A Kinetic Study: Catalase

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CHM1311, Section Z10

Demonstrator:

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INTRODUCTION:

This lab provides students with an investigation into the effectiveness of different catalysts in reducing the activation energy of reactions, specifically the breakdown of hydrogen peroxide. With that in mind, important topics to understand before the conduction of the experiment, or any similar one, are rate laws of various orders, energy diagrams regarding the activation of a reaction, and the effect of catalysis on said graphs.

To understand chemical reactions it is imperative to understand chemical kinetics. This branch of chemistry, which deals with the rate of chemical reactions. the factors affecting the rate of reactions and different reaction mechanisms, is the primary unit of study for which this lab will demonstrate**.**

Rate of Reaction: Rate of a chemical reaction is the change in the concentration of the reactants or products per unit time. It is expressed in mol L^{-1} .

R (reactants)⟶P(products)

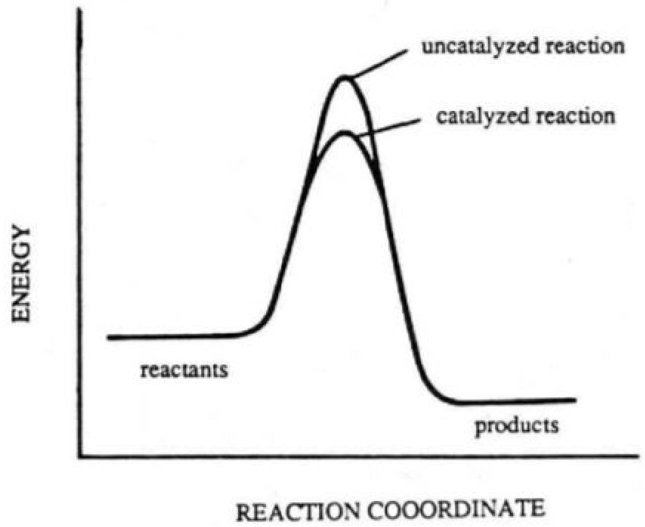
$$\text{Then the rate of disappearance of reactants} = -\frac{d[R]}{dt}$$

$$\text{rate of appearance of products} = \frac{d[P]}{dt}$$

Regarding factors affecting the rate of reaction, it is important to understand that various components will affect the speed at which a reaction takes place, with varying impact. The list below details six points which impact the rate of a reaction (in no particular order).

1. Concentration of reactants
2. Temperature
3. Surface area of reactant
4. Radiations and catalysis
5. Pressure of gas
6. Nature of the reactants

For this lab, it is most important to take into account the concentration of reactants, temperature of reaction, and the catalysis of reactions, with this final option having the greatest importance. To fully grasp this concept, one must first answer two questions: what is a catalyst, and how does it affect the rate of reaction? A catalyst, according to *Catalysis: New Reaction Pathways, Not Just a Lowering of the Activation Energy* by [author], is “a substance that is not involved in the stoichiometry of the reaction but that increases the reaction rate” **[1]**. The activation energy is defined as the energy required for a reaction to proceed. As shown by the image below, the catalyst causes the energy path to require less total energy for the reaction to proceed, thus allowing for the reaction to progress more quickly.



[1]

Rate Law Expressions and Rate Constant:

The rate law of a reaction is an expression which relates the **rate of a reaction** to the **concentrations** of the reactants via a **rate constant (k)**.

For a chemical reaction,

$$A + B \rightarrow \text{Products}$$

Therefore, the rate of this reaction is the following expression: **Rate =** $\mathbf{k[A]^a\ [B]^b}$

Where *a* and *b* are partial orders. Therefore, a+b=**order of the reaction.**

Another important note is that both a and b (the powers to which each concentration is raised) are determined experimentally, and have no relation to the stoichiometric coefficients of the reactants. The rate for a given reaction is an important tool that helps us to calculate the specific order of a reaction and the order of a reaction is important as it enables us to classify specific chemical reactions easily and efficiently, as well as the manner in which they will progress, allowing for appropriate selection of reagents in industrial situations. Knowledge of the reaction order allows us to understand various factors within the reaction including the rate law, units of the rate constant, half life, and much more.

Lastly, the Arrhenius equation must be properly understood when using reaction data to calculate activation energy. The Arrhenius equation is a linear model which plots the natural logarithm of a series of reaction rates against the reciprocal temperature corresponding to each rate as follows:

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

The slope of this linear relation is equal to the negative of the quotient between the activation energy of the reaction (**E_a**) divided by the ideal gas constant (**R**). Thus, after plotting the function, activation energy can be calculated by multiplying the slope value by R. Using calculated values for E_a, it is possible to compare the effectiveness of catalysts by identifying which reduces the activation energy to the lowest amount.

PROCEDURE: (See materials list as outlined in Lab Manual)

PART A: Extracting the catalase from spinach:

1. Spinach is to be chopped and blended until it is in a liquid solution with a certain amount of water. The amount is not important, so long as each group draws catalase from the same solution, which keeps the concentration constant.
2. The blended spinach is then strained to remove all clumps of spinach leaf, and the liquid solution should be collected in a container.
3. To keep the enzymes stable, the catalase container was kept in ice

PART B: determining the partial order of reaction with respect to hydrogen peroxide

4. Assign each of the three teams a volume of hydrogen peroxide (H_2O_2) from the 0.75% w/w% stock solution for use in calculating the calculation of hydrogen peroxide that is to be used. One group is to use 1.0 mL, another will use 2.0 mL, and the third pair will use 4.0 mL. The other partner should ensure that the gas pressure sensor is plugged into the LabQuest apparatus, and that the device is graphing pressure over time.
5. Allow each group to obtain their respective amount of H_2O_2, as well as 10.0 mL of catalase from the stock solution. Both samples should be collected in 10.0 mL graduated cylinders.
6. Fill the remainder of the graduated cylinder containing H_2O_2 with the appropriate amount of distilled water, so that it is a 10.0 mL dilute solution of H_2O_2 and water. Calculate its concentration, and record this to the data table.
7. Fill a small glass test tube with the H_2O_2 solution. Add the catalase solution to a 125 mL Erlenmeyer flask. Ensure that, before adding any solution, both containers are clean and dry, and have been rinsed thoroughly with distilled water.
8. Using a pair of forceps, carefully insert the small test tube of H_2O_2 solution into the Erlenmeyer flask, ensuring that none of the solution spills over into the catalase, which would start the reaction before collection of data.

9. Fit the rubber stopper with tubing to the Erlenmeyer's opening. Ensure that the seal is secure by attaching the gas pressure sensor to the tubing, and then sealing the tubing from the atmosphere by closing the blue valve.
10. Once all of this is complete, tap collect on the LabQuest apparatus, and shake the Erlenmeyer vigorously so that the small test tube containing the H_2O_2 solution mixes thoroughly with the catalase. Doing so will cause the H_2O_2 to decompose, releasing oxygen gas and altering the pressure of the closed system.
11. Once the time has expired or the reaction has plateaued, record the initial pressure, final pressure, and the time of reaction to the data table on Google Sheets. The spreadsheet, when provided with this data, will automatically calculate the rate of reaction for a given trial, using the formula $(p(f) - p(i))/t$.
12. Repeat steps 1-8 for trials 2 and 3. Record all data.

PART C: DETERMINING THE ACTIVATION ENERGY OF THE CATALASE-CATALYZED REACTION

1. From this point forward, all lab groups are to use 2.0 mL of H_2O_2 for all trials (both in part C and part D). While one group member collects the hydrogen peroxide and catalase, the other should prepare an ice bath/hot bath, depending on the group, and attach the thermometer probe to the LabQuest device. Otherwise, there are few differences. Follow steps 1-6 from part B of the lab.
2. Before beginning the mixing process, immerse the closed system in either the hot bath or the cold bath and allow the system to come to an equilibrium. Ensure that the blue valve is open, otherwise (primarily in the hot bath system) pressure can build up to a critical point, which may lead to damage to equipment and potential harm to students. If your group is that for which no temperature-altering bath is required, you may skip this step and proceed to step three.
3. Repeat steps 7-9 from part B. Record all data to the appropriate section of the spreadsheet.

PART D: DETERMINING THE ACTIVATION ENERGY OF THE KI-CATALYZED REACTION

1. Repeat steps 1-6 from part B, once again using 2.0 mL of H_2O_2 in all lab groups. Use the same temperature of system in each group (i.e. same ice bath if you are the ice bath group, etc.). One key difference is that instead of collecting 10.0 mL of catalase, 10.0 ml of potassium iodide (KI) will be used.
2. Repeat the equilibrating process from part C, ensuring once again that the blue valve is open so as not to build up substantial pressure.
3. Commence the mixing process once equilibrium is reached, and once again record the initial pressure, final pressure, and time of reaction to the appropriate sections of the spreadsheet.

OBSERVATIONS:

Data Tables:

Part B:

Table 1: Rate of reaction with 1.0 mL of H_2O_2 *and 10.0 mL of catalase*

Initial pressure	Final pressure	Time of reaction	Rate of Reaction
101.93 kPa	104.84 kPa	26 s	0.1119230769 kPa/s
101.99 kPa	104.64 kPa	25 s	0.106 kPa/s
101.67 kPa	104.71 kPa	28 s	0.1086 kPa/s

Table 2: Rate of reaction with 2.0 mL of H_2O_2 *and 10.0 mL of catalase*

Initial pressure	Final pressure	Time of reaction	Rate of Reaction
102.67 kPa	107.36 kPa	24 s	0.1954167 kPa/s
102.78 kPa	107.23 kPa	30 s	0.1483333 kPa/s
102.59 kPa	106.72 kPa	26 s	0.1588462 kPa/s

Table 3: Rate of reaction with 4.0 mL of H_2O_2 *and 10.0 mL of catalase*

Initial pressure	Final pressure	Time of reaction	Rate of Reaction
101.5 kPa	108.2 kPa	36 s	0.18528 kPa/s
101.53 kPa	107.87 kPa	32 s	0.19813 kPa/s
101.53 kPa	108.2 kPa	27 s	0.24704 kPa/s

Part C:

Table 4: Rate of reaction 2.0 mL of H_2O_2 *and 10.0 mL of catalase when placed in an ice bath*

Initial pressure	Final pressure	Time of reaction	Rate of Reaction
100.83 kPa	103.16 kPa	25 s	0.0932 kPa/s
101.67 kPa	104.19 kPa	27 s	0.09333333333 kPa/s
101.21 kPa	103.8 kPa	27 s	0.09592592593 kPa/s

Table 5: Rate of reaction with 2.0 mL of H_2O_2 *and 10.0 mL of catalase*

Initial pressure	Final pressure	Time of reaction	Rate of Reaction
102.72 kPa	107.30 kPa	28 s	0.1635714 kPa/s
102.91 kPa	107.10 kPa	23 s	0.1821739 kPa/s
102.52 kPa	106.36 kPa	31 s	0.1238710 kPa/s

Table 6: Rate of reaction with 2.0 mL of H_2O_2 *and 10.0 mL of catalase when placed in a hot bath*

Initial pressure	Final pressure	Time of reaction	Rate of Reaction
101.69 kPa	106.59 kPa	43.5 s	0.1126436782 kPa/s
101.69 kPa	107.42 kPa	58.5 s	0.09794871795 kPa/s
101.82 kPa	106.2 kPa	49.5 s	0.08848484848 kPa/s

Part D:

Table 7: Rate of reaction 2.0 mL of H_2O_2 *and 10.0 mL of potassium iodide when placed in an ice bath*

Initial pressure	Final pressure	Time of reaction	Rate of Reaction
101.22 kPa	99.15 kPa	25 s	-0.0828 kPa/s
102.58 kPa	919.73 kPa	26 s	-0.1096153846 kPa/s
101.73 kPa	99.34 kPa	26 s	-0.09192307692 kPa/s

Table 8: Rate of reaction 2.0 mL of H_2O_2 and 10.0 mL of potassium iodide

Initial pressure	**Final pressure**	**Time of reaction**	**Rate of Reaction**
102.76 kPa	104.27 kPa	63 s	0.0239683 kPa/s
102.58 kPa	103.94 kPa	53 s	0.0256604 kPa/s
102.62 kPa	103.69 kPa	65 s	0.0164615 kPa/s

Table 9: Rate of reaction of 2.0 mL of H_2O_2 and 10.0 mL of potassium iodide when placed in a hot bath

Initial pressure	**Final pressure**	**Time of reaction**	**Rate of Reaction**
101.88 kPa	106.07 kPa	41 s	0.102195122 kPa/s
N/A	N/A	N/A	N/A
N/A	N/A	N/A	N/A

DISCUSSION OF RESULTS & SOURCES OF ERROR:

In total, twenty-seven different graphs of pressure vs. time can be produced from the results of the lab, each corresponding to one part of the lab and one lab group, respectively. For the purpose of saving length, one graph from each trial will be presented. Additional graphs may found in the appendix section dedicated to said graphs. Using the data from tables 1, 4, and 7 for the ice bath group, the following graphs are produced:

Figure 1: Decomposition of hydrogen peroxide and its resultant oxygen produced (demonstrated by a change in pressure)

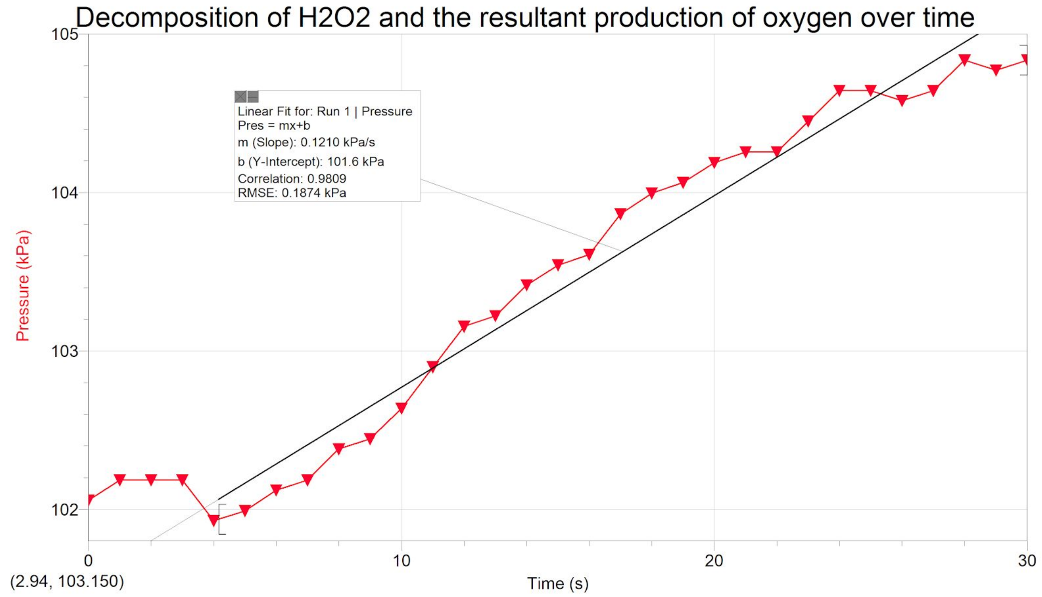


Figure 2: Decomposition of hydrogen peroxide and its resultant oxygen produced when catalyzed by the catalase enzyme, and when placed in an ice bath

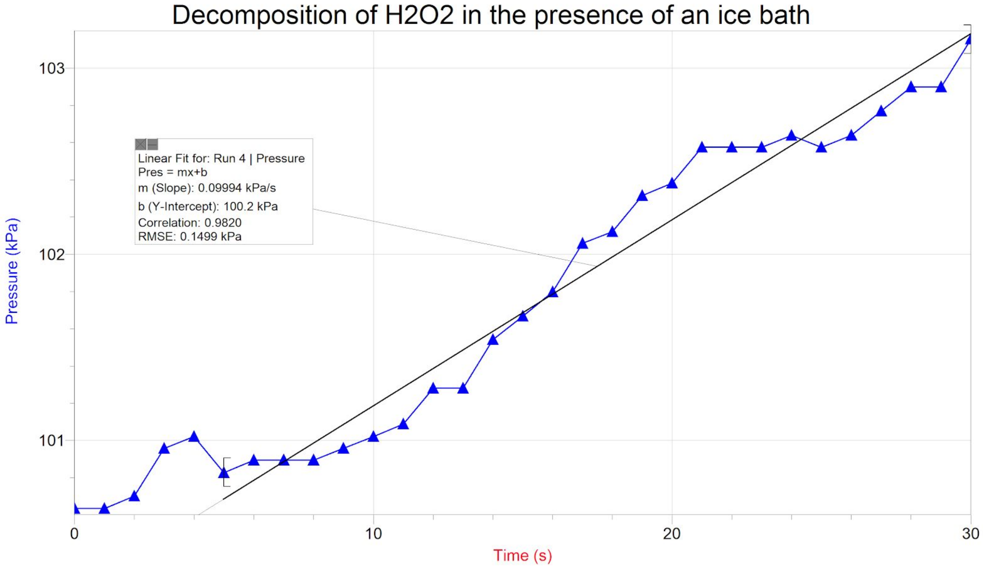
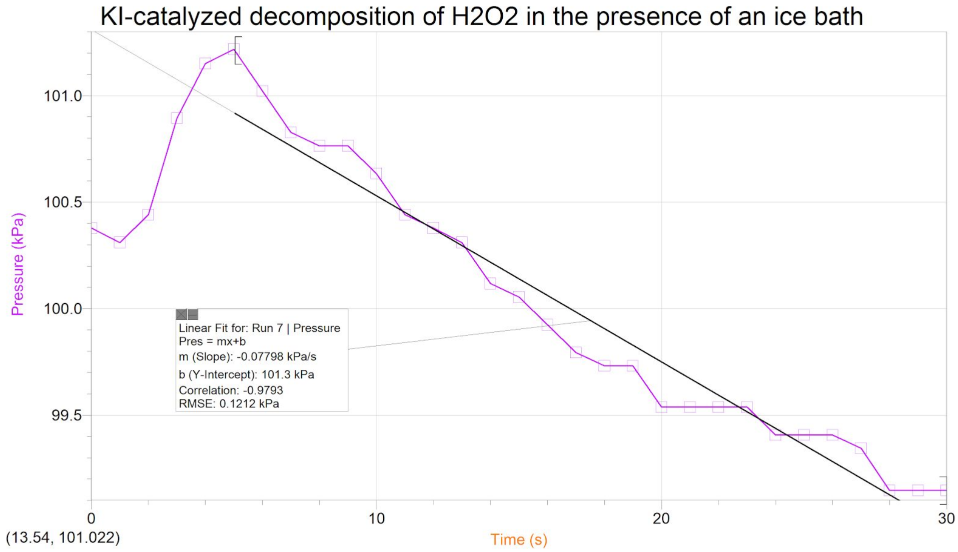


Figure 3: Decomposition of hydrogen peroxide and its resultant oxygen produced when catalyzed by potassium iodide, and when placed in an ice bath



From these first three graphs, it is apparent that the rate of reaction from the room temperature reaction is greater than the reaction occuring in the ice bath. This is an expected result, as the temperature of the surroundings is one of the factors which affects the rate of reaction. However, the third graph, depicting the reaction catalyzed by potassium iodide, displays an unexpected result. Under no circumstances should the pressure have dropped during the reaction. Thus, there must have been some source of error for which the pressure dropped. However, this cannot be due to a gap in the container, as otherwise the pressure would have remained constant at the atmospheric level, which on that day was closer to 101 kPa, rather than the drop below 100 kPa as was observed. The most logical error that would produce this drop is that the Erlenmeyer contained traces of other reactants interfering with the reaction, which caused more H_2O_2 to be produced rather than consumed. Thus, oxygen from the gas mixture would be consumed (along with liquid water from the H_2O_2 solution), and the overall amount of gas in the system would decrease, as well as the overall pressure. Conversely, if there was no source of error, this would indicate that, when combined with potassium iodide, the reverse reaction is catalyzed rather than the forward reaction. However, based on the proposed reaction mechanism:

$$H_2O_{2(aq)} + I^-_{(aq)} \leftrightarrow H_2O_{(l)} + OI^-_{(aq)}$$
$$H_2O_{2(aq)} + OI^-_{(aq)} \leftrightarrow H_2O_{(l)} + O_{2(g)} + I^-_{(aq)}$$

this is unlikely, since the iodine ion clearly acts as a catalyst, since it has no impact on the stoichiometry of the overall reaction. Therefore, the first case stands as the most logical reasoning for the drop in pressure, as potassium iodide simply cannot catalyze the reverse reaction.

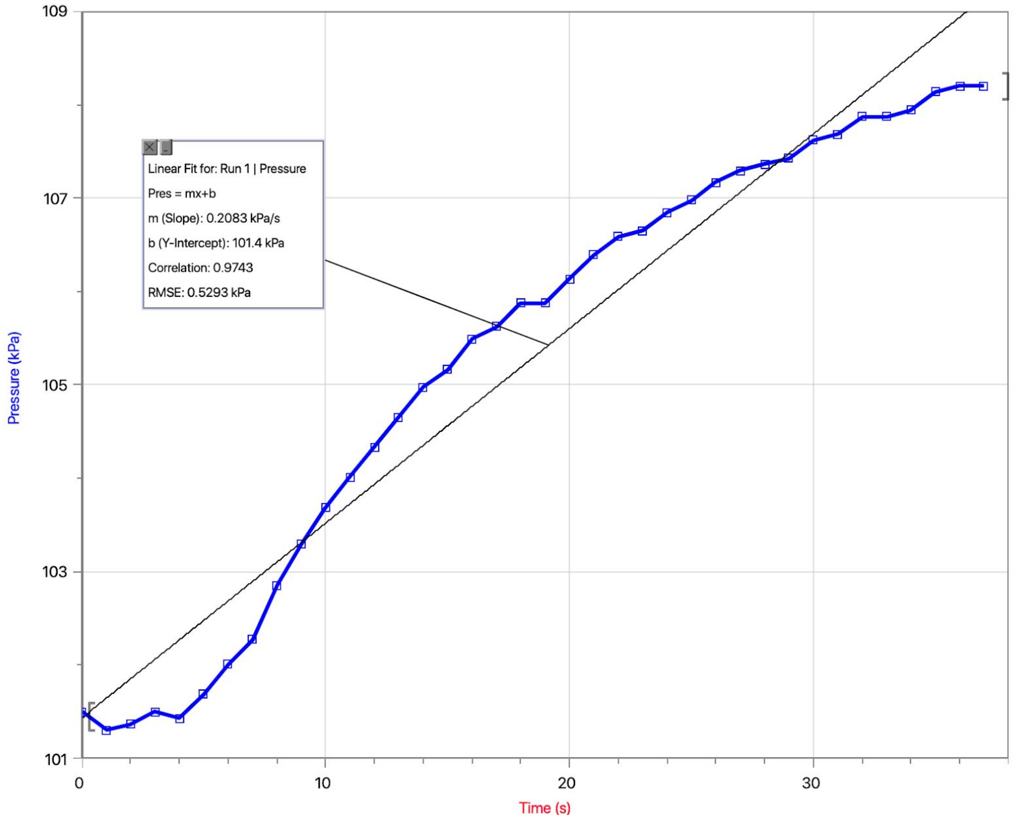


Figure 4: The decomposition of H_2O_2 with the use of 4mL of the enzyme catalase

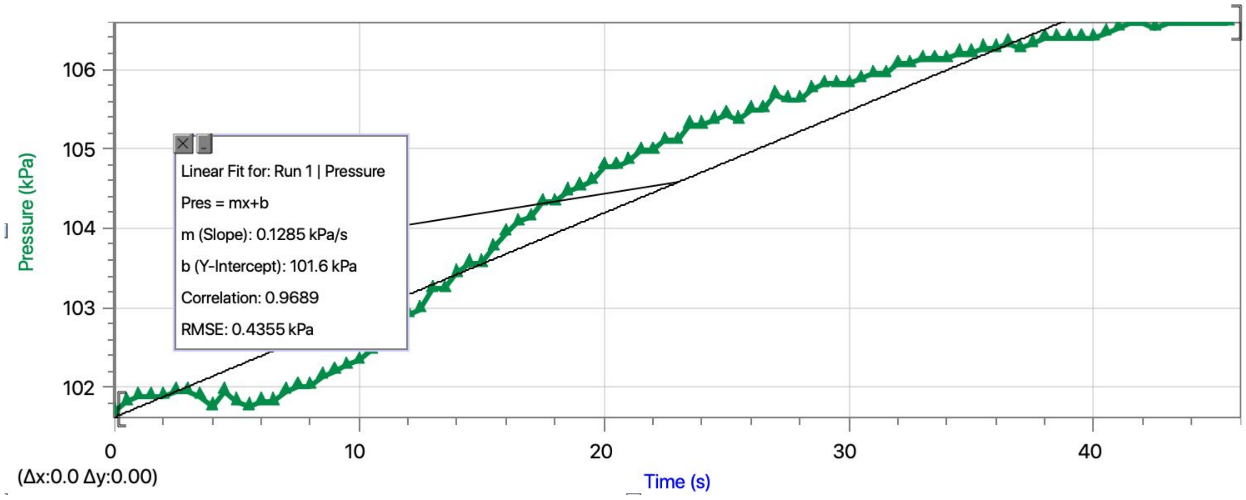


Figure 5: The decomposition of H_2O_2 catalyzed by the enzyme catalase in the presence of a hot bath.

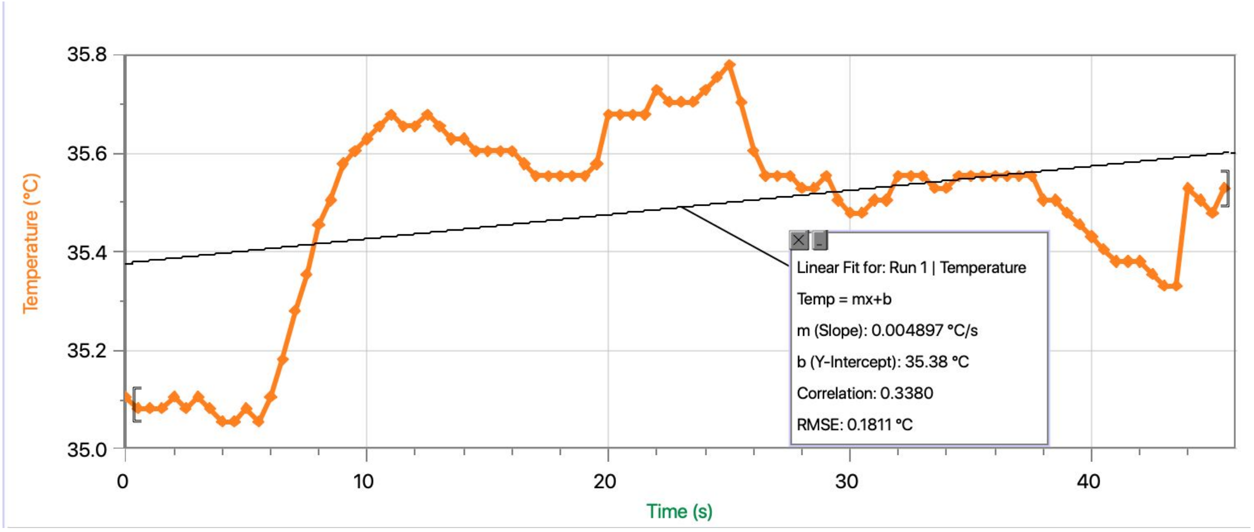


Figure 6: The decomposition of H_2O_2 catalyzed by potassium iodide in the presence of a hot bath.

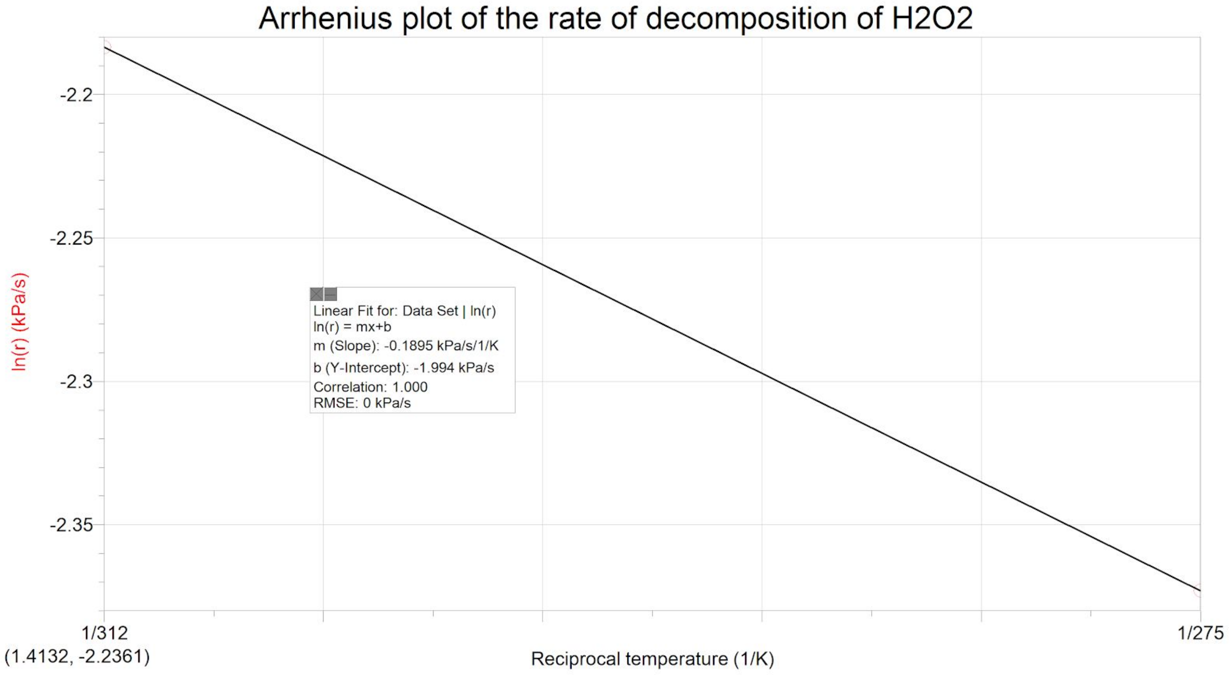
The graphs above show that increasing the volume of hydrogen peroxide, increases its overall concentration which increases the forward and reverse reaction, thus increasing the rate. The graph showing the decomposition of hydrogen peroxide, when catalyzed by catalase in a hot bath (figure 6), shows how increasing the temperature to a level high enough to not degrade the catalase enzyme, will increase its rate. This is evident from the differences in the graphs between the decomposition of hydrogen peroxide using an ice bath.The increase in the rate comes from the higher amount of collisions occurring with the increased heat energy which causes more reactions to take place and thus, raises the rate. The graph displaying the decomposition of hydrogen peroxide with the use of potassium iodide as a catalyst and a hot bath shows how the pressure fluctuates between increasing and decreasing, but overall is seen to increase over time until it begins to equilibrate.

Determination of partial order of H_2O_2

According to the data in table 1 table 2, we observed that when other factors such as temperature and the amount of catalyst are controlled, the higher concentration of H_2O_2, the larger the rate of reaction. In this experiment, the difference of concentration of H_2O_2 between three trials is demonstrated by the volume of H_2O_2 we attained at the beginning of each trial; So as the total volume of each trial is equal, there generates a different concentration of H_2O_2, which is also the only independent variable. By calculation (shown in sample calculation), we noticed that when the concentration of H_2O_2 is doubled, the rate of reaction is also doubled, in other words, in this reaction, the partial order for H_2O_2 is first order. However, a main source of error in the measurement of initial rate is the selection of the time section, in which H_2O_2 is abundant so that the reaction rate can be considered as a constant and equal to the initial rate. Due to Rate=△Pressure/Δt, a little error in time selection can cause a huge difference for rate measurement.

From all previous figures, graphs of the activation energy can be produced using ln(k) as the y-axis values and 1/T (in Kelvin) as the temperature values. The slope of this graph, based on the Arrhenius equation, is the negative of the quotient between the activation energy of the reaction (**E_a**) divided by the ideal gas constant (**R**). Thus, the following figures are produced:

Figure 7: Arrhenius plot of the decomposition of hydrogen peroxide when catalyzed by catalase



Using this plot, and the slope of the graph as the quotient E_a/R, the activation energy for this reaction can be calculated to be approximately 1.5756 J when catalyzed by catalase. However, the accuracy of this graph is questionable, since a loss of data occurred for the room temperature trials, meaning that only two data points were available to construct the above graph. No comparison is available for KI, because (due to the previously noted negative graphs when conducting trials with KI in the ice bath) no negative values exist in the ln function. Thus, only one data point is available, making any attempt at creating a graph and evaluating activation energy invalid. Thus, it can only be assumed that catalase is more effective in catalyzing the decomposition of H_2O_2.

Additional sources of error:

The most significant source of error that occured in the lab investigation was due to the lack of trials done in part D. A misconception occurred assuming that the independent variable for heat was also to be changed; instead of 35 degrees Celsius to 0 degrees celsius. As a result, 3 trials worth of valuable lab time were deemed unusable and only 1 trial with 35 degrees celsius water for the independent variable was used. The lack of data for this trail creates a large margin of error and little to none buffer space for inaccuracies. Having more trials would minimize any such problems.

Having control on the independent variable was also a problem. Based on the location of the instrument, the temperature would greatly differ. Simply breathing too close to the probe would fluctuate the temperature of the water. This can be seen in figure 6.

Pressure was not consistent within the entire flask as gas was leaking out despite putting a rubber stopper. As the experiment would begin, pressure would fluctuate. This can be seen in figure 4. Going from high to low pressure without the experiment beginning can only indicate that. During the trials, if external pressure was not placed on the stopper to keep it in place by the experimenter, the stopper would come off. This had resulted in 2 trials stopping midway because proper data was not collected. If enough pressure was being placed to remove the stopper from the flask, it can only be assumed that pressure was also leaking from the flask.

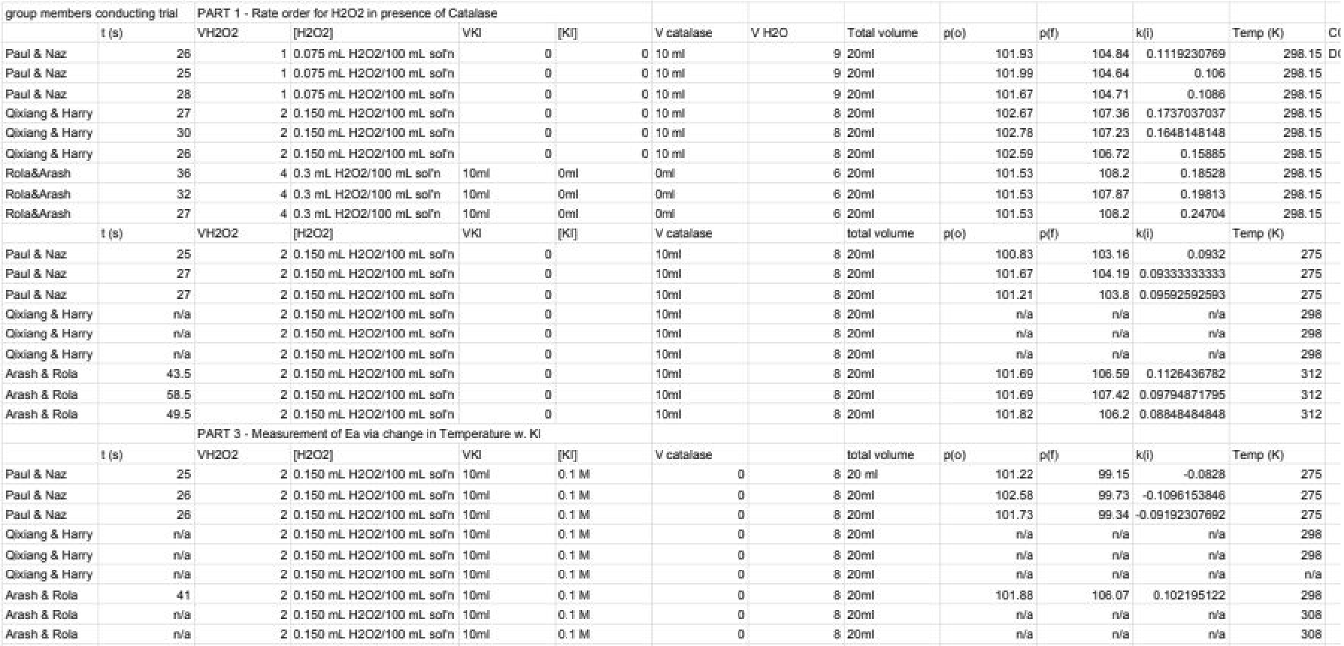
CONCLUSION: Based on the results of the lab, catalase appears to be the more effective catalyst for this reaction. However, due to large losses of data, as well as data inaccuracies, the results of this lab are inconclusive, and more trials are necessary to determine the accuracy of the results obtained during this lab.

REFERENCES:

[1]: Haim, A. (1989). *Réseau de bibliothèques - Library Network*. [online] ACS Publications. Available at: https://pubs-acs-org.proxy.bib.uottawa.ca/doi/pdf/10.1021/ed066p935 [Accessed 4 Dec. 2019].

APPENDICES:

RAW DATA



group members conducting trial		PART 1 - Rate order for H2O2 in presence of Catalase											
	t (s)	VH2O2	[H2O2]	VKI	[KI]	V catalase	V H2O	Total volume	p(o)	p(f)	k(i)	Temp (K)	C(
Paul & Naz	26	1	0.075 mL H2O2/100 mL sol'n	0	0	10 ml	9	20ml	101.93	104.84	0.1119230769	298.15	D(
Paul & Naz	25	1	0.075 mL H2O2/100 mL sol'n	0	0	10 ml	9	20ml	101.99	104.64	0.106	298.15	
Paul & Naz	28	1	0.075 mL H2O2/100 mL sol'n	0	0	10 ml	9	20ml	101.67	104.71	0.1086	298.15	
Qixiang & Harry	27	2	0.150 mL H2O2/100 mL sol'n	0	0	10 ml	8	20ml	102.67	107.36	0.1737037037	298.15	
Qixiang & Harry	30	2	0.150 mL H2O2/100 mL sol'n	0	0	10 ml	8	20ml	102.78	107.23	0.1648148148	298.15	
Qixiang & Harry	26	2	0.150 mL H2O2/100 mL sol'n	0	0	10 ml	8	20ml	102.59	106.72	0.15885	298.15	
Rola&Arash	36	4	0.3 mL H2O2/100 mL sol'n	10ml	0ml	0ml	6	20ml	101.53	108.2	0.18528	298.15	
Rola&Arash	32	4	0.3 mL H2O2/100 mL sol'n	10ml	0ml	0ml	6	20ml	101.53	107.87	0.19813	298.15	
Rola&Arash	27	4	0.3 mL H2O2/100 mL sol'n	10ml	0ml	0ml	6	20ml	101.53	108.2	0.24704	298.15	
	t (s)	VH2O2	[H2O2]	VKI	[KI]	V catalase		total volume	p(o)	p(f)	k(i)	Temp (K)	
Paul & Naz	25	2	0.150 mL H2O2/100 mL sol'n	0		10ml	8	20ml	100.83	103.16	0.0932	275	
Paul & Naz	27	2	0.150 mL H2O2/100 mL sol'n	0		10ml	8	20ml	101.67	104.19	0.09333333333	275	
Paul & Naz	27	2	0.150 mL H2O2/100 mL sol'n	0		10ml	8	20ml	101.21	103.8	0.09592592593	275	
Qixiang & Harry	n/a	2	0.150 mL H2O2/100 mL sol'n	0		10ml	8	20ml	n/a	n/a	n/a	298	
Qixiang & Harry	n/a	2	0.150 mL H2O2/100 mL sol'n	0		10ml	8	20ml	n/a	n/a	n/a	298	
Qixiang & Harry	n/a	2	0.150 mL H2O2/100 mL sol'n	0		10ml	8	20ml	n/a	n/a	n/a	298	
Arash & Rola	43.5	2	0.150 mL H2O2/100 mL sol'n	0		10ml	8	20ml	101.69	106.59	0.1126436782	312	
Arash & Rola	58.5	2	0.150 mL H2O2/100 mL sol'n	0		10ml	8	20ml	101.69	107.42	0.09794871795	312	
Arash & Rola	49.5	2	0.150 mL H2O2/100 mL sol'n	0		10ml	8	20ml	101.82	106.2	0.08848484848	312	
		PART 3 - Measurement of Ea via change in Temperature w. KI											
	t (s)	VH2O2	[H2O2]	VKI	[KI]	V catalase		total volume	p(o)	p(f)	k(i)	Temp (K)	
Paul & Naz	25	2	0.150 mL H2O2/100 mL sol'n	10ml	0.1 M	0	8	20 ml	101.22	99.15	-0.0828	275	
Paul & Naz	26	2	0.150 mL H2O2/100 mL sol'n	10ml	0.1 M	0	8	20ml	102.58	99.73	-0.1096153846	275	
Paul & Naz	26	2	0.150 mL H2O2/100 mL sol'n	10ml	0.1 M	0	8	20ml	101.73	99.34	-0.09192307692	275	
Qixiang & Harry	n/a	2	0.150 mL H2O2/100 mL sol'n	10ml	0.1 M	0	8	20ml	n/a	n/a	n/a	298	
Qixiang & Harry	n/a	2	0.150 mL H2O2/100 mL sol'n	10ml	0.1 M	0	8	20ml	n/a	n/a	n/a	298	
Qixiang & Harry	n/a	2	0.150 mL H2O2/100 mL sol'n	10ml	0.1 M	0	8	20ml	n/a	n/a	n/a	n/a	
Arash & Rola	41	2	0.150 mL H2O2/100 mL sol'n	10ml	0.1 M	0	8	20ml	101.88	106.07	0.102195122	298	
Arash & Rola	n/a	2	0.150 mL H2O2/100 mL sol'n	10ml	0.1 M	0	8	20ml	n/a	n/a	n/a	308	
Arash & Rola	n/a	2	0.150 mL H2O2/100 mL sol'n	10ml	0.1 M	0	8	20ml	n/a	n/a	n/a	308	

SAMPLE CALCULATIONS:

Determining the partial order of H2O2:

Trial 1(with 1.0ml H_2O_2):

$C(H_2O_2)$=75%w/w m(solution)=1g [Assume that the density is equal to that of water]
*$m(H_2O_2)$=1g*75%=0.75*

$n(H_2O_2)$=0.75g/(34g/mol)=0.0221mol

V(catalyst)=10ml $V(H_2O_2)$=10ml

V(total)=V(catalyst)+$V(H_2O_2)$=20ml=0.02L

$C_{final}(H_2O_2)=n(H_2O_2)/V(total)$=1.103mol/L

Average reaction speed (Rate)= (0.1119+0.106+0.1086)/3=0.1088 (KPa/s)

Trial 2(with 2.0ml H_2O_2):

$C(H_2O_2)=75\%w/w$ $m(solution)=2g$ *[Assume that the density is equal to that of water]*
$m(H_2O_2)=2g*75\%=0.75$

$n(H_2O_2)=1.5g/(34g/mol)=0.04412mol$

$V(catalyst)=10ml$ $V(H_2O_2)=10ml$

$V(total)=V(catalyst)+V(H_2O_2)=20ml=0.02L$

$C_{final}(H_2O_2)=n(H_2O_2)/V(total)=2.206mol/L$

Average reaction speed (Rate)= (0.1954+0.1483+0.1588)/3=0.1675 (KPa/s)

Ratio between H_2O_2 concentrations in two trials:

$c2(H_2O_2)/ c1(H_2O_2)=(2.206mol/L)/(1.103mol/L)=2$

Ratio between reaction rates in two trails:

$Rate2/Rate1=(0.1675KPa/s)/(0.1088KPa/s)=1.540$

Calculating the rate of reaction:

Using data from Trial 1 of the 1.0 mL group:

$$r = (p_f - p_i)/t$$

$$r = ((104.84\ kPa) - (101.93\ kPa))/(26\ s)$$

$$r = 0.1119230769\ kPa/s$$

Calculating activation energy from the slope of a graph:

Using data from Figure #

$$-E_a/R = m$$

$$E_a = -Rm$$

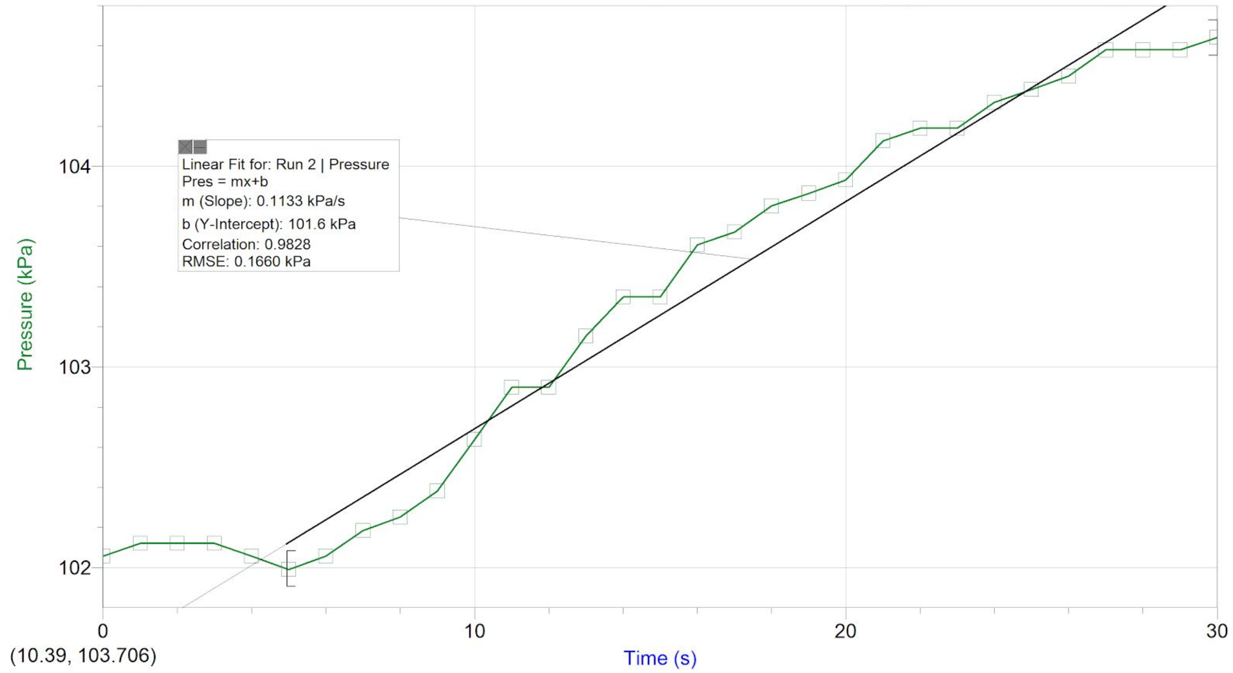
$$E_a \approx 1.5756\ J$$

ADDITIONAL TABLES:

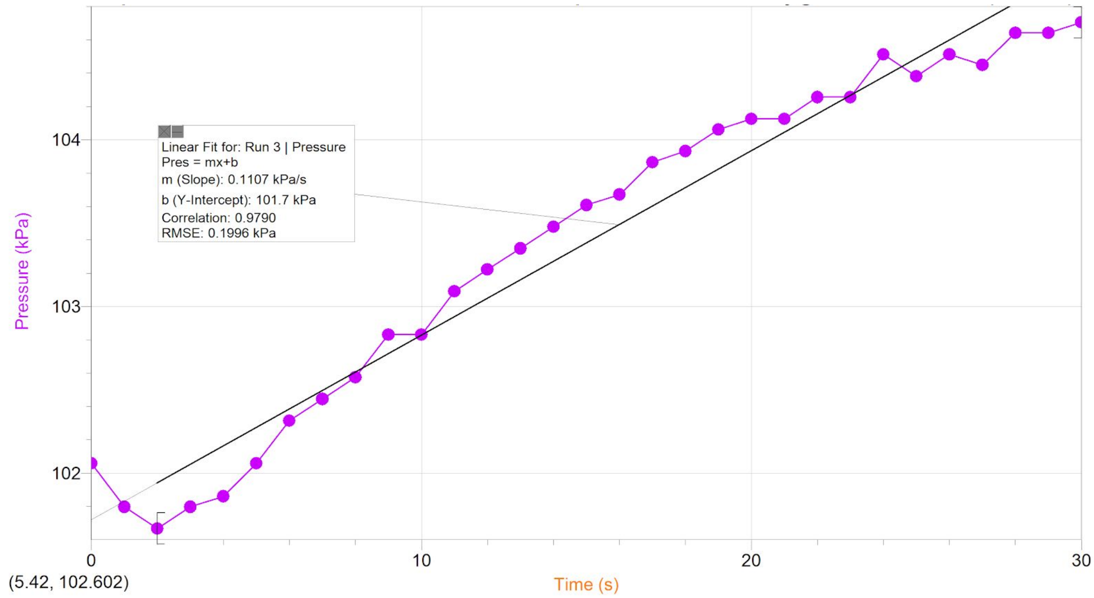
N/A

ADDITIONAL GRAPHS:

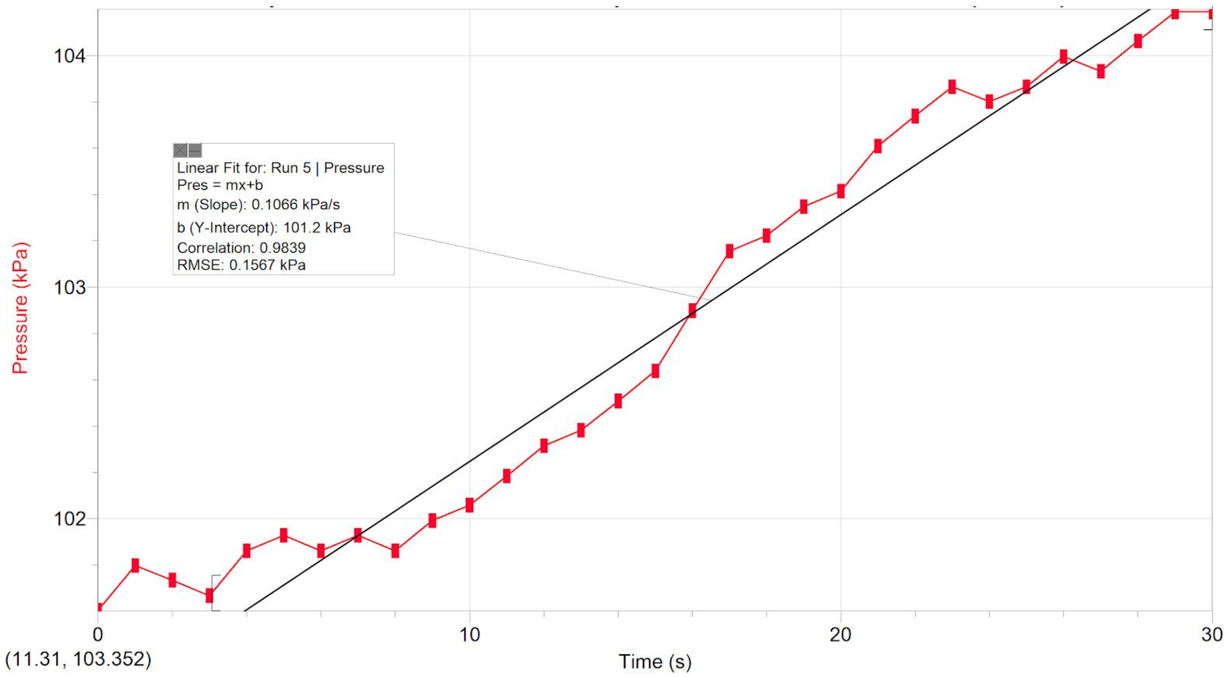
Decomposition of H_2O_2 *and the relevant production of oxygen over time (trial 2)*



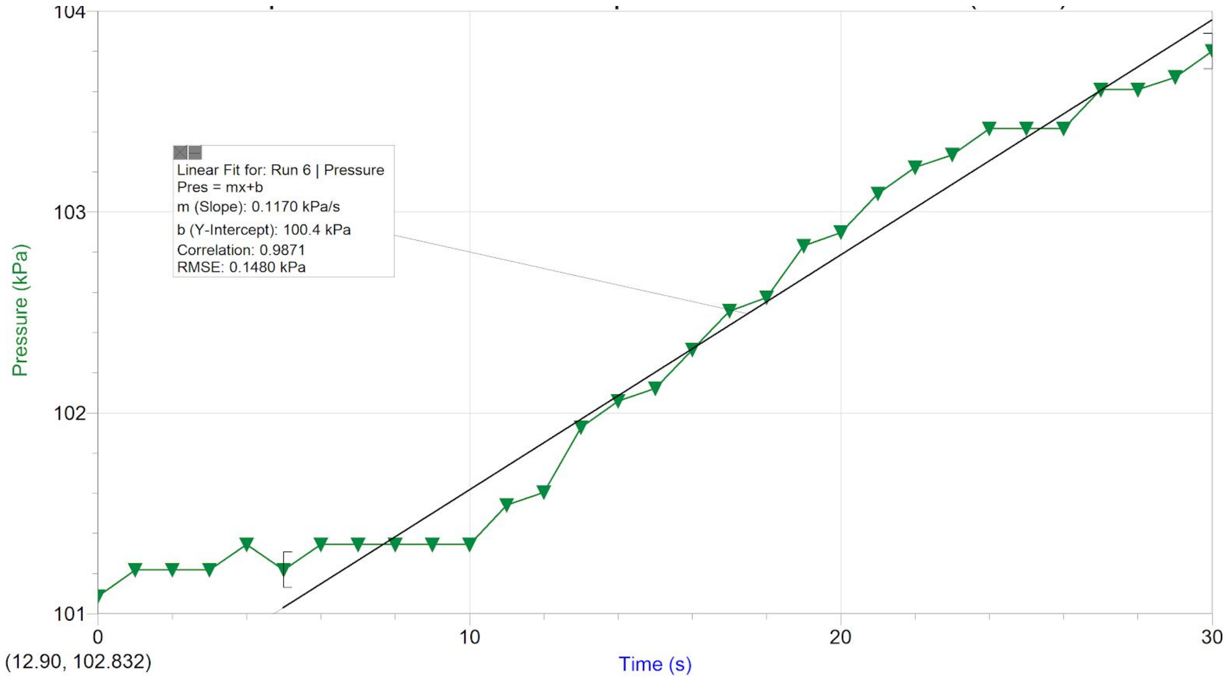
Decomposition of H_2O_2 *and the relevant production of oxygen over time (trial 3)*



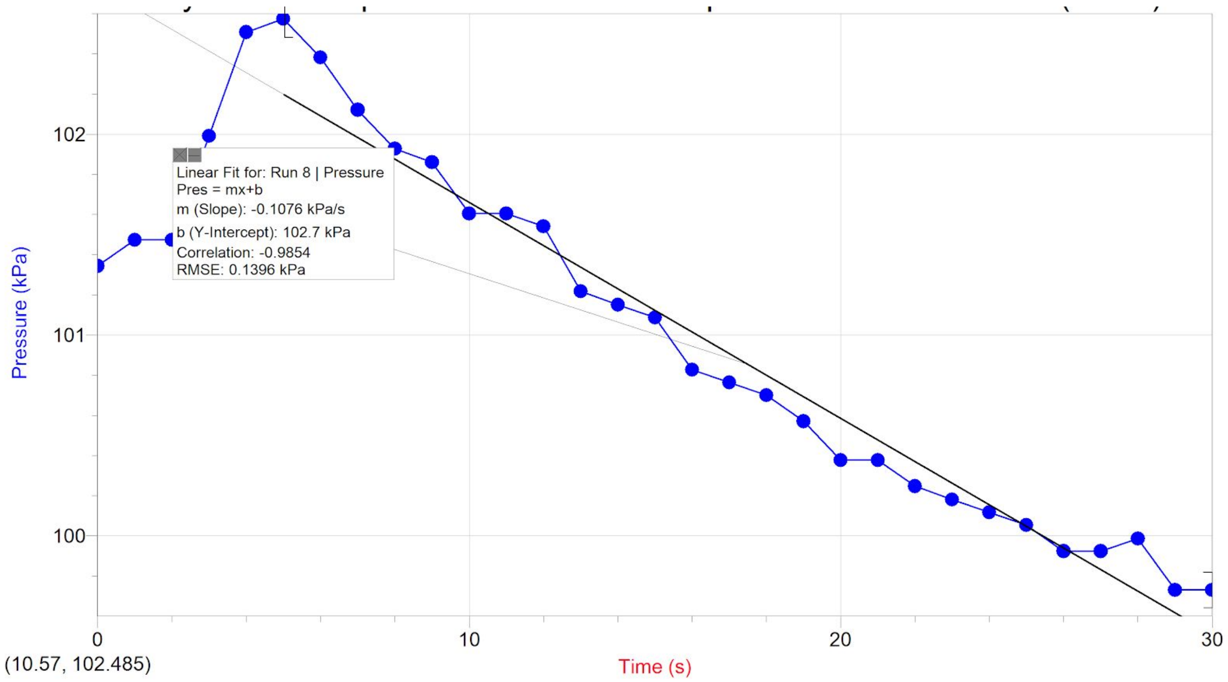
Decomposition of H_2O_2 *in the presence of ice bath (trial 2)*



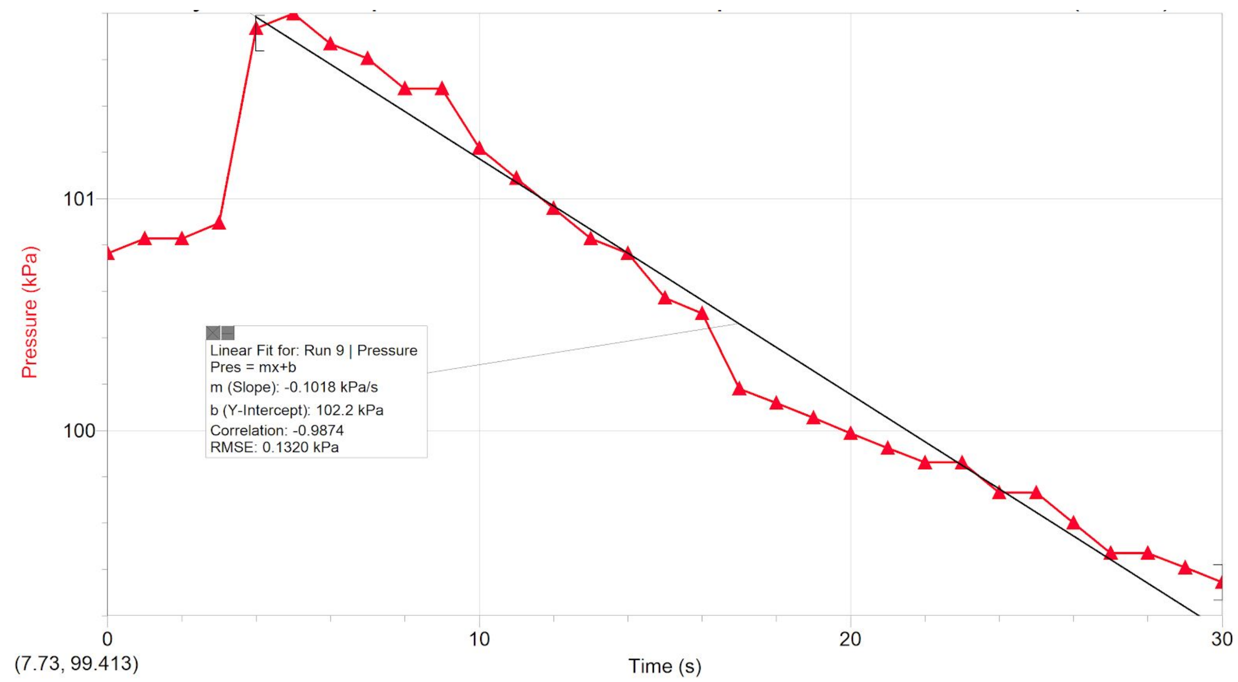
Decomposition of H_2O_2 in the presence of ice bath (trial 3)



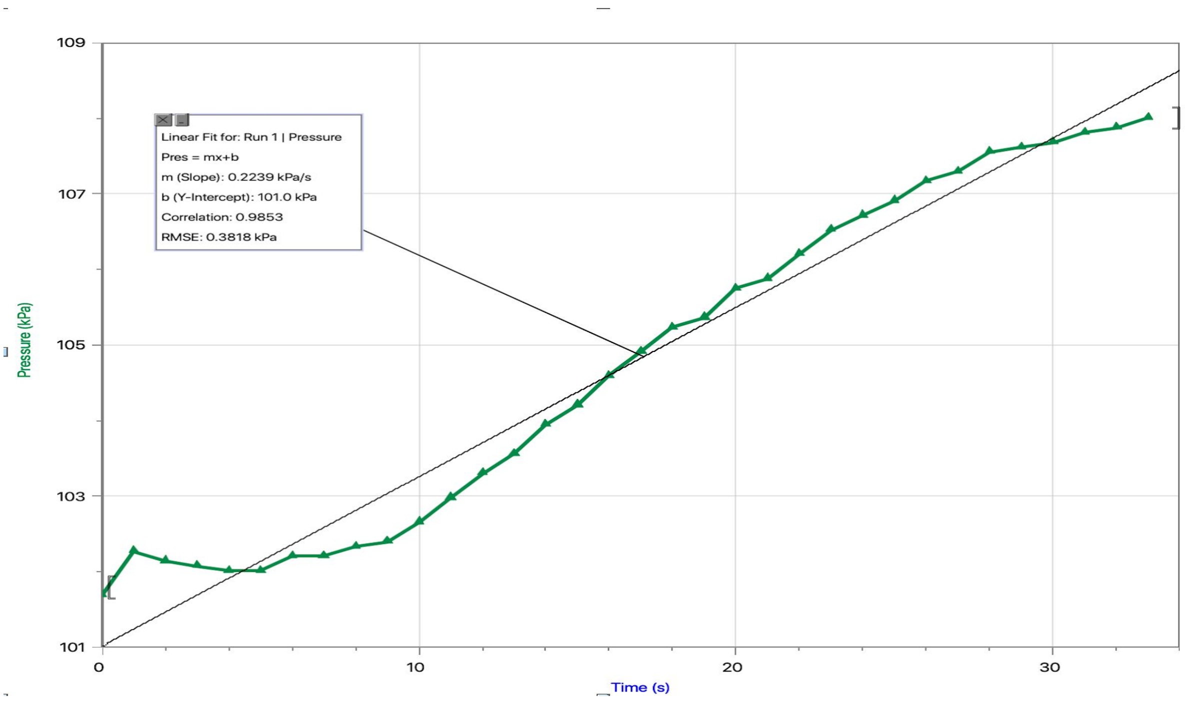
KI- Catalyzed decomposition of H_2O_2 *in the presence of ice bath (trial 2)*



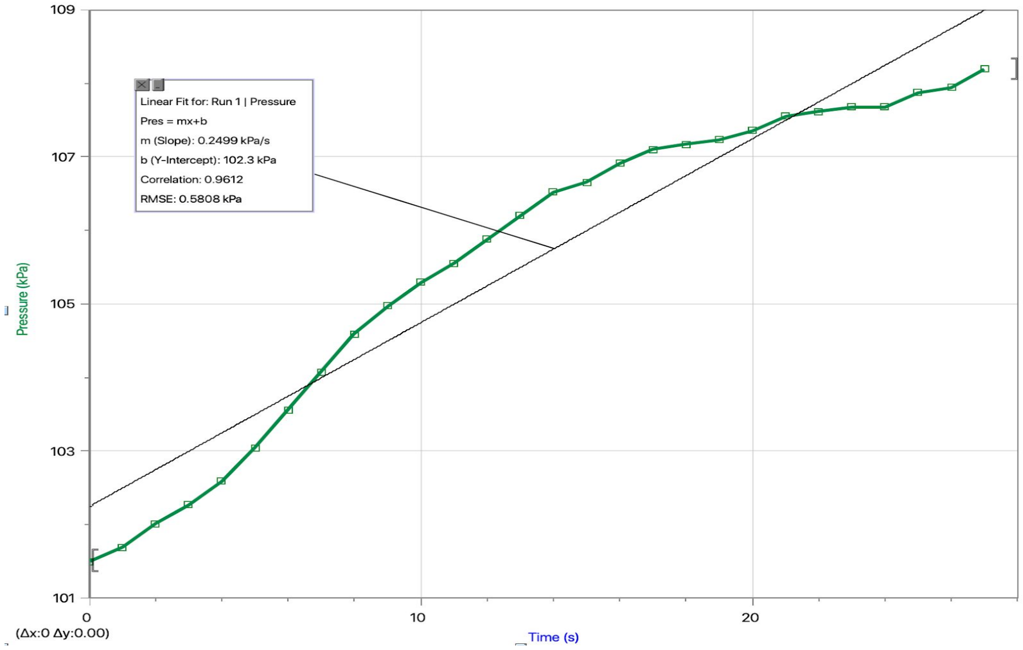
KI- Catalyzed decomposition of H_2O_2 *in the presence of ice bath (trial 3)*



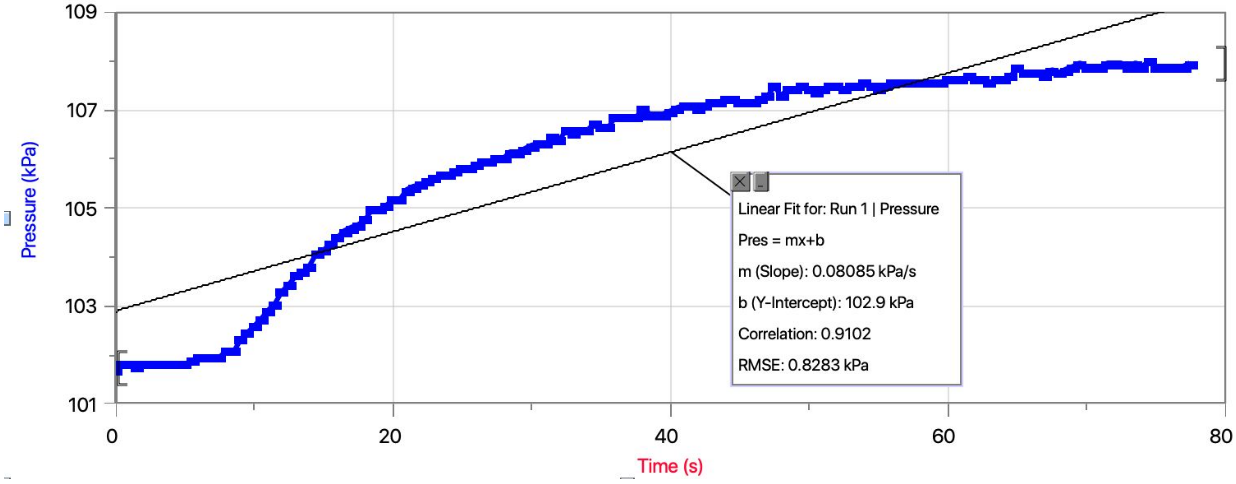
Decomposition of H_2O_2 *Catalyzed by 4mL of Catalase (trial 2)*



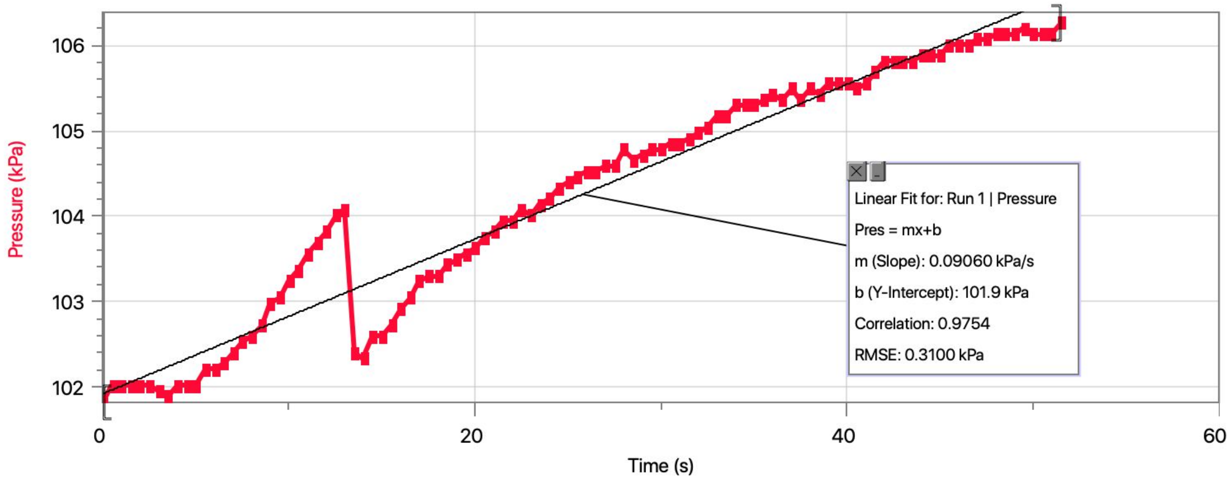
Decomposition of H_2O_2 *Catalyzed by 4mL of Catalase (trial 3)*

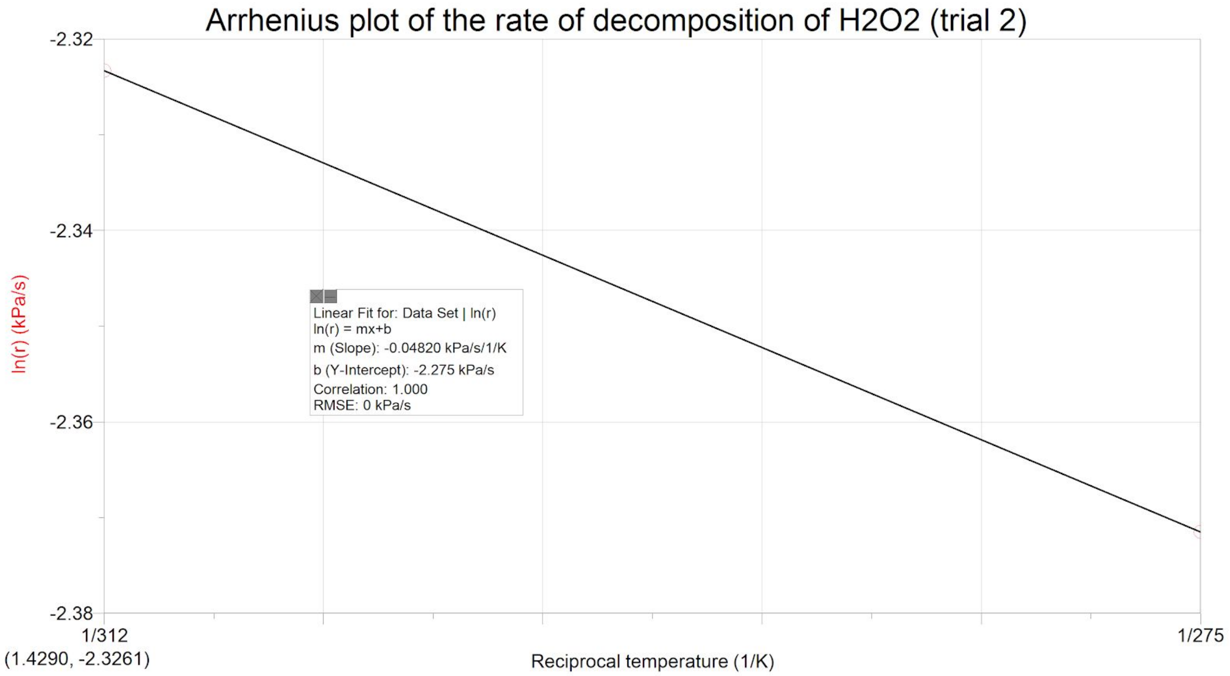


Decomposition of H_2O_2 *in the presence of a hot bath (trial 2)*



Decomposition of H_2O_2 *in the presence of a hot bath (trial 3)*





Arrhenius plot of the rate of decomposition of H2O2 (trial 3)