelab grade.