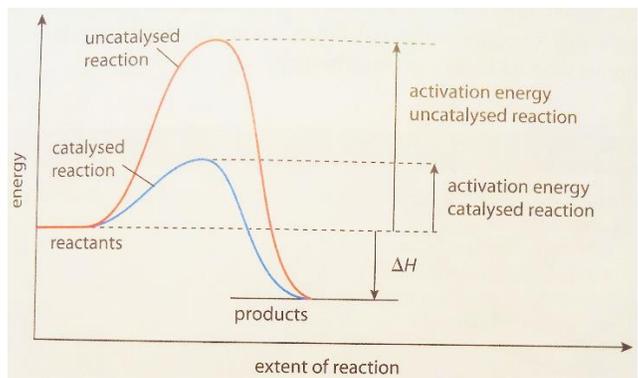


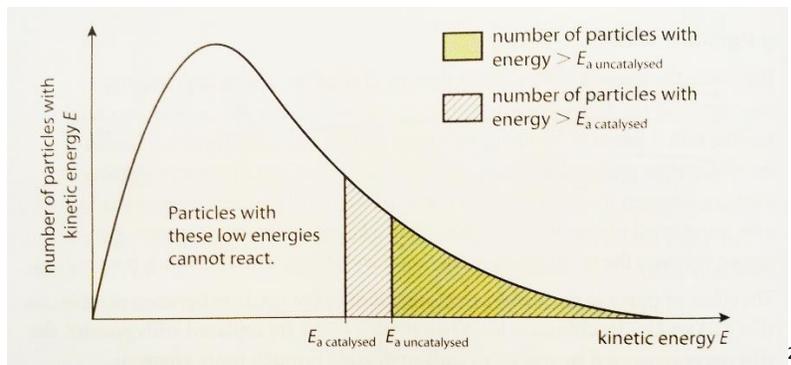
To determine whether H^+ ions in HCl act as a catalyst in the reaction between hydrochloric acid and sodium thiosulfate

Introduction:

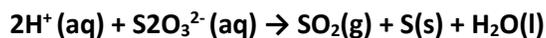
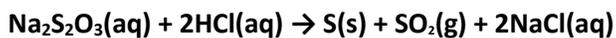
A catalyst is a substance that increases the rate of the reaction by lowering down the reaction's activation energy. Reactions happen as a result of successful collisions and for that, particles need to have a certain amount of kinetic energy. This energy is known as activation energy. Catalysts provide an alternate pathway to the reaction that requires less energy, therefore allowing more successful collisions in a shorter amount of time. This is explained in diagram below.



The diagram below explains this concept using a Maxwell-Boltzmann distribution curve and how catalysts alter activation energy.



The stoichiometric and net ionic equations for the reaction between hydrochloric acid and sodium thiosulfate are:



¹ Pearson Baccalaureate IB Chemistry Higher Level- p;213

² Pearson Baccalaureate IB Chemistry Higher Level- p;214

Majority of catalysts don't undergo chemical change and by the endpoint of the reaction they're the same chemically as they were in the beginning. However, there are some reactions that are more complex in which the reactants or the products themselves temporarily act as catalysts in some steps of the reaction mechanism. This investigation hopes to find out whether hydrogen ions in the acid act as a catalyst by calculating the activation energies of the reaction at different concentrations of the acid. The Arrhenius equation ($k \propto e^{-E_a/RT}$ or $k=Ae^{-E_a/RT}$) can be used to calculate the activation energy.

Research Question:

Based on the activation energies for the reaction between HCl and Na₂S₂O₃ using different concentrations of hydrochloric acid, do the H⁺ ions in HCl act as a catalyst?

Hypothesis:

The H⁺ ions react as a catalyst. The higher the concentration of hydrochloric acid, the lower the activation energy should be; indicating catalytic activity.

Variables:

<p>Independent</p> <ul style="list-style-type: none"> • Temperature of the reactants • Concentration of HCl solution 	<p>Dependent</p> <p>Time taken for solution to become opaque/ catalytic activity</p>
<p>Controlled</p> <ol style="list-style-type: none"> 1. Temperature of the surroundings 2. Concentration of Na₂S₂O₃ solution 3. Volume of HCl and Na₂S₂O₃ solutions 4. Apparatus used 5. Stirring of the reactants once they were mixed 	

A note on the independent variables:

This investigation is a bit unusual in that it requires two variables to be independent, however it's important to note that they weren't independent simultaneously. A non-continuous method was used where all three concentrations of HCl- 0.5 mol dm⁻³, 0.3 mol dm⁻³ and 0.1 mol dm⁻³ were manipulated one by one with a range of temperatures repeated each time. It can be thought of as trials of the same experiment except that the concentration of acid was also changed.

Method of control:

Air conditioning remained constant in the lab throughout. 0.1 mol dm⁻³ concentration of Na₂S₂O₃ solution was used at all times. 10 cm³ volume was controlled using measuring cylinders for both the HCl and Na₂S₂O₃. Same measuring cylinders, beakers, Bunsen burner and thermometer were used each time. Each time the reactants were mixed in the beaker and similar intensity of stirring was tried.

Apparatus:

10 cm³ measuring cylinder (± 0.05 ml)

10 cm³ measuring cylinder (± 0.1 ml)

Thermometer (± 0.1 °C)

Stopwatch (± 1 s)

Two 100 cm³ thick glass beakers

Bunsen burner

Tripod

Gauze

Stirring rod

Droppers

Filter paper (for making a mark)

Materials:

0.1 mol dm⁻³ Na₂S₂O₃ solution

0.5 mol dm⁻³ HCl solution

Distilled water

Match sticks

Markers

Wipes

Procedure:

1. Designate a beaker, a measuring cylinder and a dropper for both Na₂S₂O₃ and HCl solutions with some kind of a marking on them so that they don't get mixed up and cause impurities before the actual reaction. Use the same ones throughout this investigation.
2. Make a bold cross on the filter paper. This will be used to check the opaqueness of the reaction.
3. 0.5 mol dm⁻³ HCl solution will be used first as it needs no further dilutions. Measure 10 cm³ of it and 10 cm³ of Na₂S₂O₃ solution and empty them into two separate beakers. This is the room temperature test, so it needs no heating. Place the beaker with sodium thiosulfate over the mark on the filter paper.

4. Pour the acid into the sodium thiosulfate with one hand and start the stopwatch with the other. Immediately put the thermometer in and stir the solutions.
5. The solution gradually becomes opaque. Stop the stopwatch the moment the cross is no longer visible and note the temperature in the thermometer followed by noting the time on the stopwatch.
6. Wash the tip of the thermometer and the beaker used for mixing HCl and Na₂S₂O₃ solutions with distilled water.
7. The next 4 temperatures will be based on a 'guiding temperature' that is roughly in the 30s, 40s, 50s and 60s degrees Celsius. This is tricky as usually the solutions will heat more than the value on the thermometer and it might be hard to stick to the guiding temperature. Heat the solutions until the thermometer reads about 10 degrees less than the desired amount.
8. Repeat steps 3, 4, 5 and 6 but this time after heating the solutions as described in step 6. In total 5 different times should be achieved with 5 different temperatures.
9. Now this experiment has to be repeated with a lower HCl concentration of 0.3 mol dm⁻³. This can be achieved by having 6 mL of acid in the measuring cylinder and filling in the rest of 4 mL with distilled water. Repeat steps 3-8 with the only difference being the dilution of HCl as described in this step.
10. The whole process so far has to be repeated yet again but with 0.1 mol dm⁻³ concentration of HCl. This can be achieved by having only 2 mL of acid in the measuring cylinder and filling in the rest with 8 mL distilled water each time. The rest is a repetition of step 9.
11. The practical part of this investigation is over. Note down any qualitative observations during the experiments along with the temperatures and time taken previously noted and clear the working area

Risk Assessment:

Hydrochloric acid solution is corrosive to eyes and skin. It is moderately toxic by ingestion and inhalation. Sodium thiosulfate solution is a body tissue irritant. Both should be handled carefully. The reaction generates sulfur dioxide which is also a skin and eye irritant. Therefore this investigation should be performed in a well-ventilated lab. The beakers with solutions in them will be hot on heating so they should be held with caution. Heat resistant gloves, safety goggles and a lab coat are recommended to avoid direct contact with any of the chemicals with eyes or skin.

Photographs of the Setup and Apparatus:



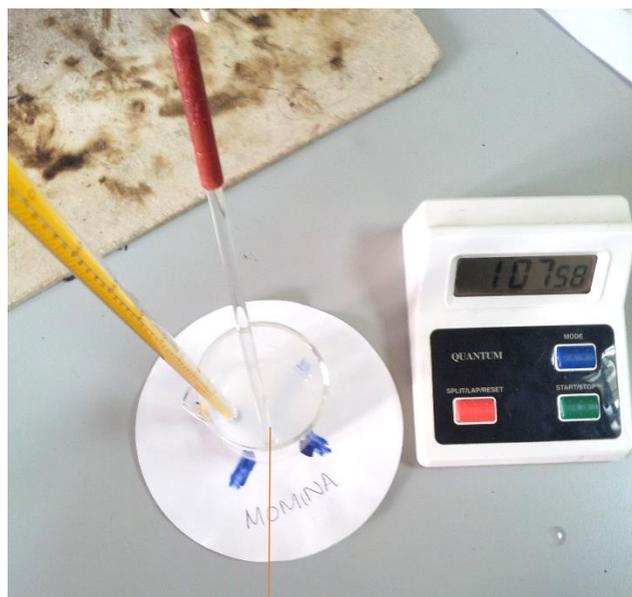
± 0.1 °C thermometer and stirring rod

100 mL beakers containing HCl and $\text{Na}_2\text{S}_2\text{O}_3$ solutions placed on the tripod at the same time

Bunsen burner with a blue flame



Just after the start of the reaction, the solutions are still almost clear and the blue cross is visible. Stirring rod is used to even out the concentration of reactants in the beaker, thermometer measures the rising temperature.



Almost at the end of the reaction with the timer reading 1 minute 7.58 seconds. The products have now formed and the solution is white and almost opaque, with the cross underneath barely visible.

Data collection:

Conc. of HCl/ mol dm ⁻³	Temp./ °C ±0.1	Time/ seconds ±1
0.5	24.0	59
	35.0	29
	41.0	24
	53.0	16
	62.0	7
0.3	24.8	74
	34.0	38
	42.0	26
	50.0	20
	63.0	9
0.1	22.6	94
	35.0	55
	47.0	30
	57.0	17
	64.0	13

Table 1: Raw data table

Qualitative analysis:

- As the reaction between hydrochloric acid and sodium thiosulfate progressed, the clear and colorless solution gradually became white, then off-white and opaque.
- This process was much faster in higher temperatures in each of these tests.
- Using 0.5 mol dm⁻³ HCl, the resulting solution of products was more yellowish/ off-white compared to the 0.1 mol dm⁻³ concentration.

Data Processing:

The data collected above needs to be processed in order for it to be used in the Arrhenius equation and therefore find the activation energy. Firstly, the temperature needs to be converted in Kelvin, so 273 is added to the degrees Celsius values. 24°C, for example, becomes 297 K.

The Arrhenius equation is or $k=Ae^{-E_a/RT}$ where A is the Arrhenius constant which gives the frequency of successful collisions based on collision geometry and its units are the same as k , the rate constant of a first order reaction. E_a is activation energy, R is the gas constant at

8.31 J K⁻¹ mol⁻¹ and T is temperature in K. In order to produce a linear equation, natural log of both sides is taken resulting in $\ln k = -E_a/RT + \ln A$. A graph of $\ln k$ against $1/T$ will give a straight line with the gradient $= -E_a/R$. After finding the gradient from the graph, activation energy is easy to derive as it's simply multiplying the gradient with the gas constant (changing the subject of the formula).

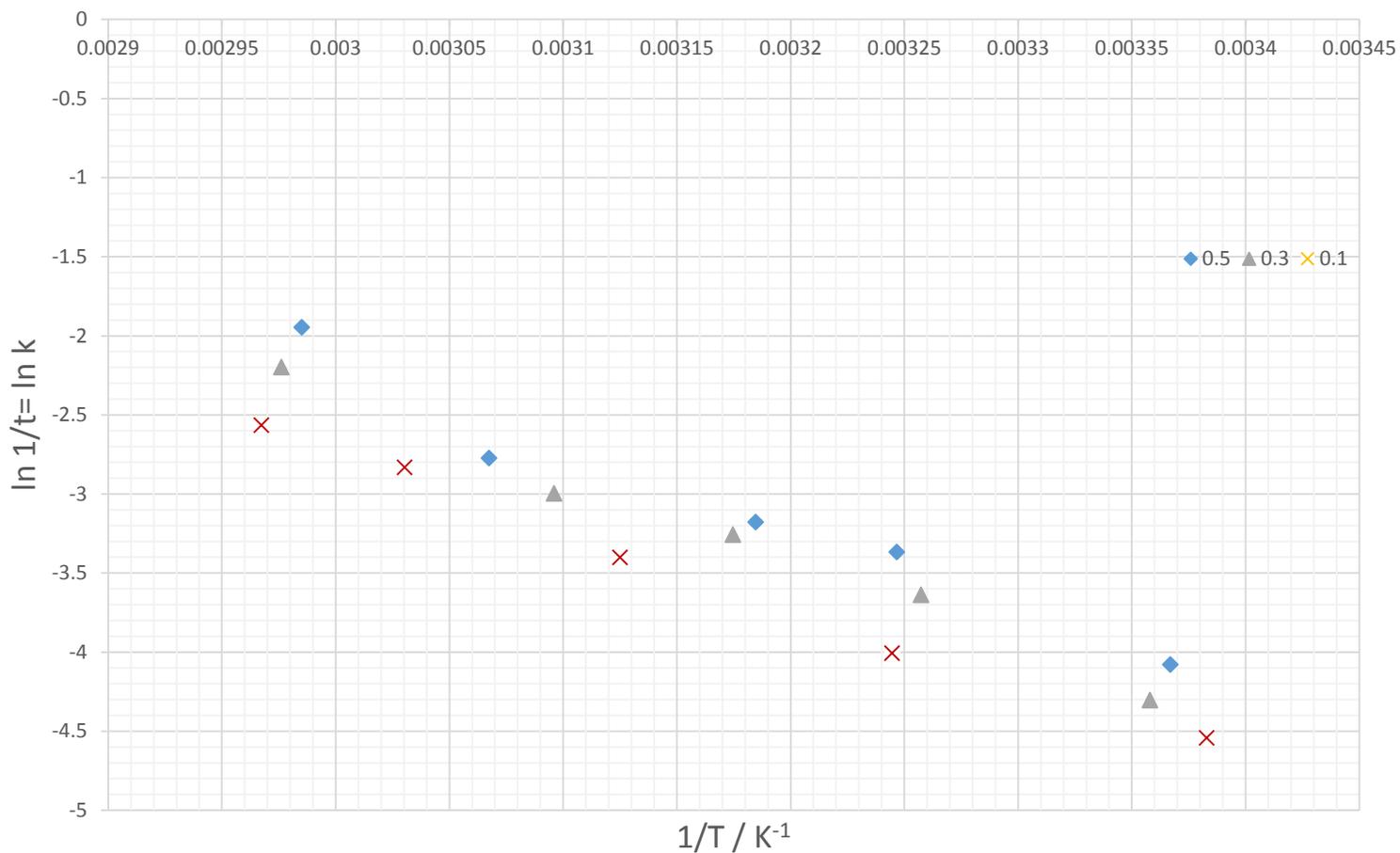
Conc. of HCl/ mol dm ⁻³	Temperature/ °C	Temperature (T)/ K	Time (t)/ s	1/t / s ⁻¹	1/T / K ⁻¹	ln 1/t
0.5	24.0	297	59	0.016949	0.003367	-4.07754
	35.0	308	29	0.034483	0.003247	-3.3673
	41.0	314	24	0.041667	0.003185	-3.17805
	53.0	326	16	0.0625	0.003067	-2.77259
	62.0	335	7	0.142857	0.002985	-1.94591
0.3	24.8	297.8	74	0.013514	0.003358	-4.30407
	34.0	307	38	0.026316	0.003257	-3.63759
	42.0	315	26	0.038462	0.003175	-3.2581
	50.0	323	20	0.05	0.003096	-2.99573
	63.0	336	9	0.111111	0.002976	-2.19722
0.1	22.6	295.6	94	0.010638	0.003383	-4.54329
	35.2	308.2	55	0.018182	0.003245	-4.00733
	47.0	320	30	0.033333	0.003125	-3.4012
	57.0	330	17	0.058824	0.00303	-2.83321
	64.0	337	13	0.076923	0.002967	-2.56495

Table 2: Processed data table

NB: 3 significant figures are not used in this table because that would lead to $1/T$ column having the same value of K⁻¹ in many of the trials, affecting the later calculations.

Now this data can be plotted on graph to produce a linear function. As explained before processing, this Arrhenius plot will help in deriving the activation energy. The graph is overleaf along with the gradients taken between 2nd and 4th intervals of each of the 3 concentrations of HCl using $\frac{\Delta y}{\Delta x}$.

Arrhenius plot for deriving activation energy



Conc. Of HCl/ mold m ⁻³	Gradient calculated	Activation Energy (Gradient × R)/ J mol ⁻¹	Activation energy / kJ mol ⁻¹
0.5	-3317.41	-27567.7	-27.7
0.3	-3977.93	-33056.6	-33.1
0.1	-5477.75	-45520.1	-45.5

Table 3: Activation energy calculation (final value in 3 significant figures)

Uncertainties:

Total percentage uncertainty:

$$\text{Percentage uncertainty} = \left(\frac{\text{net uncertainty}}{\text{measured value}} \right) \times 100\%$$

The uncertainties mentioned before can be used to form an expression of the percentage uncertainty. Since there were no before and after values recorded, the net uncertainty is the same as uncertainty of the apparatus. 100 mL of HCl and Na₂S₂O₃ were used each time but the temperature and time taken varied, so they'll be described using variables x and y respectively and calculated below in Table 4 (all values are up to 3 decimal places)

(Measuring cylinders) (Thermometer) (Timer)

$$\begin{array}{c} \downarrow \qquad \qquad \downarrow \qquad \qquad \swarrow \\ \frac{0.05}{10} \times 100 + \frac{0.1}{10} \times 100 + \frac{0.1}{x} \times 100 + \frac{1}{y} \times 100 \\ \\ = \pm \left(0.5 + 1 + \frac{10}{x} + \frac{100}{y} \right) \% \end{array}$$

Table 4: Total percentage uncertainty calculation

Conc. of HCl/ mol dm ⁻³	Temp./ °C ±0.1	Percentage uncertainty of the temperature	Time/ seconds ±1	Percentage uncertainty of the time taken	Total percentage uncertainty
0.5	24	±0.417 %	59	1.695	±3.612 %
	35	±0.286 %	29	3.448	±5.234 %
	41	±0.244 %	24	4.167	±5.911 %
	53	±0.189 %	16	6.250	±7.939 %
	62	±0.161 %	7	14.286	±15.947 %
0.3	24.8	±0.403 %	74	1.351	±3.255 %
	34	±0.294 %	38	2.632	±4.426 %
	42	±0.238 %	26	3.846	±5.584 %
	50	±0.200 %	20	5.000	±6.700 %
	63	±0.159 %	9	11.111	±12.770 %
0.1	22.6	±0.442 %	94	1.064	±3.006 %
	35.2	±0.284 %	55	1.818	±3.602 %
	47	±0.213 %	30	3.333	±5.046 %
	57	±0.175 %	17	5.882	±7.558 %
	64	±0.156 %	13	7.692	±9.349 %

Absolute uncertainty:

Total percentage uncertainties can be converted into the absolute uncertainty by multiplying them with the activation energy derived in Table 3.

For example, in the first interval for 0.5 mol dm⁻³ HCl, $\frac{3.611582}{100} \times -27.5677$ will give a value of -0.99563 kJ mol⁻¹. This is the maximum possible amount more or less the activation energy previously calculated in table 3. Table 5 below shows absolute uncertainty calculations for all of the other intervals and concentrations of HCl up to 3 decimal places.

Conc. of HCl/ mol dm ⁻³	Total percentage uncertainty	Activation Energy/ kJ mol ⁻¹	Absolute uncertainty/ kJ mol ⁻¹
0.5	±3.612 %	-27.568	±0.996
	±5.234 %		±1.443
	±5.911 %		±1.629
	±7.939 %		±2.189
	±15.947 %		±4.396
0.3	±3.255 %	-33.057	±1.076
	±4.426 %		±1.463
	±5.584 %		±1.846
	±6.700 %		±2.215
	±12.770 %		±4.221
0.1	±3.006 %	-45.520	±1.368
	±3.602 %		±1.640
	±5.046 %		±2.297
	±7.558 %		±3.440
	±9.349 %		±4.255

Table 5: Absolute uncertainty calculation

Conclusion:

The results seem to indicate that H^+ ions indeed act as a catalyst in the reaction between sodium thiosulfate and hydrochloric acid. This supports the hypothesis as with 0.5 mol dm^{-3} concentration of HCl solution, the activation energy was more than 5 kJ mol^{-1} lower compared to that of 0.3 mol dm^{-3} concentration and 18 kJ mol^{-1} lower compared to the 0.1 mol dm^{-3} concentration. The higher the concentration of HCl, the lower the activation energy indicating catalytic activity. However, the gradients did not look very different. The 2nd, 3rd and 4th points only were used for activation energy calculation each time. The points on the graph were not completely on a straight line either.

Increasing the concentration of the H^+ ions speeds up the rate of reaction. This would be true even if they weren't acting as a catalyst, because with a higher concentration there is a greater probability of molecules achieving the right geometric orientation and sufficient energy for successful collisions. But increasing the concentration doesn't alter the activation energy at all. Only catalysts can provide an alternate pathway that requires less activation energy as seen in the diagrams in the introduction. With a higher concentration of H^+ ions, more of the catalyst is present and it catalyzes the reaction between more molecules. This is why 0.5 mol dm^{-3} has the least amount of activation energy; and 0.1 mol dm^{-3} has the most.

H^+ ions in this reaction act as a type of homogenous catalyst; that is they're in the same aqueous solution state as the rest of reactants. The reaction between hydrochloric acid and sodium thiosulfate, according to the results, can be said to be autocatalytic- it is catalyzing itself. One of the reactants is also the catalyst, and as the concentration of the reactant is decreases as the reaction progresses, the concentration of the catalyst also decreases. This is unusual because most ordinary catalysts aren't altered chemically by the end point of the reaction.

Evaluation:

The most prominent problem with this investigation is that there are no trials. Because of the repetitive nature of the experiment itself and that it was a non-continuous method of investigation, it was very time consuming. A lack of trials makes the results unreliable and puts the whole conclusion into question because this is what can be said with one trial only; it's unpredictable whether other trials would have given the same results, especially given that there isn't a massive difference in the activation energies and some intervals have a very large total percentage uncertainty.

The experimental design could use many improvements. Heating the solutions as was done in this investigation was tricky and twice, the solution had to be discarded because the temperature was too high above the guiding temperature. This unnecessarily wasted time in an investigation that had many time constraints.

On the apparatus and data collection, the measuring cylinder for HCL and $\text{Na}_2\text{S}_2\text{O}_3$ were not of the same type and had different uncertainties whereas all apparatus should have been similar. The short reactions also had a much bigger percentage uncertainty as seen in Table 4, which were increased a great deal because only seconds were recorded even though microseconds were also visible on the stopwatch.

With some of the higher temperature intervals, it was hard to pour the acid in, to start the stopwatch and immediately put in the thermometer and stiller all at once. The reactions would last only a few

seconds sometimes and the reaction time to do all that could have cause many random errors. This investigation should ideally be performed in pairs as it makes managing everything easier.

The surrounding temperatures are supposed to be a controlled variable but the practical of this experiment was done over lessons on two days and the outside heat and sunlight conditions would have been different, causing the AC to cool in a different way. This might have had an impact on some of the results.

Improvements:

Improvement	How it enhances the investigation
1. Use of a water bath	Instead of using a Bunsen burner, a water bath would have been far more effective. The heat would be distributed more uniformly, heating the beakers in a way where the temperature can be controlled more easily to achieve it close to the guiding temperatures.
2. Carry out the investigation with at least 5 trials	Makes the results more conclusive as more extensive data is available for manipulation and limits the impact of random errors on the results.
3. Use a burette	Burettes have a lesser uncertainty compared to measuring cylinders. This would reduce the total percentage uncertainty and make the investigation more precise.
4. Use a digital thermometer	Not only will it reduce the total percentage uncertainty, it will also make it easier to record the temperature at the right time since a lot of the other equipment has to be handled too and this won't need to be handled in the same way. Also reduces the parallax errors.
5. Record microseconds on the stopwatch	Reduces the percentage uncertainty a great deal. This would have a big impact especially on the higher temperature intervals.

Works cited:

Brown, C. and Ford, M. 2009. *Chemistry*. Harlow, Essex: Pearson Education.