Chemistry Lab Report

Experiment Title

To determine the effect of Intermolecular Forces on the Enthalpy of Solution for various metal hydroxides

Name: Abhay Dalmia
Session: May 2014
School: Gandhi Memorial International School

Abstract:

In this experiment, energetics theories and principles were used to calculate the enthalpy of solution of several hydroxides. Calorimetry carried out in a polystyrene cup was used to measure the enthalpy of solution. A temperature probe was used for accurate data tabulation.

The purpose of this experiment is to determine the affect of intermolecular forces, and hence reactivity on the enthalpy of solution. The enthalpy solution values were tabulated, and then conclusions were drawn as to the difference in reactivity and intermolecular forces.

Through this experiment, it was concluded that the stronger the cation in the compound, makes it harder to disassociate, and hence is more endothermic that the enthalpy of solution of a cation that is weaker in its compound. Hence, the enthalpy of solution of sodium hydroxide was found to be more endothermic than that of potassium hydroxide, as sodium has a smaller ionic radius than potassium, and hence attracts its anion strongly.

Similarly, it was found that cations with a relative positive charge of 2, have a more endothermic enthalpy of solution than those with a positive charge of 1, since they attract their anion counterpath with greater intermolecular forces, hence takes more energy to break bonds. Since bond breaking is endothermic, the enthalpy is also more endothermic.

Overall the experiment was quite successful since it fulfilled the aim and its conclusions were useful. Even uncertainty was kept to a minimum. However, there was great uncertainty in the case of calcium hydroxide due to its limited solubility in water.

CONTENTS

INTRODUCTION

The enthalpy change of solution [ΔH^{θ}_{sol}] is the enthalpy change when one mole of solute dissolves in an infinite volume of water. In practice, a point of infinite dilution is not possible, so instead, it is common practice to reach a point where further dilution has no significant effect on the enthalpy of solution. Then results are extrapolated to obtain accurate and precise results.[1]

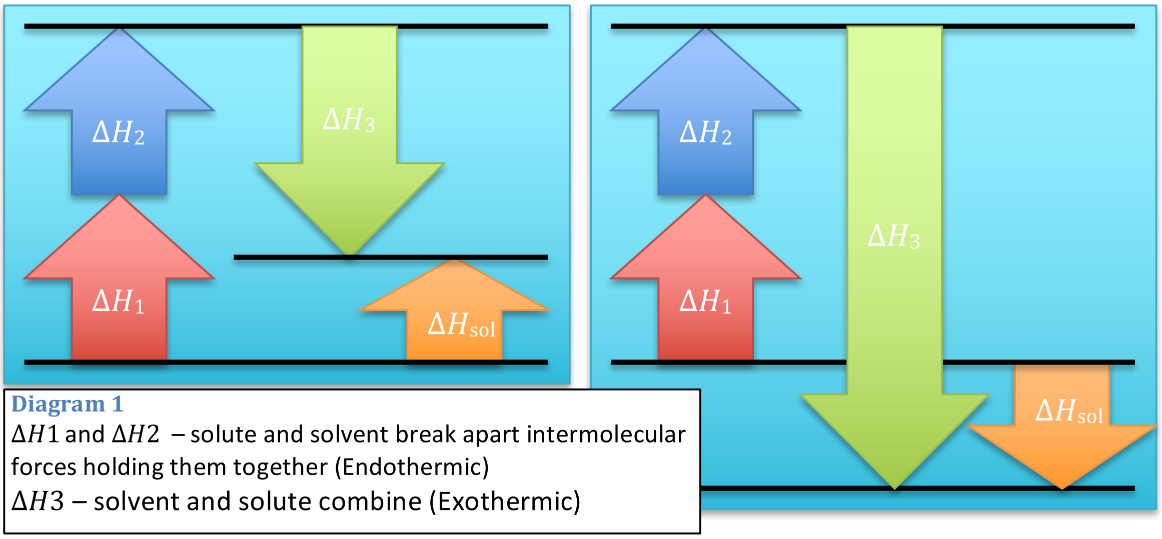


Diagram 1
$\Delta H1$ and $\Delta H2$ – solute and solvent break apart intermolecular forces holding them together (Endothermic)
$\Delta H3$ – solvent and solute combine (Exothermic)

The enthalpy of a solution is affected by three main factors (three other enthalpies) that depend on the solvent and the solute and their interaction with each other. If the energy required to break the intermolecular forces in either the solvent or solute decreases, the enthalpy of solution becomes more exothermic or less endothermic.

Therefore, a solute or solvent with greater intermolecular forces than another will have a more exothermic enthalpy or a less endothermic enthalpy.[2]

Intermolecular forces are the attractive forces that operate between atoms of a molecule. In ionic compounds (as this experiment will be concerned with) they are known as electrostatic forces of attraction. The greater or stronger these electrostatic forces are, the more energy is required to separate the ionic compound. Breaking bonds is endothermic; therefore, the overall reaction becomes less exothermic because more net intake of energy takes place.[3]

Two main factors affecting the intermolecular forces of an ionic compound are –

1. Charge of Ions involved in the ionic compound
2. Ionic Radii of the elements involved

---

[1] Talbot, Christopher, Richard Harwood, and Christopher Coates. *Chemistry for the IB Diploma*. London: Hodder Murray, 2009. Print.

[2] Javeed, Zafir. "Enthalpy of Solution - ChemWiki." *Enthalpy of Solution*. UC Davis ChemWiki by University of California, n.d. Web. 09 Mar. 2013. <http://chemwiki.ucdavis.edu/Physical_Chemistry/Physical_Properties_of_Matter/Solutions/Page_Title>.

[3] "Intermolecular Forces." *Intermolecular Forces*. Georgia Southern University, n.d. Web. 09 Mar. 2013. <http://cosm.georgiasouthern.edu/chemistry/general/molecule/forces.htm>.

DESIGN

ASPECT 1: *DEFINING THE PROBLEM*

RESEARCH QUESTION

How does the strength of ionic bonds in various metal hydroxides of the first and second group elements, affect the enthalpy of solution when dissolved in water, while concentration, temperature and pressure are kept constant?

HYPOTHESIS

I hypothesize that as the strength of ionic bonds in compounds increase, the enthalpy of solution will become less exothermic or more endothermic. In other words, the enthalpy of solution will numerically increase because now more energy is needed to separate the ionic compounds. As it absorbs more energy, the net endothermic part of the reaction increases (it needs to absorb more energy). This results in the sum of all endothermic and exothermic forces of the reaction to tend more towards endothermic (increases numerically as endothermic reactions are noted by a positive enthalpy of reaction.

Firstly, I hypothesize, that when it comes to changing the cation, the smaller ions like Sodium (Na^+) will have a more endothermic reaction than larger ions like Potassium (K^+). Since they both have the same charge, the main factor that will affect the enthalpy of solution would be the size of the atom.

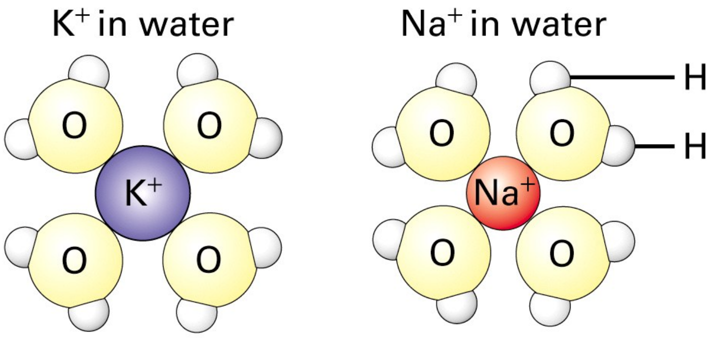


Diagram 2

As seen in Diagram 2, the sodium ion is much smaller in size compared to the potassium ion. The electrostatic forces between sodium and its compounds is higher than that of potassium and its compounds because its nucleus is closer to the anion's electrons forming a stronger bond due to higher attractive forces. Though potassium is more reactive than sodium due to its larger ionic radius, in its compounds, potassium is easier to separate than sodium since in sodium, the distance between ion centers is lesser which leads to greater attraction between the cation (Na^+) and the anion (constant in this case).

This follows for Ca^{2+} and Mg^{2+} where Mg^{2+} will have a more endothermic reaction when compared to Ca^{2+} with water. Since Magnesium has a stronger lattice with a constant anion, the lattice enthalpy is much greater which in turn causes its reaction (mainly enthalpy of solution) to be more endothermic in comparison to Calcium.

Secondly, I hypothesize that when the charge of the cation is different (Na^+ and Mg^{2+}), then, the enthalpy of solution of Mg^{2+} will be higher. Since Mg^{2+} and Na^+ are in the same period and are consecutive, there is very less difference between their ionic radius. But, magnesium has a +2 charge in its compounds compared to sodium's +1. The higher charge leads to a stronger attractive force between ions that are harder to separate. More energy is required to separate them.

So, I hypothesize that Magnesium's enthalpy of solution will be more endothermic that Sodium's because Magnesium has a greater charge which leads to stronger lattices in its ionic compounds. It takes more energy to separate the compound. This higher intake of energy is endothermic which results in a more endothermic enthalpy of solution.

The above two hypothesis' assume that the anion is constant and only the cation changes. Evaluating the change when both cation and anion are changed would be impossible because having 2 independent variables changed at the same time makes it impossible to know what change of independent variable caused which change in dependent variable. Here the independent variable will be the change of cation. The dependent variable or the variable I will be measuring making all these changes will be the enthalpy of solution and other suitable reactions.

Changing the anion of the compound will have similar affects compared to changing the cation. The more electronegative the atom of the ion (the more attractive the negative ion), the more energy it will require for the compound to be separated in order to dissolve it in a solvent.

So, even when the negative ion is changed, the ions having higher attractive forces (F^-) will have a more endothermic enthalpy compared to less attractive ions (Cl^-).

I will keep the negative ion constant. I will also control other factors such as the size of particles, concentration of solutions used and thermodynamic conditions.

ASPECT 2: *CONTROLLING VARIABLES*

STATEMENT and CONTROL OF VARIABLES

Independent Variable –

1. The compound (4 different hydroxides) – The independent variable is the variable that I will change in the experiment. Since the purpose of my experiment is to determine the affects on enthalpy of solution due to strength of ionic bonds, I will attempt to change the strength of the ionic bonds in an ionic compound.

I will change the cation of the compound in the reactions planned. This will allow me to form a relationship between the cation. The elements I will be using as the cations will be Sodium, Potassium, Magnesium and Calcium when possible. Hence, I will use the hydroxides of these 4 metals and conduct the same reaction for all 4 of these hydroxides. By comparing the different enthalpies, a relationship will be formed based on the strength of bonds in those 4 compounds.

Dependent Variable –

1. Temperature (heat evolved or absorbed by the reaction) - The dependent variable is the variable I expect to change due to changing the independent variable. This is also the variable I will be measuring. For the experiments, I will be mainly measuring the enthalpy of solution. For some types of compounds, I will also be measuring the enthalpy of other suitable reactions. For example, for the hydroxides, I will be measuring the enthalpy of neutralization as a solution.

Subtracting the enthalpy of neutralization from this will also yield the enthalpy of solution. This will be used as an error checking technique to affirm the enthalpy of solution derived from other reactions. This is measured to 2 decimal places significance level with a temperature probe, every 5 seconds.

Controlled Variables –

1. Size of Solid Particles - The size of particle of the solid (powdered, chips or blocks) will not directly affect the enthalpy of any reactions but it does affect the rate. A change in particle size will make some reactions end faster and others slower. The slower the reaction is, the more heat is lost to the surrounding which leads to a lower reading of enthalpy and more uncertainty. If some reactions are slower and others faster, then in some reactions, more heat will be lost and in others less. Since in this experiment, the comparison of enthalpies is more important than the absolute value, I will be aiming to keep the rate of the reaction constant. Therefore, I will use all solids as small chips. All solid particles are mashed up and passed through a sieve to confirm that sizes are congruent.

2. Concentration of Liquid Reagents - I will also be keeping the concentration of liquid reagents constant. An increase in concentration will affect he rate of the reaction. Additionally, a change in concentration could affect the enthalpy of a solution depending on which reagent is the limiting reagent. If the concentration of the limiting reagent is increased, there will be more enthalpy change. I will use reagents of $1 moldm^{-3}$.

3. Temperature of Surroundings - The temperature and pressure of the surroundings will be kept constant because either can cause a change in the enthalpy. The higher the surrounding temperature, the more endothermic the reaction will tend to be. Even when the reaction is exothermic, less heat will be lost to the surroundings due to a harsher gradient of heat energy. The surrounding is already saturated with heat so it will be harder for it to absorb the energy released by the solution.

4. Pressure of Surroundings - A change in surrounding pressure will change the rate of the reaction that in turn will indirectly affect the enthalpy of the reactions.

5. Number of Reactions (Hess' Law) *(read below for further explanation)* - For most compounds, the enthalpy of solution can be found directly. For others, I will have to use Hess' Law following a number of reactions to find out the enthalpy of the reaction that I want. However, I will have to keep the number of reactions constant. The more the number of reactions following the Hess' Law the more uncertainty and lower the enthalpy will be compared to the actual value. Each step involves heat loss. Therefore, the more the number of steps, the more heat loss occurs. Though the absolute value doesn't matter, the more the number of reactions, the more error there will be and less reliability when comparing data. Therefore the number of steps to the final reaction has to be kept constant.

6. Physical State of Reactants and Products - Lastly, the physical state of the reactants and products makes a huge and monumental difference to the calculated enthalpy. If in one reaction the reactant is a solid and in other a gas, then the reaction with the gaseous reactant will yield and more exothermic enthalpy. In the reaction with the solid reactant, energy is absorbed to convert the reactant into its gaseous state. This endothermic change is absent from the reaction with the already gaseous reactant.

Similarly, when the reaction yields a gaseous product, some energy has been used to convert the otherwise solid product to a gaseous product. Both of the cases above could lead to erroneous data readings. Therefore, all reactants will be in solid and liquid state (one liquid and one solid) and all products will be kept constant in solution state.

7. Insulation - The insulation directly affects the enthalpy change during experiments. It determines the amount of energy conserved or lost. Keeping this constant will make my experiment more precise but wont affect accuracy. I will try to insulate the polystyrene cup as well as possible and then keep the insulation constant.

8. Anion of Compound - I will also be keeping the anion constant. I will use hydroxides (hydroxide ion) for this experiment. Changing anion of the compound leads to a change in behavior and properties of the compound. It might also behave as a covalent compound instead. This would change the enthalpy values. Therefore, I will keep it constant.

Additional Information

I will conduct all the experiments in the same surrounding on the same day. I will use all the same apparatus (cleaned between trials and experiments) in order to keep uncertainty and other factors constant.

The mass of the reactant is not variable. Enthalpy change of a solution is measured based on the energy evolved or absorbed per mole of the reactant (independent variable). Therefore, the mass only has to be noted, but the end value will be irrespective of mass, as it will convert all masses to mole and all values to kJ/mol. I will still try to keep the mass about constant for linearity of data.

ASPECT 3: *DEVELOPING A METHOD FOR COLLECTION OF DATA*

APPARATUS

1. Beaker (1000 ml)
2. Polystyrene Cup as container for reaction
3. Cotton for Insulation
4. Data Logger with Temperature Probe ±0.01°C
5. Graduated Measuring Cylinder (100 ml) to measure liquid reagents ±0.5ml
6. Liquid Reagents
    a. HCl – 1 mol dm^{-3}
    b. Distilled Water
7. Solid Reagents
    a. Hydroxides
8. Beakers for Liquid Reagent
9. Crucible for Solid Reactants
10. Electronic Mass Balance ±0.001g

DIAGRAM OF SETUP

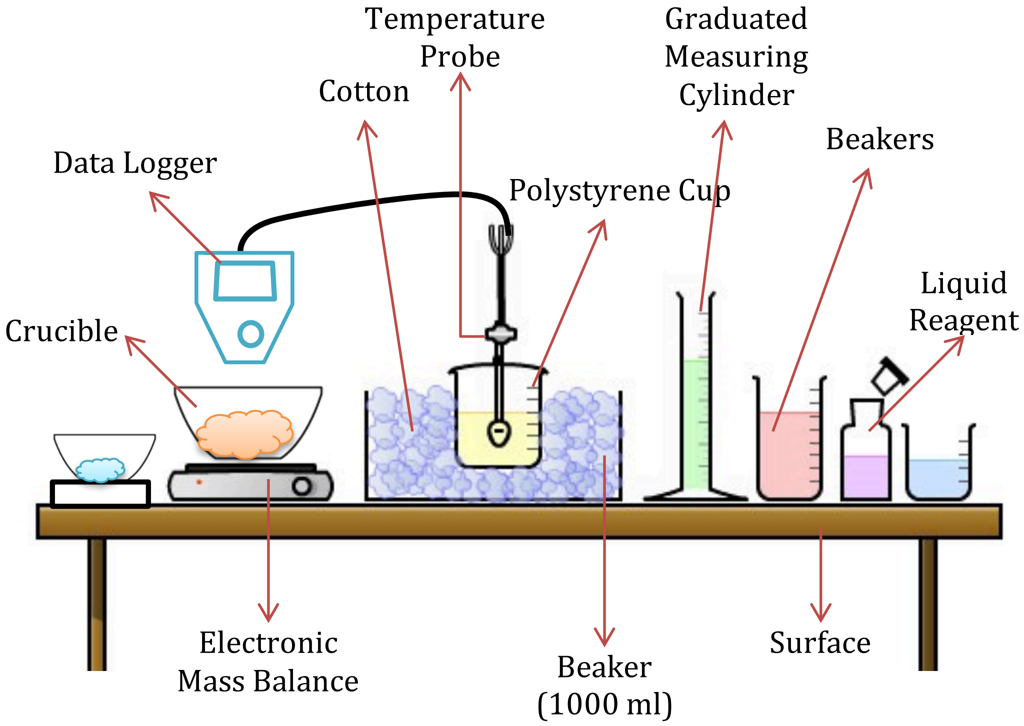


Diagram 3

PROCEDURE

1. First, I setup the Beaker (1000 ml), polystyrene cup and the insulation. This part of the set up will remain same throughout. I put the polystyrene cup inside the beaker and coated all sides with as much cotton as possible. Between trials and reactions, I would simply take the polystyrene cup out, clean it and reuse it again.

2. Then I added the cap with the temperature probe. I set the data logger to take temperature readings every 5 seconds. The temperature probe allowed the collection of sufficient data, as it could collect temperature data every 5 seconds to a certain accuracy. Taking readings every 5 seconds is very important since the exact progression of reaction is important when drawing and extrapolating graphs.

3. Then, I setup the distilled water. I used a constant of 60 cm^3. Hence, for every trial and experiment, I first measured out 60ml of distilled water using the measuring cylinder (the greatest source of uncertainty). This distilled water is then poured into the polystyrene cup.

4. Then, I set the temperature probe into the distilled water and started the probe. I measured about some amount of hydroxide, first starting with sodium hydroxide. I tried to take a constant of 5g of mass, even though it was not required, in order to directly compare results.

5. At exactly 60 seconds of the temperature probe taking data, I added the solid to the polystyrene cup. Then I allowed the temperature probe to continue taking data for a total of 300 seconds. Later, graphs will be extrapolated from the cooling graph to 60 seconds, to compensate for part of the heat loss to surrounding.

6. When I put the mass into the distilled water, I closed the cap of the cup as fast as possible and tried to limit the heat loss to surrounding by further covering the cup with cotton insulation.

7. Then I cleaned the polystyrene cup and did more trials. By repeating the steps described above. I further repeated the experiment for four hydroxides. Measurements of mass of solid were conducted using the crucible and an electronic mass balance, to great precision.

Safety and Precautions Taken -

1. Since heat was evolved in this reaction, I wore eye goggles to protect from the heat and the unknown amount of heat that might be evolved by dissolving the mass in the distilled water.

2. Enthalpy of solutions were first calculated theoretically by using data of other reactions and by using Hess Law, to make sure that reactions were safe and would not lead to temperatures above 60°C that could be potentially dangerous.

3. Hence, only 2.5g of potassium hydroxide as opposed to 5g of other reactants were used, as potassium hydroxide reacts more exothermically according to calculations and 5g could have lead to large spikes in temperature.

Formula's Used and Data Presented –

1. The temperature data as opposed to time collected by the temperature probe will first be graphed and the garphs will be extrapolated, to find the change in temperature, with minimum heat loss.
2. Once this change in temperature is calculated, the heat energy evolved due to this change in temperature will be calculated in joules through the equation –
$$\Delta H = m \cdot c \cdot \Delta T$$
3. After the heat energy change was calculated for each trial and all hydroxides, this heat energy had to be further changed to standard heat change or enthalpy for the solution, so that data could be compared and analyzed. This is done through the equation –
$$\text{Standard Enthalpy} = \frac{\text{Heat Evolved}}{\text{No. of Moles } (n)}$$
Where number of moles refers to the moles of solid, of either hydroxides, were dissolved in the distilled water. The number of moles can be calculated using stoichiometric equations –
$$n(\text{moles}) = \frac{\text{mass}}{\text{molar mass}}$$
where mass was also measured in some step of the procedure.

4. It is this standard enthalpy of solution which is finally presented.

DATA COLLECTION PROCESSING

ASPECT 1: *RECORDING RAW DATA*

DATA ON HYDROXIDES

Hydroxides Used –

1. $NaOH$
2. KOH
3. $Ca(OH)_2$
4. $Ba(OH)_2$

REACTION

This reaction involves reacting metal hydroxides with water. When metal hydroxides are mixed with water, they disassociate into metal and hydroxide ions. This dissociation leads to a change in enthalpy. This enthalpy change could be positive or negative. In other words, the absolute temperature could either increase or decrease.

This change in enthalpy will be calculated in order to conclusively decide on the strength/reactivity of the cation in the compound.

$$NaCl_{(s)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

The enthalpy change of solution is the change that occurs when one mole of solute dissolves in large excess. In this experiment, one mole will not be dissolved, but the final enthalpy will be scaled up to be calculated.

Table 1: Mass and Volume readings for all reactions concerning hydroxides

		REACTION 1 (With Water)			
		Before*	After**	**Mass** ***	**Vol.** ****
$NaOH$	Trial 1	57.832	62.632	4.800	60
	Trial 2	65.134	69.684	4.550	60
	Trial 3	57.826	62.15	4.324	60
KOH	Trial 1	65.147	67.527	2.380	60
	Trial 2	57.828	60.190	2.362	60
	Trial 3	65.126	67.699	2.573	60
$Ca(OH)_2$	Trial 1	57.838	62.618	4.780	60
	Trial 2	57.826	62.719	4.893	60
	Trial 3	65.127	69.912	4.785	60
$Ba(OH)_2$	Trial 1	57.832	62.666	4.834	60
	Trial 2	57.831	62.649	4.818	60
	Trial 3	65.128	69.923	4.795	60
* The mass of the crucible (empty) – $(\pm 0.001)\ grams$					
** The mass of the crucible (with the solid) – $(\pm 0.001)\ grams$					
*** The mass of the solid – $(\pm 0.002)\ grams$					
**** Volume of Liquid Used – $(\pm 0.5)\ cm^3$					

Qualitative Data for measuring mass

1. Though it’s not necessary to take the same mass of solid, I approximately tried to take it as a constant.

2. I used two crucibles to make measuring the solids easier. However, the masses of the two crucibles were different. That however doesn’t affect my results because my experiment is only concerned with the change in mass.

3. The volume had the highest uncertainty. However, the best apparatus possible to measure $60cm^3$ of the liquid was the graduated measuring cylinder. I could have used the 10ml and 25ml pipette to make up 60 ml but that would take too long. So, I settled with the measuring cylinder.

4. The electronic mass scale was quite accurate but was almost too precise. Passing wind and minute atmospheric changes would change the reading. This proved problematic since in this experiment, precise measurements of mass are imperative.

NaOH

Table 2: Temperature Reading for Reaction 1 of NaOH

TIME	TRIAL 1 *	TRIAL 2 *	TRIAL 3 *
0	296.4	297.4	297.7
5	296.4	297.4	297.6
10	296.4	297.4	297.6
15	296.4	297.5	297.6
20	296.4	297.4	297.7
25	296.4	297.3	297.6
30	296.4	297.5	297.6
35	296.4	297.4	297.7
40	296.4	297.4	297.6
45	296.4	297.4	297.6
50	296.4	297.5	297.6
55	296.4	297.4	297.6
60	296.4	297.5	297.7
65	296.4	297.4	297.9
70	296.7	297.6	298.1
75	297.3	298.1	298.6
80	297.9	298.8	299.9
85	298.7	300.2	301.6
90	299.5	301.4	303.4
95	300.3	302.6	305.6
100	301.1	303.8	306.3
105	301.9	305.0	308.1
110	302.6	306.1	309.4
115	303.4	307.3	310.2
120	304.3	308.4	310.9
125	305.4	309.3	312.6
130	306.4	309.9	313.8
135	307.3	310.7	313.9
140	308.1	311.2	314.0
145	309.0	311.9	314.0
150	309.7	312.3	314.1
155	310.4	312.9	314.0
160	311.2	313.1	314.0
165	311.8	313.3	314.0
170	312.2	313.6	314.1
175	312.6	313.6	314.1
180	312.8	313.7	314.0
185	313.1	313.7	313.9
190	313.3	313.7	314.0
195	313.2	313.7	314.0
200	313.2	313.7	313.8
205	313.2	313.6	313.9
210	313.3	313.6	313.9
215	313.2	313.5	313.8
220	313.1	313.6	313.7
225	313.1	313.5	313.6
230	313.0	313.5	313.7
235	313.0	313.4	313.6
240	313.0	313.5	313.7
245	312.9	313.3	313.6
250	312.9	313.4	313.5
255	312.9	313.3	313.5
260	312.9	313.3	313.6
265	312.8	313.3	313.5
270	312.8	313.3	313.5
275	312.6	313.3	313.4
280	312.6	313.2	313.5
285	312.6	313.1	313.4
290	312.5	313.2	313.4
295	312.5	313.0	313.4
300	312.5	312.9	313.3
305	312.3	313.0	313.3
310	312.3	313.0	313.2
315	312.3	312.9	313.4
320	312.2	312.9	313.3
325	312.2	312.9	313.2
330	312.2	312.9	313.0
335	312.2	312.9	313.1
340	312.2	312.9	313.1
345	312.2	312.8	313.0
350	312.1	312.8	313.0
355	312.1	312.8	312.9
360	312.1	312.7	312.9
365	312.1	312.7	312.7
370	312.1	312.6	312.9
375	312.1	312.7	312.8
380	312.1	312.8	312.8
385	312.1	312.7	312.8
390	312.0	312.6	312.7
395	311.9	312.6	312.8
400	312.0	312.6	312.7
405	311.9	312.6	312.7
410	311.9	312.6	312.7
415	311.9	312.5	312.7
420	311.8	312.5	312.6
425	311.8	312.5	312.7
430	311.8	312.4	312.6
435	311.8	312.4	312.6
440	311.7	312.3	312.6
445	311.7	312.4	312.5
450	311.7	312.3	312.6
455	311.6	312.3	312.5
460	311.6	312.3	312.5
465	311.5	312.2	312.5
470	311.6	312.2	312.6
475	311.5	312.2	312.5
480	311.5	312.2	312.5
485	311.5	312.2	312.4
490	311.4	312.2	312.5
495	311.4	312.2	312.5
500	311.4	312.2	312.5
505	311.4	312.1	312.5
510	311.4	312.1	312.4
515	311.4	312.1	312.4
520	311.4	312.1	312.5
525	311.3	312.0	312.4
530	311.3	312.0	312.3
535	311.2	312.0	312.4
540	311.2	312.1	312.4
545	311.2	312.1	312.3
550	311.2	312.0	312.3
555	311.2	312.0	312.3
560	311.2	311.9	312.3
565	311.2	312.0	312.2
570	311.1	311.9	312.3
575	311.1	311.9	312.3
580	311.2	311.8	312.3
585	311.1	311.8	312.3
590	311.1	311.8	312.4
595	311.0	311.8	312.4
600	311.0	311.8	312.3

* Temperature Reading from Data Logger – ($K \pm 0.1$)

KOH

Table 3: Temperature Reading for Reaction 1 of KOH

TIME	TRIAL 1 *	TRIAL 2 *	TRIAL 3 *
0	298.3	297.7	298.6
5	298.4	297.7	298.7
10	298.5	297.7	298.6
15	298.4	297.7	298.6
20	298.3	297.6	298.6
25	298.4	297.7	298.6
30	298.5	297.7	298.7
35	298.6	297.7	298.7
40	298.4	297.6	298.7
45	298.4	297.6	298.6
50	298.5	297.6	298.7
55	298.6	297.6	298.7
60	298.4	297.7	298.8
65	298.6	298.4	300.3
70	299.2	301.6	301.9
75	301.0	305.4	302.7
80	303.1	307.3	304.3
85	304.8	306.2	304.6
90	305.8	306.1	306.1
95	306.2	306.2	306.4
100	306.5	306.2	306.9
105	306.6	306.2	306.7
110	306.5	306.3	306..9
115	306.5	306.2	306.8
120	306.6	306.2	306.9
125	306.6	306.3	306.9
130	306.6	306.3	306.8
135	306.5	306.3	306.7
140	306.5	306.3	306.6
145	306.5	306.2	306.7
150	306.6	306.2	306.6
155	306.5	306.2	306.7
160	306.5	306.1	306.7
165	306.5	306.2	306.6
170	306.5	306.2	306.6
175	306.4	306.2	306.7
180	306.5	306.2	306.6
185	306.5	306.1	306.6
190	306.5	306.1	306.5
195	306.5	306.1	306.5
200	306.5	306.2	306.6
205	306.5	306.1	306.5
210	306.5	306.1	306.5
215	306.4	306.1	306.5
220	306.4	306.1	306.5
225	306.3	306.2	306.5
230	306.3	306.1	306.5
235	306.3	306.1	306.4
240	306.3	306.0	306.3
245	306.4	306.1	306.5
250	306.4	306.0	306.4
255	306.3	306.1	306.4
260	306.4	306.0	306.3
265	306.4	306.1	306.2
270	306.2	306.1	306.1
275	306.4	306.1	306.2
280	306.3	306.1	306.2
285	306.2	306.0	306.1
290	306.2	306.0	306.2
295	306.2	306.1	306.1
300	306.2	306.0	306.1
305	306.2	306.1	306.0
310	306.2	306.0	306.1
315	306.3	306.0	306.0
320	306.2	306.1	305.9
325	306.2	306.0	306.1
330	306.1	306.0	306.0
335	306.2	306.0	305.9
340	306.1	306.0	305.9
345	306.2	306.0	305.9
350	306.1	306.0	305.9
355	306.2	306.0	305.9
360	306.1	306.0	305.9
365	306.2	306.0	305.9
370	306.1	306.0	305.9
375	306.1	305.9	305.8
380	306.2	306.0	305.9
385	306.2	306.1	305.9
390	306.0	306.0	305.8
395	306.0	305.9	305.9
400	306.1	306.0	305.8
405	306.1	305.7	305.7
410	306.0	305.9	305.6
415	306.1	305.9	305.8
420	306.1	305.9	305.8
425	306.0	305.9	305.7
430	305.9	305.8	305.6
435	306.1	305.9	305.7
440	305.9	305.8	305.7
445	306.0	305.8	305.6
450	306.1	305.7	305.7
455	305.9	305.8	305.7
460	305.9	305.9	305.6
465	306.0	305.8	305.6
470	306.0	305.8	305.7
475	306.0	305.9	305.6
480	306.0	305.8	305.6
485	305.9	305.8	305.6
490	305.8	305.8	305.6
495	306.0	305.8	305.6
500	305.8	305.9	305.5
505	305.8	305.8	305.5
510	306.0	305.7	305.5
515	305.9	305.8	305.5
520	305.9	305.7	305.5
525	306.0	305.8	305.4
530	305.9	305.7	305.5
535	305.9	305.7	305.4
540	305.7	305.7	305.4
545	305.9	305.8	305.5
550	305.9	305.7	305.4
555	305.7	305.9	305.4
560	305.9	305.7	305.3
565	305.7	305.7	305.5
570	305.8	305.7	305.3
575	305.7	305.6	305.3
580	305.8	305.7	305.3
585	305.7	305.6	305.3
590	305.8	305.5	305.3
595	305.7	305.7	305.2
600	305.7	305.5	305.3

* Temperature Reading from Data Logger – $(Kelvin \pm 0.1)$

Ca(OH)$_2$

Table 4: Temperature Reading for Reaction 1 of Ca(OH)$_2$

TIME	TRIAL 1 *	TRIAL 2 *	TRIAL 3 *
0	298.3	297.9	298.1
5	298.4	297.9	298.1
10	298.3	297.9	298.1
15	298.3	297.9	298.1
20	298.3	297.9	298.1
25	298.3	297.9	298.2
30	298.4	297.9	298.1
35	298.3	297.9	298.1
40	298.3	297.9	298.2
45	298.2	297.9	298.2
50	298.3	297.9	298.1
55	298.4	297.9	298.1
60	298.4	297.9	298.1
65	298.3	297.9	298.2
70	298.8	298.8	298.5
75	298.8	298.9	298.6
80	298.2	298.8	298.6
85	298.6	298.9	298.6
90	298.6	298.9	298.6
95	298.6	298.8	298.6
100	298.6	298.8	298.6
105	298.6	298.8	298.6
110	298.6	298.8	298.6
115	298.6	298.8	298.6
120	298.8	298.8	298.6
125	298.6	298.8	298.6
130	298.6	298.7	298.6
135	298.6	298.6	298.6
140	298.6	298.8	298.6
145	298.5	298.8	298.6
150	298.6	298.7	298.6
155	298.7	298.7	298.5
160	298.7	298.7	298.6
165	298.6	298.7	298.6
170	298.6	298.7	298.6
175	298.7	298.7	298.6
180	298.6	298.6	298.5
185	298.6	298.7	298.6
190	298.7	298.7	298.6
195	298.7	298.6	298.6
200	298.6	298.7	298.6
205	298.6	298.6	298.6
210	298.7	298.7	298.6
215	298.8	298.6	298.5
220	298.6	298.7	298.5
225	298.6	298.6	298.6
230	298.7	298.6	298.5
235	298.6	298.6	298.5
240	298.6	298.6	298.6
245	298.8	298.7	298.5
250	298.6	298.6	298.5
255	298.6	298.6	298.6
260	298.5	298.6	298.5
265	298.7	298.6	298.5
270	298.5	298.5	298.5
275	298.6	298.5	298.6
280	298.5	298.5	298.5
285	298.6	298.4	298.5
290	298.6	298.5	298.5
295	298.5	298.5	298.5
300	298.6	298.6	298.5
305	298.6	298.6	298.6
310	298.6	298.6	298.5
315	298.6	298.6	298.5
320	298.5	298.6	298.5
325	298.6	298.5	298.5
330	298.5	298.6	298.5
335	298.6	298.5	298.5
340	298.6	298.6	298.5
345	298.6	298.6	298.5
350	298.6	298.6	298.5
355	298.5	298.6	298.5
360	298.6	298.5	298.4
365	298.5	298.5	298.5
370	298.5	298.6	298.5
375	298.6	298.6	298.5
380	298.5	298.6	298.5
385	298.5	298.5	298.5
390	298.5	298.4	298.5
395	298.6	298.4	298.4
400	298.5	298.4	298.4
405	298.5	298.5	298.5
410	298.5	298.4	298.4
415	298.5	298.5	298.5
420	298.5	298.5	298.4
425	298.5	298.4	298.5
430	298.5	298.4	298.5
435	298.5	298.4	298.5
440	298.5	298.4	298.5
445	298.5	298.4	298.4
450	298.5	298.4	298.5
455	298.5	298.4	298.5
460	298.5	298.4	298.5
465	298.5	298.4	298.5
470	298.5	298.4	298.5
475	298.5	298.3	298.4
480	298.5	298.4	298.5
485	298.5	298.4	298.5
490	298.4	298.4	298.5
495	298.4	298.4	298.4
500	298.4	298.4	298.4
505	298.4	298.3	298.4
510	298.4	298.3	298.4
515	298.4	298.4	298.4
520	298.4	298.4	298.4
525	298.5	298.3	298.4
530	298.5	298.4	298.4
535	298.4	298.3	298.4
540	298.5	298.3	298.4
545	298.5	298.3	298.4
550	298.4	298.4	298.4
555	298.5	298.4	298.3
560	298.4	298.3	298.4
565	298.4	298.3	298.4
570	298.4	298.3	298.4
575	298.4	298.3	298.3
580	298.4	298.3	298.4
585	298.4	298.4	298.4
590	298.4	298.3	298.4
595	298.4	298.3	298.3
600	298.4	298.3	298.4

* Temperature Reading from Data Logger – (K ± 0.1)

$Ba(OH)_2$

Table 5: Temperature Reading for Reaction 1 of $Ba(OH)_2$

TIME	TRIAL 1 *	TRIAL 2 *	TRIAL 3 *
0	298	297.9	297.6
5	298.1	297.9	297.6
10	298.1	297.9	297.6
15	298.1	297.8	297.6
20	298.1	297.9	297.6
25	298.1	297.9	297.6
30	298.1	297.9	297.6
35	297.9	297.9	297.6
40	297.9	298.0	297.6
45	298	297.9	297.6
50	298.1	297.9	297.6
55	297.9	297.9	297.6
60	298	297.9	297.6
65	297.8	297.2	297.6
70	296.3	295.6	295.9
75	295.4	294.5	294.8
80	295.3	294.4	294.7
85	295.1	294.5	294.8
90	295.2	294.4	294.8
95	295.4	294.5	294.8
100	295.4	294.4	294.8
105	295.4	294.5	294.8
110	295.4	294.4	294.5
115	295.4	294.5	294.6
120	295.3	294.5	294.7
125	295.5	294.5	294.7
130	295.3	294.5	294.8
135	295.4	294.5	294.8
140	295.4	294.4	294.7
145	295.4	294.5	294.7
150	295.4	294.5	294.7
155	295.4	294.5	294.8
160	295.4	294.6	294.8
165	295.5	294.6	294.7
170	295.4	294.6	294.8
175	295.6	294.6	294.7
180	295.4	294.4	294.7
185	295.5	294.6	294.7
190	295.4	294.5	294.7
195	295.4	294.5	294.8
200	295.5	294.6	294.8
205	295.4	294.6	294.7
210	295.4	294.6	294.8
215	295.3	294.7	294.8
220	295.7	294.9	294.8
225	295.4	294.7	295.0
230	295.6	294.7	294.9
235	295.7	294.7	294.9
240	295.4	294.7	294.9
245	295.7	294.8	294.9
250	295.6	294.8	295.0
255	295.6	294.9	294.9
260	295.7	294.7	294.9
265	295.7	294.9	295.1
270	295.7	294.8	295.1
275	295.6	294.9	295.1
280	295.6	295.0	295.1
285	295.7	295.0	294.9
290	295.7	295.1	295.1
295	295.7	295.0	295.1
300	295.6	295.0	295.1
305	295.7	295.1	295.2
310	295.7	295.1	295.1
315	295.7	295.1	295.3
320	295.6	295.0	295.4
325	295.7	295.1	295.2
330	295.7	295.1	295.4
335	295.6	295.0	295.3
340	295.7	295.1	295.3
345	295.7	295.0	295.3
350	295.7	295.1	295.3
355	295.7	295.1	295.3
360	295.7	295.1	295.4
365	295.7	295.1	295.4
370	295.7	295.3	295.4
375	295.9	295.4	295.4
380	295.7	295.3	295.4
385	295.8	295.3	295.4
390	295.6	295.3	295.5
395	295.9	295.4	295.5
400	295.7	295.2	295.5
405	295.7	295.3	295.5
410	295.7	295.4	295.5
415	295.7	295.3	295.5
420	295.8	295.3	295.6
425	295.8	295.3	295.5
430	295.8	295.4	295.6
435	295.8	295.4	295.5
440	295.8	295.4	295.6
445	295.9	295.4	295.5
450	295.9	295.6	295.6
455	295.8	295.4	295.6
460	295.9	295.5	295.6
465	295.8	295.5	295.5
470	295.9	295.5	295.6
475	295.8	295.4	295.7
480	295.9	295.5	295.9
485	295.9	295.5	295.6
490	295.9	295.5	295.7
495	295.9	295.5	295.8
500	295.9	295.7	295.6
505	295.9	295.6	295.8
510	295.9	295.7	295.7
515	295.9	295.4	295.8
520	296.0	295.6	295.7
525	295.9	295.6	295.8
530	296.0	295.5	295.7
535	296.0	295.6	295.8
540	296.0	295.6	295.7
545	296.0	295.6	295.8
550	295.9	295.6	295.8
555	296.0	295.6	295.8
560	296.2	295.7	295.8
565	296.1	295.6	295.8
570	296.0	295.7	295.8
575	296.2	295.6	295.9
580	296.1	295.7	295.8
585	296.2	295.7	295.8
590	296.1	295.7	295.9
595	296.1	295.6	295.9
600	296.1	295.7	295.9

* Temperature Reading from Data Logger – (K $\pm$ 0.1)

Qualitative Data for Conducting Experiment

1. The calcium hydroxide was only partially soluble in water. When stirring with stirrer, powdered solid could still be felt at the bottom of the cup. After experiment, when emptying the cup, I saw the calcium hydroxide powder was still present at the bottom of the polystyrene cup. Therefore, experiment wasn't and couldn't be completed until all reactant disappears

2. The initial temperature for a second and third trial and all consecutive experiments was a little (about 0.5 K) higher than the previous. This is because, there wasn't enough time to wait for the polystyrene cup to completely return to room temperature. Additionally, since most reactions conducted were exothermic, the temperature of the surrounding was affected mildly.

3. In Barium Hydroxide reaction, the temperature swings/more unpredictable. This may be due to the limited solubility of Barium Hydroxide.

4. For Reaction 2, I only conducted each experiment for 300 seconds instead of the previous 600 seconds and only 2 trials. I did this, because I noticed from the reactions I had already conducted, that the reaction had a stable and predictable curve coming back to room temperature. Therefore, 300 seconds were sufficient. Additionally, since this was a secondary experiment, doing this was appropriate and would allow more than enough opportunity to collect the required data

5. If a trial did not support the others conducted, I re conducted the reaction again in order to get both reliable and correct data. This went well for the second reactions that I conducted and made my results unquestionable.

6. Each solid had different solubility, therefore, there was a difference in both rate of reaction and heat evolved.

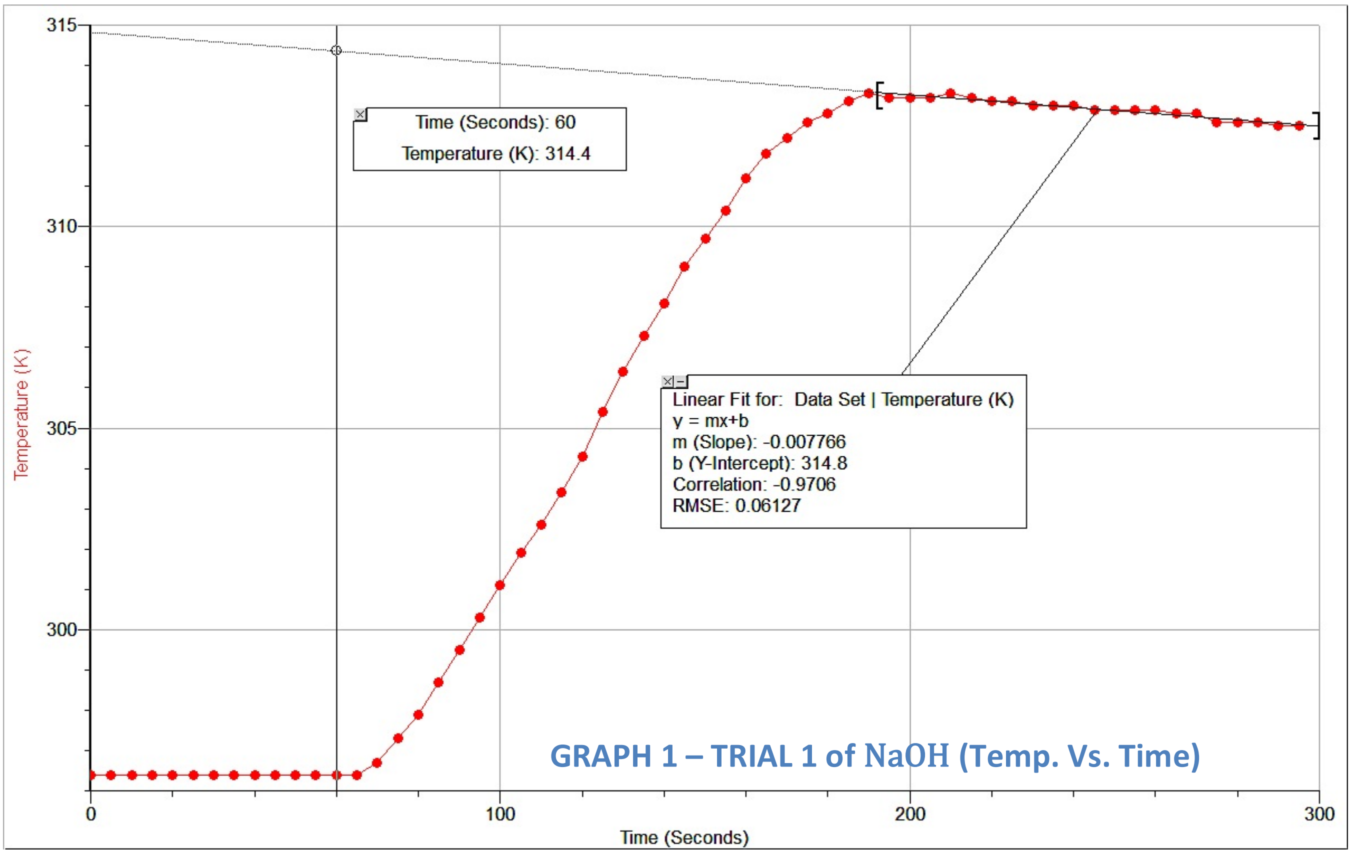
ASPECT 2 & 3: *PROCESSING and PRESENTING RAW DATA*

Using the data collected above, individual graphs were made for each trial, with temperature against time. Each graph was extrapolated, to calculate temperature with minimum energy lost to surroundings. Additionally, graphs were made with all three trials together, to evaluate both the accuracy and precision of readings relative to each other.

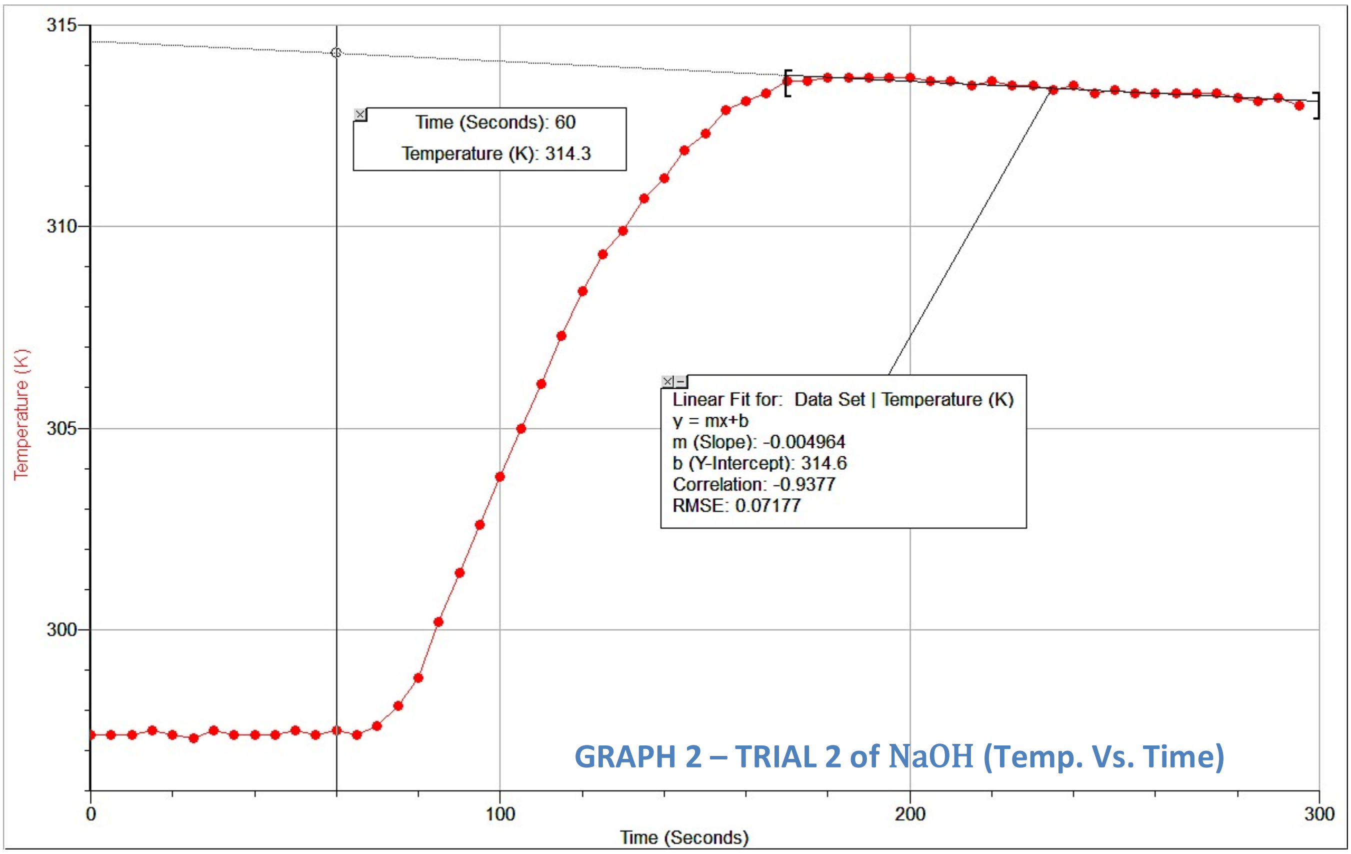
After each set of trials is extrapolated, a table of final maximum temperatures is created, to be used later in order to calculate enthalpy.

The linear fit described in the graph is the line that fits the data points when the liquid starts cooling or heating after attaining its maximum or minimum temperature respectively. It is this line that is extended to time=60s (since this was when the sample was added), and the corresponding value of temperature is found.

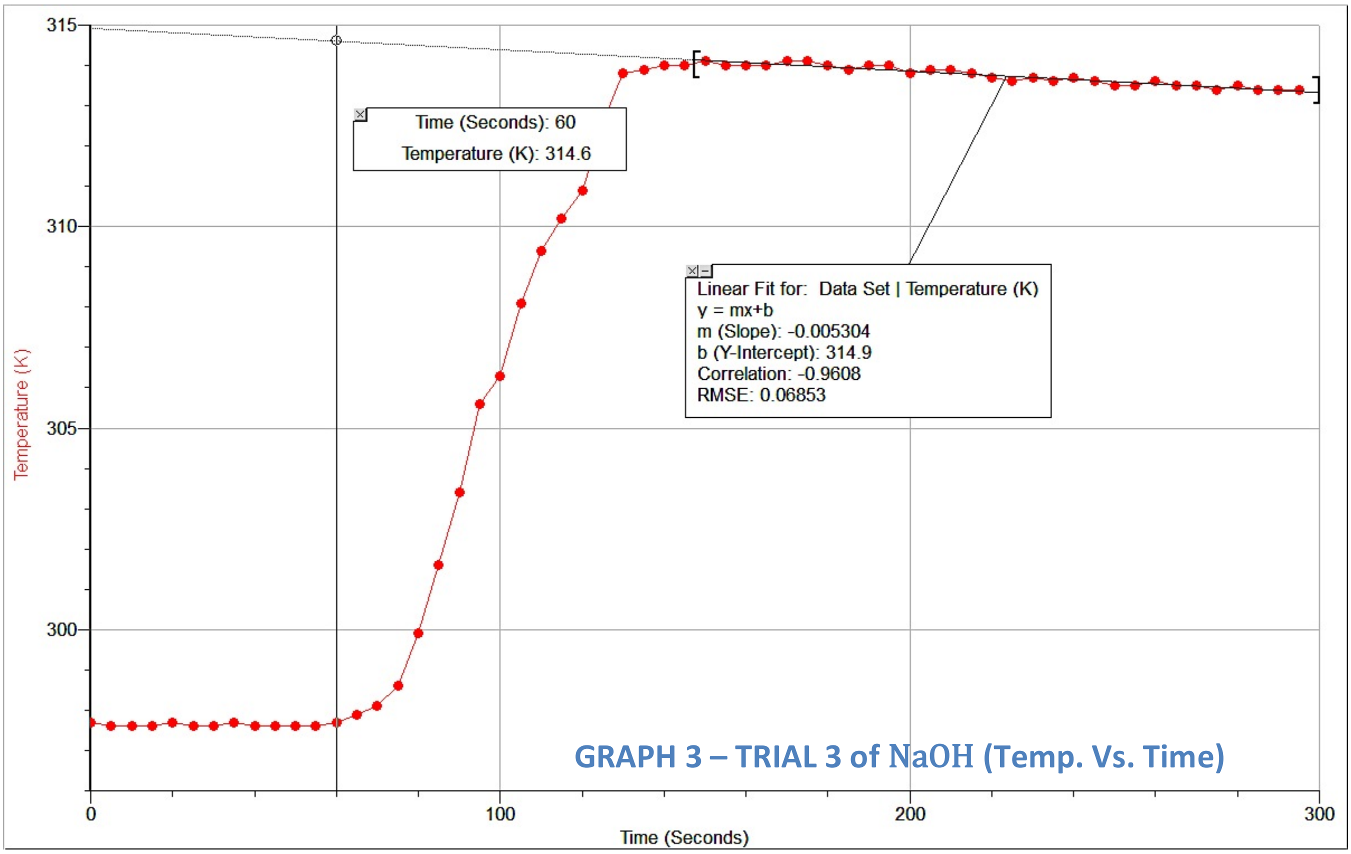
Graphs for NaOH



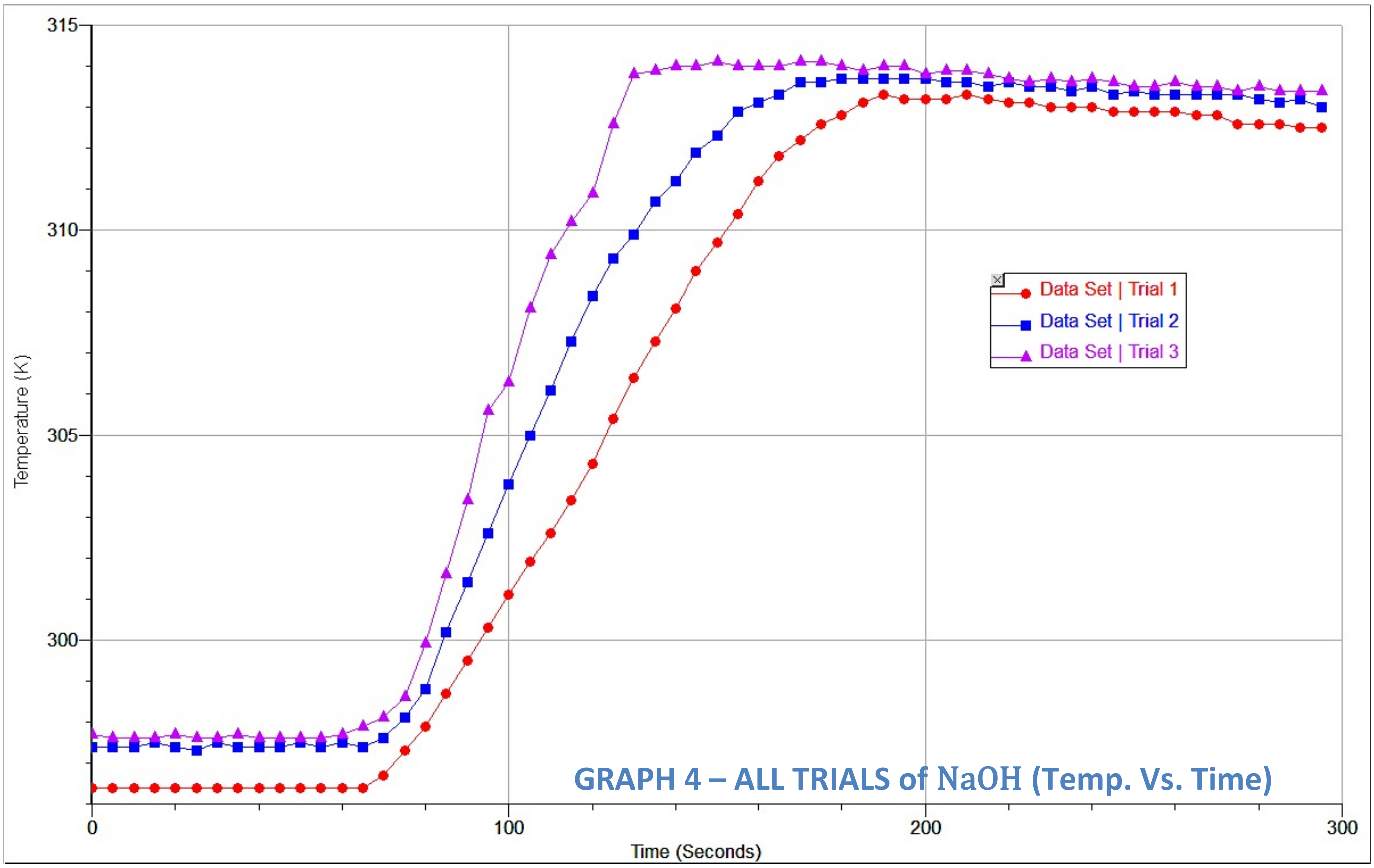
GRAPH 1 – TRIAL 1 of NaOH (Temp. Vs. Time)



GRAPH 2 – TRIAL 2 of NaOH (Temp. Vs. Time)



GRAPH 3 – TRIAL 3 of NaOH (Temp. Vs. Time)



GRAPH 4 – ALL TRIALS of NaOH (Temp. Vs. Time)

From the combined graphs, we can see that the “room temperature” or rather the container temperature at the start of the experiment between the three trails consecutively increased, but by small, and almost insignificant amounts. This variation will not lead to any great systematic or random errors because change of temperature is what will finally be used.

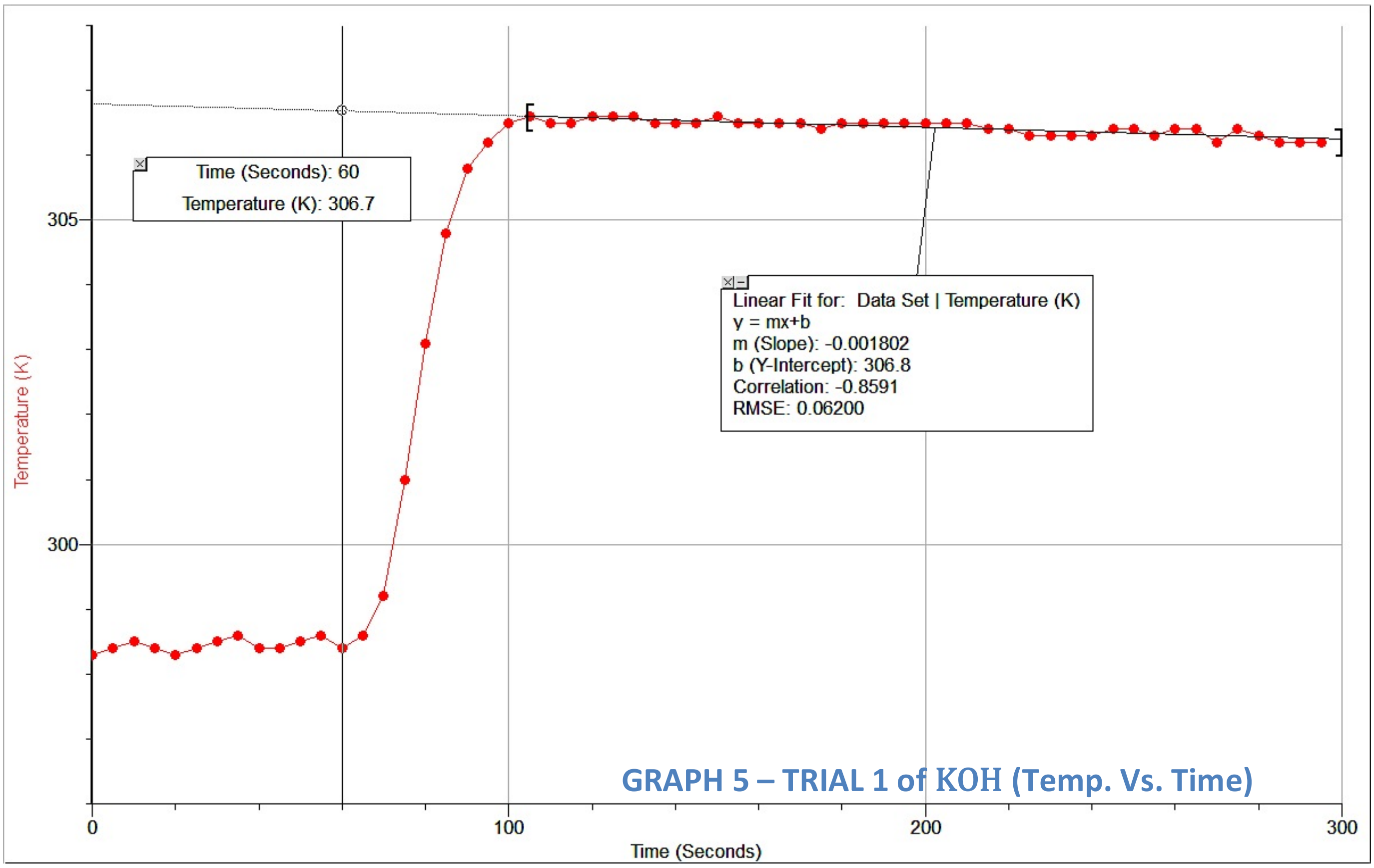
It can also be noticed that the three trials, reached about the same maximum temperature which means that the data is precise. However, the reactions took different times to reach this maximum, as seen by the three different curves. The 3rd trial especially took the smallest amount of time.

Finally singular graphs were extrapolated to further reduce the error caused due to loss of heat to surroundings. The data obtained is summarized below.

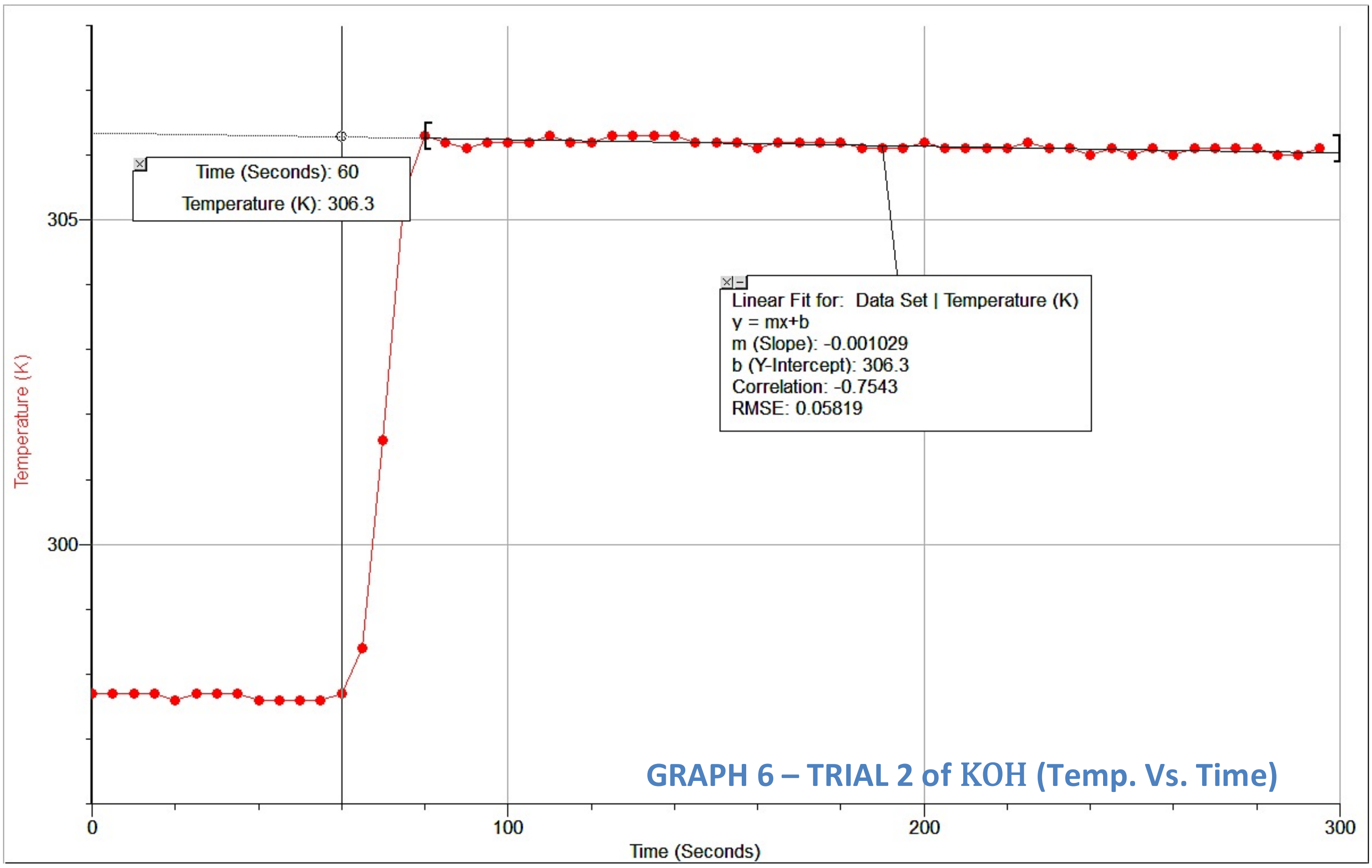
Summary Table –

	Max. Temp. Reached (K)	Room Temperature	Extrapolated Temp. (K)	Absolute Uncertainty	**Change in Temperature**
Trial 1	313.3	296.4	314.4	0.1℃	**18**
Trial 2	313.7	297.4	314.3	0.1℃	**16.9**
Trial 3	314.1	297.6	314.6	0.1℃	**17**
% Uncert.	≈ 0.25%	≈ 0.43%	≈ 0.25%		≈ **1**%

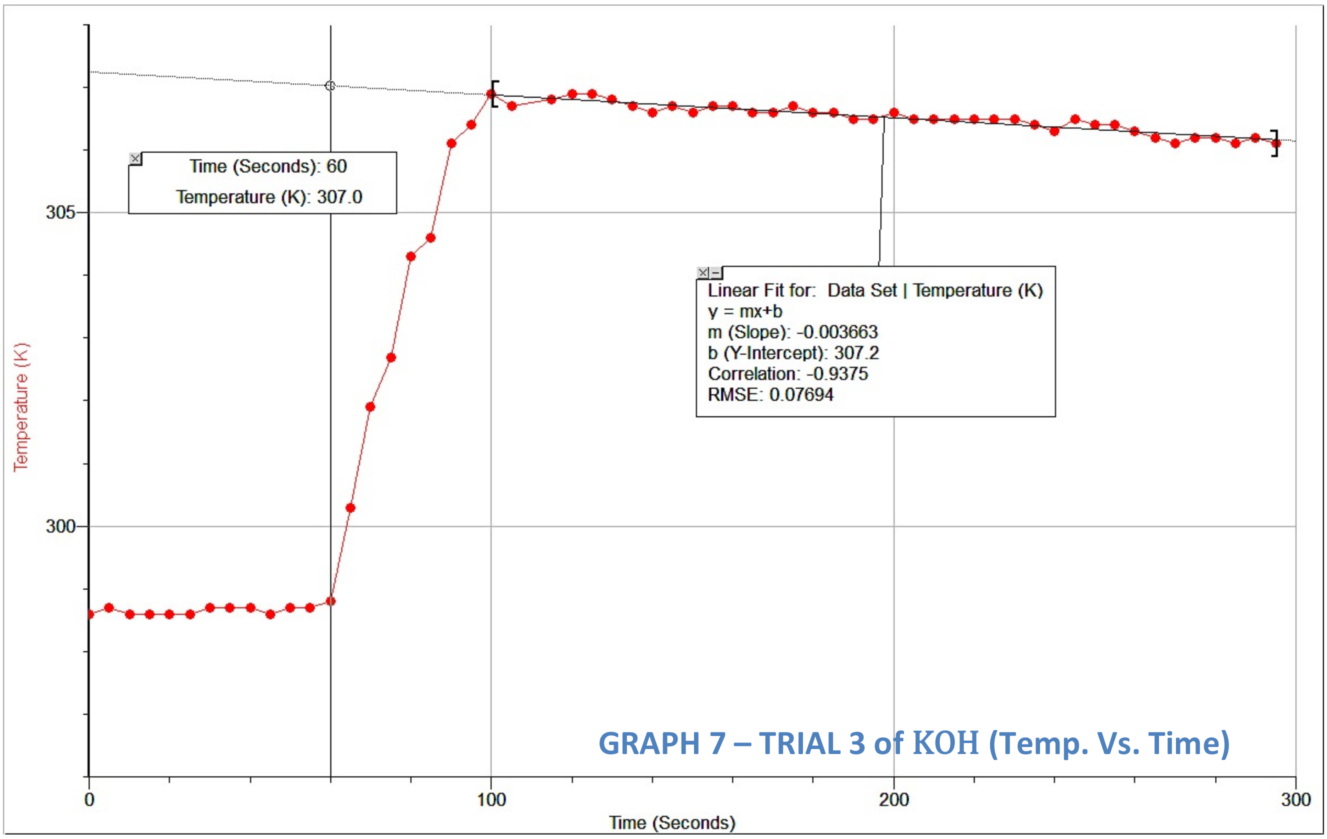
Graphs for KOH



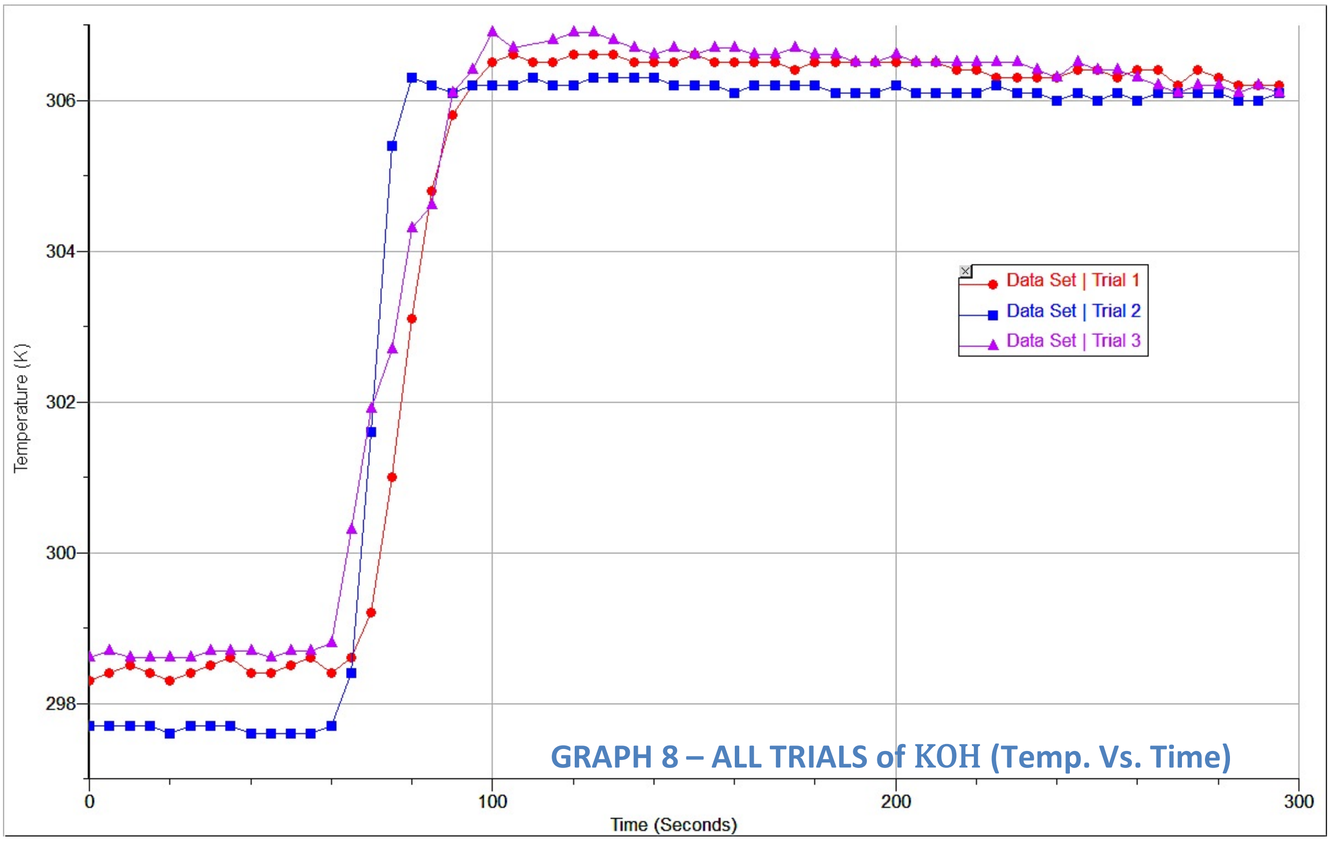
GRAPH 5 – TRIAL 1 of KOH (Temp. Vs. Time)



GRAPH 6 – TRIAL 2 of KOH (Temp. Vs. Time)



GRAPH 7 – TRIAL 3 of KOH (Temp. Vs. Time)



GRAPH 8 – ALL TRIALS of KOH (Temp. Vs. Time)

First we notice from the combined graphs that the room temperature was not exactly constant, however, again it does not make much of a difference since the change in temperature is being calculated here. However, it does create a slight difference, as the warmer the container is from room temperature, the less affect heat has on its final temperature, because at higher temperatures, heat loss has a greater effect.

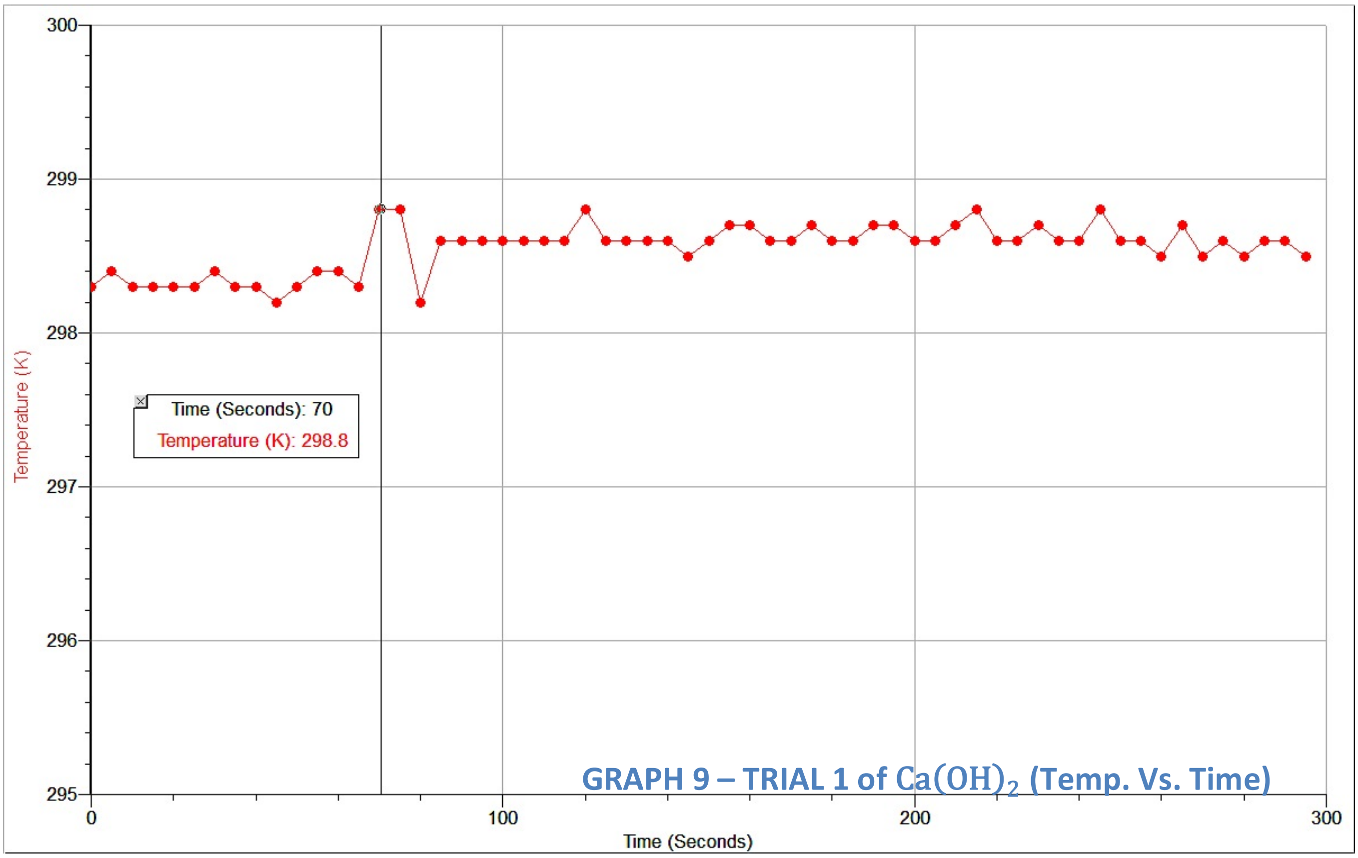
Since the temperatures reached in this experiment were lower than the one with NaOH, less heat was lost to the surrounding. This can also be seen by the minute change in maximum temperature and extrapolated temperature. Whereas in NaOH, the difference was about 0.4K, in this, the difference is one 0.1K.

Finally singular graphs were extrapolated to further reduce the error caused due to loss of heat to surroundings. The data obtained is summarized below.

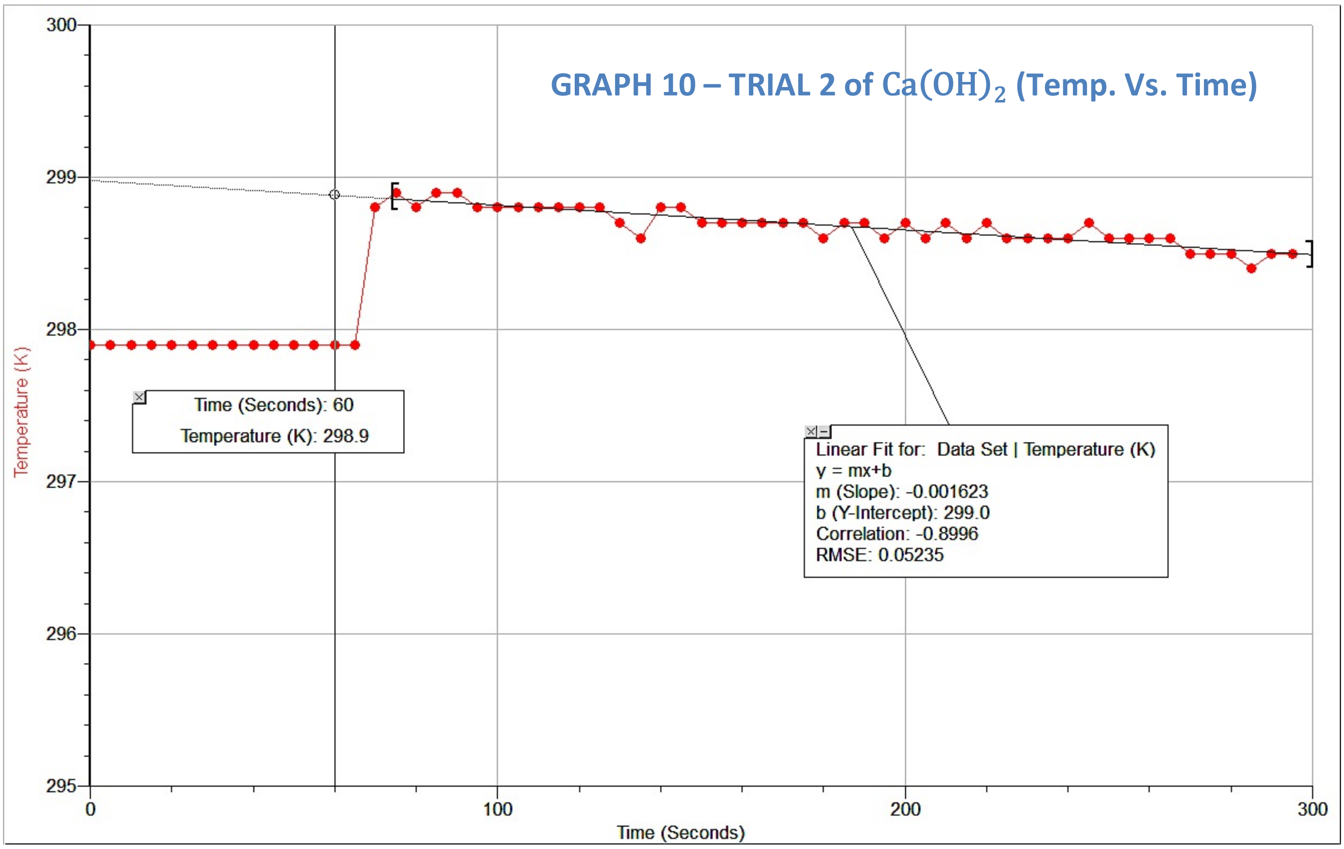
Summary Table –

	Max. Temp. Reached (K)	Room Temperature	Extrapolated Temp. (K)	Absolute Uncertainty	**Change in Temperature**
Trial 1	306.6	298.4	306.7	0.1℃	**8.3**
Trial 2	306.3	297.7	306.3	0.1℃	**8.6**
Trial 3	306.9	298.6	307.0	0.1℃	**8.4**
% Uncert.	≈ 0.30%	≈ 0.40%	≈ 0.30%		≈ **2.5**%

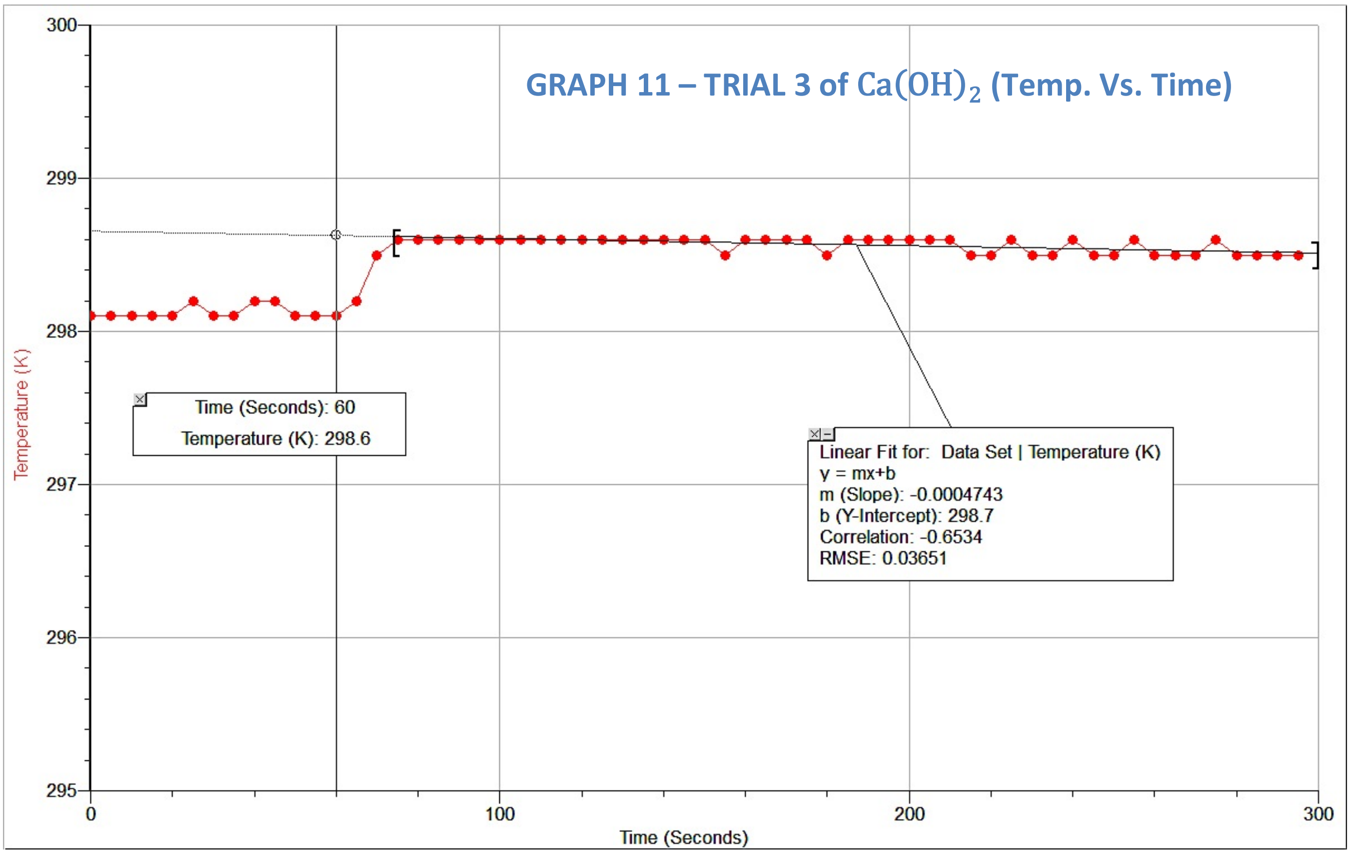
Graphs for $Ca(OH)_2$



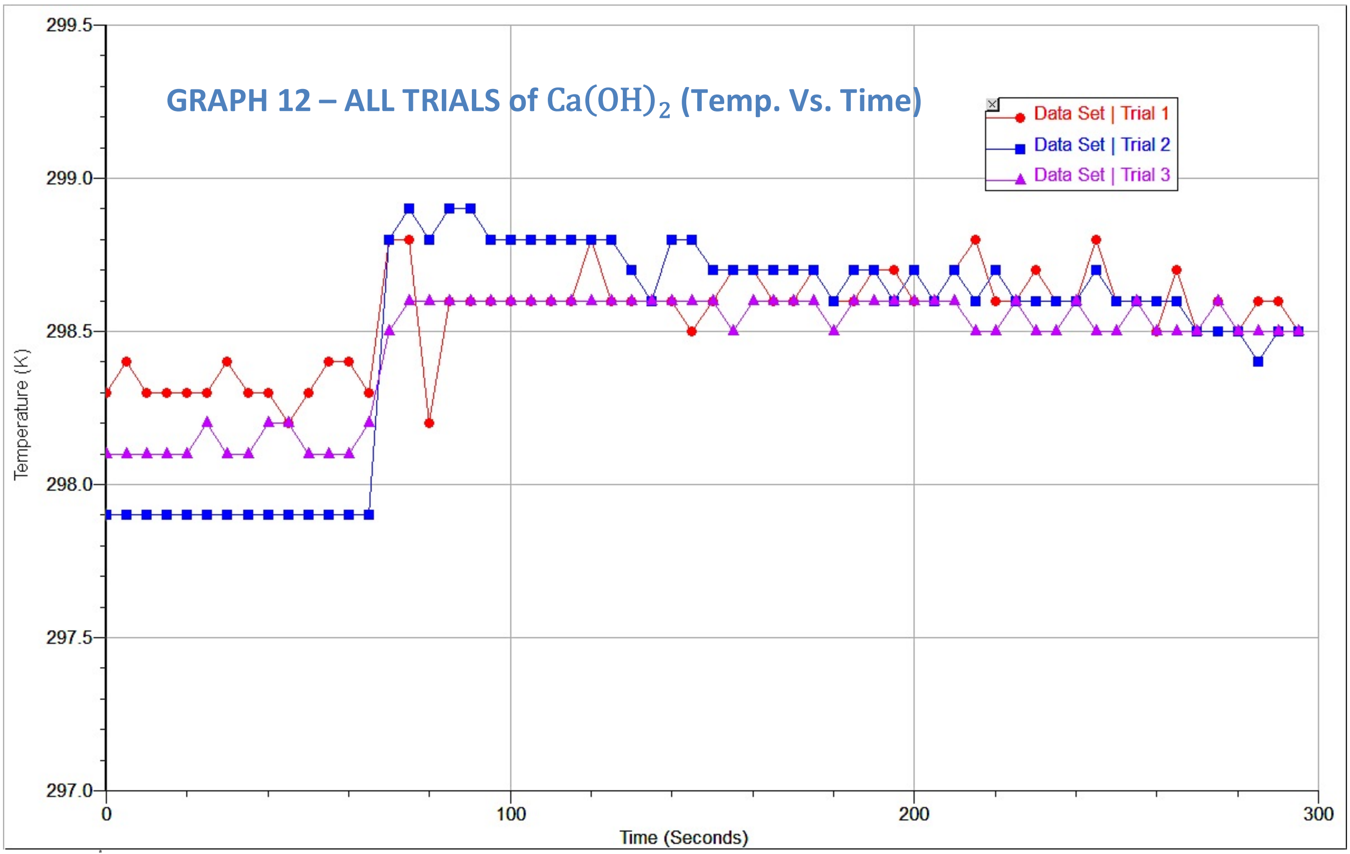
GRAPH 9 – TRIAL 1 of $Ca(OH)_2$ (Temp. Vs. Time)



GRAPH 10 – TRIAL 2 of $Ca(OH)_2$ (Temp. Vs. Time)



GRAPH 11 – TRIAL 3 of $Ca(OH)_2$ (Temp. Vs. Time)



The problem that was seen in this experiment was the limited solubility of calcium hydroxide. This proved to be a big problem, as it limited the change in temperature, so the rise in temperature was barely noticeable. Additionally, we don't know what percentage of the sample dissolved to result in that temperature. Hence, it has to be assumed that the whole sample reacted. This will add inaccuracy to enthalpy calculations conducted later.

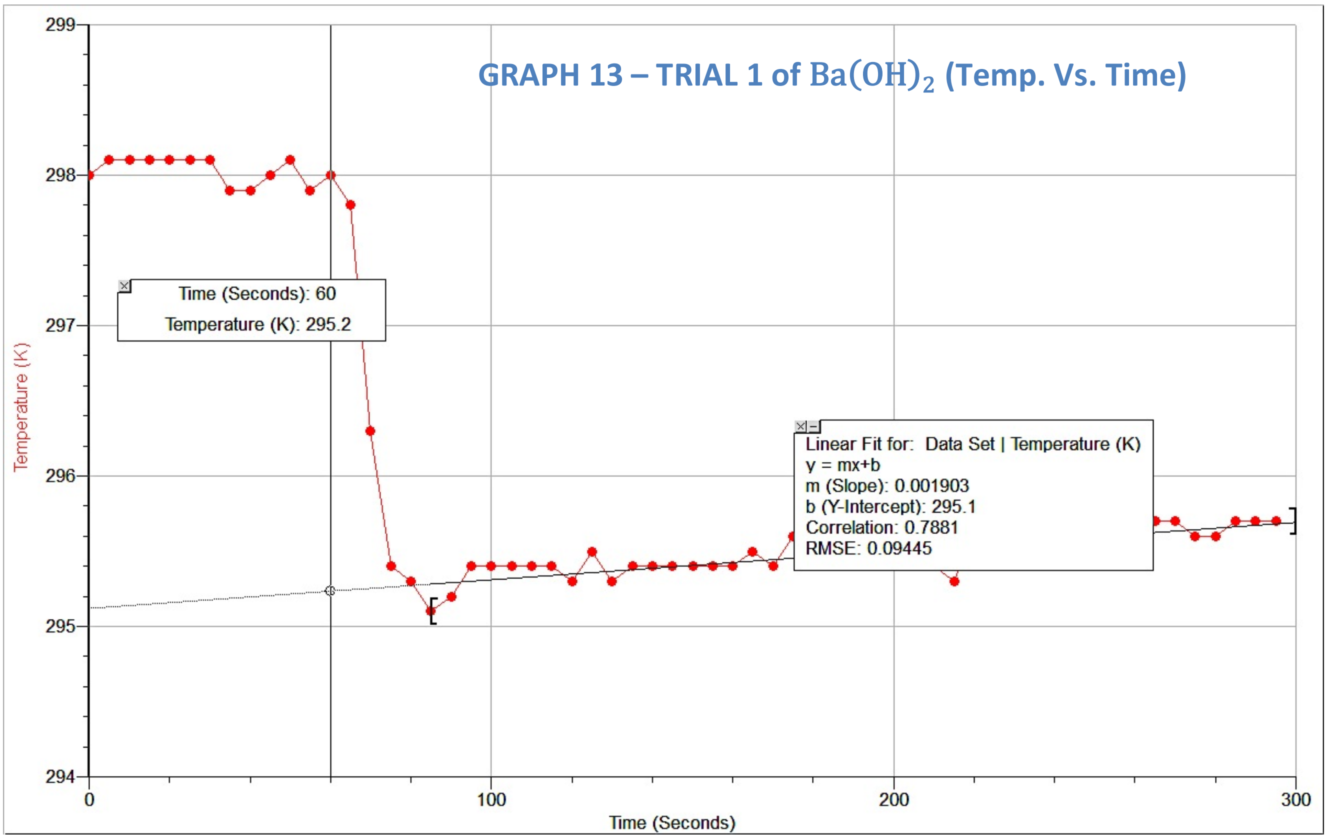
Since the temperatures reached in this experiment were even lower in this experiment compared to the previous two, the change from maximum temperature to extrapolated temperature is even lesser, often being exactly the same. This also proves that the higher the temperatures, hence the further the object from thermal equilibrium, the more heat energy is lost to the surroundings.

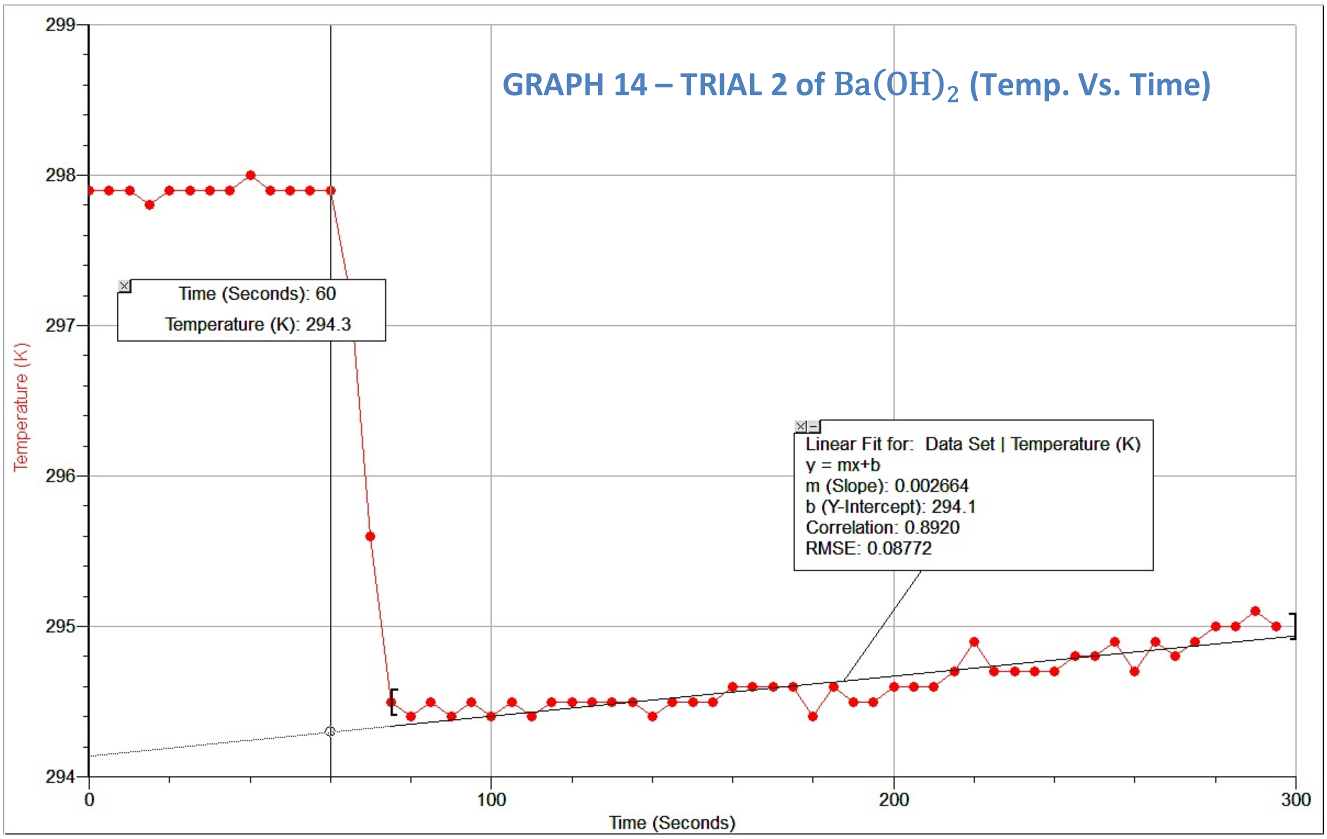
Finally, it is alarming that the percentage uncertainty is 20% for this experiment, but it seems right due to the reasons below and the low change in temperature, which could be negligible or significant. The data obtained is summarized below.

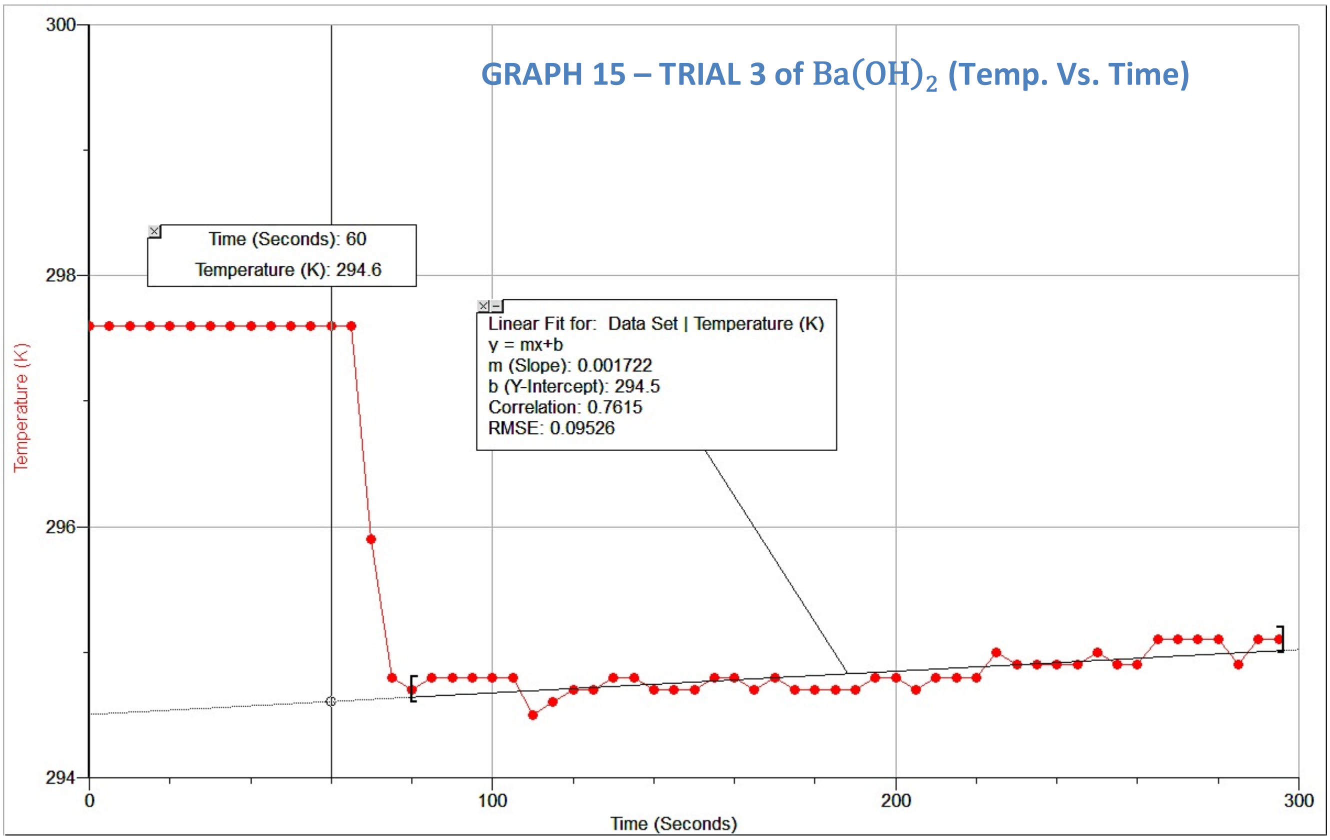
Summary Table –

	Max. Temp. Reached (K)	Room Temperature	Extrapolated Temp. (K)	Absolute Uncertainty	**Change in Temperature**
Trial 1	298.8	298.3	298.8	0.1°C	**0.5**
Trial 2	298.8	297.9	298.9	0.1°C	**1.5**
Trial 3	298.6	298.1	298.6	0.1°C	**0.5**
% unce.	≈ 0.38%	≈ 0.40%	≈ 0.38%		**≈ 20%**

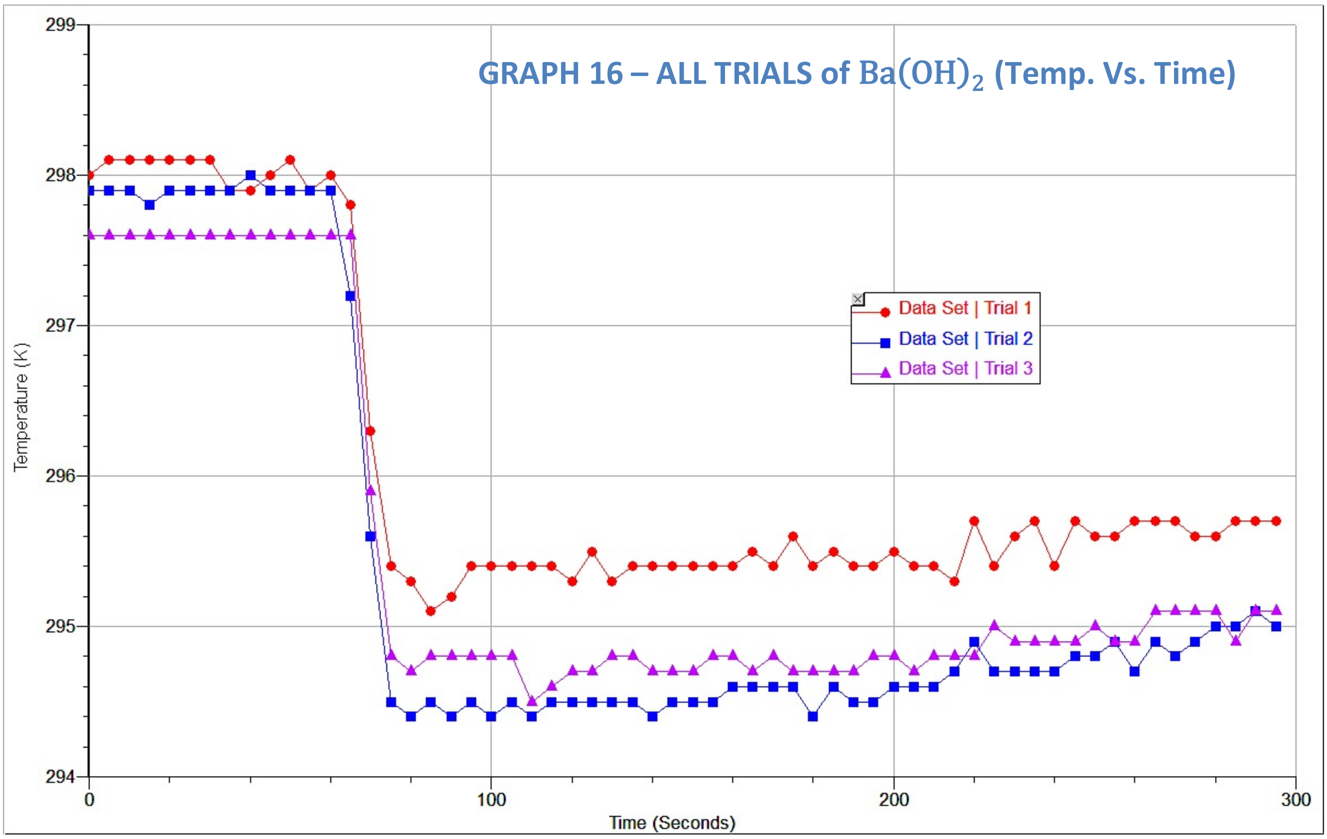
Graphs for $Ba(OH)_2$


GRAPH 13 – TRIAL 1 of $Ba(OH)_2$ (Temp. Vs. Time)


GRAPH 14 – TRIAL 2 of $Ba(OH)_2$ (Temp. Vs. Time)



GRAPH 15 – TRIAL 3 of $Ba(OH)_2$ (Temp. Vs. Time)



GRAPH 16 – ALL TRIALS of $Ba(OH)_2$ (Temp. Vs. Time)

Of the four experiments, this was the only endothermic reaction, where heat was taken in instead of given out. Hence, the temperature of the surrounding reduced rather than increased, unlike the other experiments.

Surprisingly, some of the extrapolated temperatures are higher than the minimum temperature, because normally, a lower extrapolated temperature would be expected. This hints that the lowest temperature recorded could have been an odd point.

Additionally, from the combined graphs, it can be seen that even though trial 1 did not have that much of a higher room temperature, the minimum temperature it could attain was less than the other two trials. This trial, hence could have been privy to random errors, or the mass of solid used in this trial could have been less. From table 1 above, we see that its mass was similar to trial 3, however, did not manage to attain a minimum point similar to trial 3.

Summary Table –

	Min. Temp. Reached (K)	Room Temperature	Extrapolated Temp. (K)	Absolute Uncertainty	**Change in Temperature**
Trial 1	295.2	298.0	295.2	0.1°C	**2.8**
Trial 2	295.4	297.9	294.3	0.1°C	**3.6**
Trial 3	294.5	297.6	294.6	0.1°C	**3.0**
% unce.	≈ 0.45%	≈ 0.40%	≈ 0.46%		≈ **6%**

Table 2: Summary of all important information collected for calculations

		REACTION (With Water)			
		Mass (grams)	**Volume (cm^3)**	**ΔTemp. (K)**	**% uncer. in temp.**
$NaOH$	Trial 1	4.800	60	**18**	≈ 1%
	Trial 2	4.550	60	**16.9**	
	Trial 3	4.324	60	**17**	
KOH	Trial 1	2.380	60	**8.3**	≈ 2.5%
	Trial 2	2.362	60	**8.6**	
	Trial 3	2.573	60	**8.4**	
$Ca(OH)_2$	Trial 1	4.780	60	**0.5**	≈ 20%
	Trial 2	4.893	60	**1.5**	
	Trial 3	4.785	60	**0.5**	
$Ba(OH)_2$	Trial 1	4.834	60	**2.8**	≈ 6%
	Trial 2	4.818	60	**3.6**	
	Trial 3	4.795	60	**3.0**	
% Uncertainty		0.05%	1%		

Annotated Calculations and Uncertainty

STEP 1: Calculating heat energy evolved from the above information in Table 2. Heat Energy in Joules

The heat energy evolved is absorbed by the surrounding. In this case, the liquid medium that the reaction is occurring in is considered the surrounding. Hence the heat evolved can be calculated by the heat absorbed by the water, and hence the increase in temperature of the water as summarized in table 2. The equation for heat energy required to increase the temperature of an object is given by the equation –

$$\Delta H = m \cdot c \cdot \Delta T$$

Where m represents the mass of the object/medium, c is the specific heat capacity of that object – the energy required to raise the temperature of the object by 1K and ΔT, which is the change in temperature.

This process however lends itself to some uncertainty, because not all the heat evolved is absorbed by the medium.

In the experiments above, the volume of water used is constant at 60 cm^3. This volume has to be converted to mass in order to use in the equation.

$$\text{Density}=\frac{\text{Mass}}{\text{Volume}} \gg \text{Mass}=\text{Density} \times \text{Volume}$$

Hence, since the density of water is 1g·cm^{-3} -

$$\text{Mass}= 1\text{g}\cdot\text{cm}^{-3}\times 60\ \text{cm}^3 = 60\text{g} \pm 1\%$$

The percentage uncertainty remains the same because the volume was simply multiplied by a constant. The specific heat capacity of water is 4.1855J·g^{-1}K^{-1}. This hence transforms the equation above to –

$$\Delta H = 60\text{g} \cdot 4.1855\text{J}\cdot\text{g}^{-1}\text{K}^{-1} \cdot \Delta T$$

Now the heat energy evolved can be simply calculated by plugging in the change in temperature values recorded in Table 2. The percentage uncertainty becomes the 1% added to the % of uncertainty in the temperature values (as included in Table 2). Hence for NaOH, the uncertainty of heat evolved becomes 2%.

		REACTION (With Water)		
		ΔTemp. (K)	**Heat Evolved**	**% uncert. in Heat Evolved**
$NaOH$	Trial 1	**18**	4520	≈ 1%
	Trial 2	**16.9**	4244	
	Trial 3	**17**	4269	
KOH	Trial 1	**8.3**	2084	≈ 2.5%
	Trial 2	**8.6**	2160	
	Trial 3	**8.4**	2110	
$Ca(OH)_2$	Trial 1	**0.5**	126	≈ 20%
	Trial 2	**1.5**	377	
	Trial 3	**0.5**	126	
$Ba(OH)_2$	Trial 1	**2.8**	703	≈ 6%
	Trial 2	**3.6**	904	
	Trial 3	**3.0**	753	

STEP 2: Finding the number of moles of reactant that dissolved in the fixed volume of water.

The number of moles have to be calculated from mass, because standard enthalpy values describe heat of a reaction, in moles and kilojoules. The simple equation to convert mass to moles is –

$$n(\text{moles}) = \frac{\text{mass}}{\text{molar mass}}$$

The molar mass however is specific to each compound. The molar mass, is the mass in grams that a mole of the compound would weigh. This is found by adding the atomic mass of the individual elements that make up the compound, in the correct order.

For NaOH –
$Mr(\text{Na}) = \ 22.989769$
$Mr(\text{O}) = \ 15.9994$
$Mr(\text{H}) = \ 1.00794$
Therefore, on adding - $Mr(\text{NaOH}) = \ 39.997$

Other Molar Masses of compounds used –
$Mr(\text{KOH}) = \ 56.1056$
$Mr(\text{Ca(OH)}_2) = \ 74.093$
$Mr(\text{Ba(OH)}_2) = \ 171.34$

Hence, the moles of reactant is found simply by dividing the Mass in Table 2, with the appropriate molar mass of that compound.

The percentage uncertainty of mass is 0.1%, as described in Table 2. The uncertainty remains same for all final mole values, as the mass is simply being divided by a constant. Hence uncertainty of all moles values is $\pm 0.1\%$.

		REACTION (With Water)		
		Mass	**No. of Moles**	**% uncert. in No. of Moles**
$NaOH$	Trial 1	**18**	0.120009	$\approx 0.1\%$
	Trial 2	**16.9**	0.113759	
	Trial 3	**17**	0.108108	
KOH	Trial 1	**8.3**	0.042420	$\approx 0.1\%$
	Trial 2	**8.6**	0.042099	
	Trial 3	**8.4**	0.045859	
$Ca(OH)_2$	Trial 1	**0.5**	0.064514	$\approx 0.1\%$
	Trial 2	**1.5**	0.066039	
	Trial 3	**0.5**	0.064581	
$Ba(OH)_2$	Trial 1	**2.8**	0.028212	$\approx 0.1\%$
	Trial 2	**3.6**	0.028120	
	Trial 3	**3.0**	0.027985	

STEP 3: Calculating final Enthalpy of Solution from the heat energy evolved and moles of reactant dissolved.

The heat energies found in Step 1 illustrates the heat evolved for the number of moles specified in Step 2. However, standard enthalpies describe heat energy evolved for 1 mole of reactant, in the reaction.

Hence, Standard Enthalpy $= \frac{\text{Heat Evolved}}{\text{No. of Moles } (n)}$

By dividing the heat evolved calculated in step 1 with the corresponding moles calculated in step 2, enthalpy is calculated in the form of J·mol^{-1}. This value can further be divided by 1000, to get the values in kJ·mol^{-1}, as normally stated.

		REACTION (With Water)				
		Heat Change (J)	**No. of Moles**	**Enthalpy (J/mol)**	**Average Enthalpy (kJ/mol)**	**% uncert. in Enthalpy**
$NaOH$	Trial 1	**18**	0.120009	37667	**38.155**	≈ 1.1%
	Trial 2	**16.9**	0.113759	37308		
	Trial 3	**17**	0.108108	39490		
KOH	Trial 1	**8.3**	0.042420	49137	**48.812**	≈ 2.6%
	Trial 2	**8.6**	0.042099	51301		
	Trial 3	**8.4**	0.045859	45999		
$Ca(OH)_2$	Trial 1	**0.5**	0.064514	1946	**3.198**	≈ 20.1%
	Trial 2	**1.5**	0.066039	5704		
	Trial 3	**0.5**	0.064581	1944		
$Ba(OH)_2$	Trial 1	**2.8**	0.028212	24923	**27.998**	≈ 6.1%
	Trial 2	**3.6**	0.028120	32151		
	Trial 3	**3.0**	0.027985	26921		

Therefore – Enthalpy of Solution of the four hydroxides are as follows –
(A negative sign is added for exothermic reactions, reactions that released heat)

$$\Delta H_{(s)}(\text{NaOH}) = -\,38.155 \pm 0.42\ \text{kJ}$$
$$\Delta H_{(s)}(\text{KOH}) = -\,48.812 \pm 1.27\ \text{kJ}$$
$$\Delta H_{(s)}(\text{Ca(OH)}_2) = -\,3.198 \pm 0.64\ \text{kJ}$$
$$\Delta H_{(s)}(\text{Ba(OH)}_2) = 27.998 \pm 1.70\ \text{kJ}$$

CONCLUSION AND EVALUATION

ASPECT 1: *CONCLUDING*

From the enthalpies above, we can conclude a few things regarding the intermolecular forces of the compounds and reactivity of the cations that participated in the reaction. Firstly, since the enthalpy of sodium hydroxide was more endothermic than that or potassium hydroxide, (enthalpy of solution of sodium hydroxide was less negative, where the negative sign shows exothermic reaction), it can be concluded that the hypothesis was correct, since it predicted this.

Hence, we see that the stronger bonds in sodium hydroxide required more energy to break. This meant that its enthalpy was more endothermic than potassium hydroxide's, since potassium hydroxide required less energy, and released more energy. The uncertainty was low in these cases, and could not have lead to sodium hydroxide having a more exothermic reaction, since the difference between the enthalpy of the two is far greater than the extents to which the uncertainty would allow the data to migrate to.

The second conclusion we can form is by comparing potassium hydroxide and barium hydroxide. The two hydroxides have similar ionic radii, however the cation in barium hydroxide has a +2 charge, whereas the cation in potassium hydroxide only has a +1 charge. The +2 charge, according to the hypothesis should mean that the enthalpy is more endothermic, because the compound with the +2 cation would require far more energy to break bonds than the +1, because the +2 cation has greater electrostatic forces due to the stronger attraction of negative and positive ions with a +2 charge.

This is supported by the experiments, because barium hydroxide does have a more endothermic reaction than potassium hydroxide. Infact, it has a far greater endothermic reaction than potassium hydroxide, almost differing by 60,000J.

The difference between an increased charge is far greater than the difference of increased ionic radii. From this we can make on of the most important conclusions being, when enthalpy is concerned, the differing charge of two cations in their compounds has a greater affect than any difference in atomic radii. Charge is a dominant factor when compared to atomic radii.

Unfortunately, these conclusions could not be further supported by calcium hydroxide because its experiments into enthalpy of solution had great uncertainties, due to the extremely small change in temperature and the partial solubility. However, by considering literature values found on the Internet, it agrees and reinforces the pattern noticed when considering the above three hydroxides. This might also hint at the trend continuing for all metal hydroxides for the 1st and 2nd group elements.

ASPECT 2&3: *EVALUATING PROCEDURES AND IMPROVEMENTS*

1. The uncertainty was kept at a minimum as an automatic temperature probe and electronic mass balance were used. These apparatus have very low uncertainty. However, the measure of volume had a relatively high uncertainty as it was measured using a graduated cylinder.

2. The uncertainty is very high for calcium hydroxide. This is in part because it is only partially soluble and hence, a large uncertainty is expected and required because the true volume of solid that dissolved to make the solution is not known. Additionally, the large uncertainty is also due to the small change in temperature. The smaller the change in temperature, the larger the uncertainty because the device is less sensitive, relative to the change. The solubility of other compounds could have also affected the final temperature measured and hence enthalpy calculated. This is one of the drawbacks of this experiment.

3. The solubility of some compounds related to others created huge differences in this experiment. Even among potassium hydroxide, sodium hydroxide and barium hydroxide, their slight difference in solubility would have had an affect on the final enthalpy calculated. This was much more significant for calcium hydroxide which had partial solubility and much of the mass added did not dissolve and participate in the reaction. A greater volume of water could have been used to eliminate any difference in solubility. The only problem that arises with this is that, the temperature difference, whether temperature is absorbed or evolved will be much less as there is a larger volume of water to either absorb the heat given off, or evolve the heat required.

4. As seen by the combined graphs, the values of temperature, hence enthalpies are precise. The combined graphs that illustrate data points of all thee trials, have curves that are close to each other. From this, it can be concluded that values were precise. Whether they are accurate or not cannot be inferred from this. However, significant imprecise trials can be seen for Calcium Hydroxide. This is not because the measurements themselves are imprecise, but rather because the scale of the graph taken is small, and every small deviation of 0.1K is represented by a huge spike.

5. The extrapolation also corrected for odd points as seen with Barium Hydroxide, where the extrapolated point was higher than the minimum temperature recorded, because it was an odd point.

6. For the calculation of energy and heat evolved, only the specific heat capacity of the liquid medium: water, was considered. The specific heat capacity and the change in temperature that occurs in the polystyrene cup itself were not investigated. Although polystyrene cup has a high specific heat capacity and is not a heat conductor, hence does not absorb much heat, this would still make a difference in enthalpy. Since not all polystyrene cups are same, and differ based

on thickness and percentage composition, the individual specific heat capacity of the cup would have to be calculated, which is out of scope for this investigation.

7. Experiments evolving heat are compromised due to heat loss to surrounding being major factor. Insulation was used to retain heat, however, several procedures could have been improved to reduce heat loss.
    a. The procedures were mechanical, involving opening the cap of the polystyrene cup and manually adding the content to the container. Often the cap would not be closed properly, in a hurry to measure the maximum temperature achieved.
    b. Heat was also lost due to a hole in the cup to introduce the temperature probe.
    c. Finally, heat measurements could have been made more accurate if a more accurate probe was used. The probe used calculated temperature to one decimal point. One clear advantage of the probe was that it wasn't manual and hence reduced the chances of human error.

8. The heat loss of the experiment due to the reasons above was minimized by the extrapolation of graphs. This analytical tool allowed me to account for some proportion of the heat loss, from when the solid sample was added to the liquid medium.

9. The measurement with the greatest uncertainty is volume, since the most accurate container that could be found was a graduated cylinder. Hence, more accurate method and apparatus could have been used for volume.

10. The strength of cation was observed and a conclusion was formed by considering the enthalpy of solution. Other similar reaction that all four hydroxides undergo could have additionally been used for more data. For example, the hydroxides reaction with hydrochloric acid. By investigating the change in enthalpy due to neutralization. This would have lead to more conclusive evidence for their relative strength in compounds.