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Results

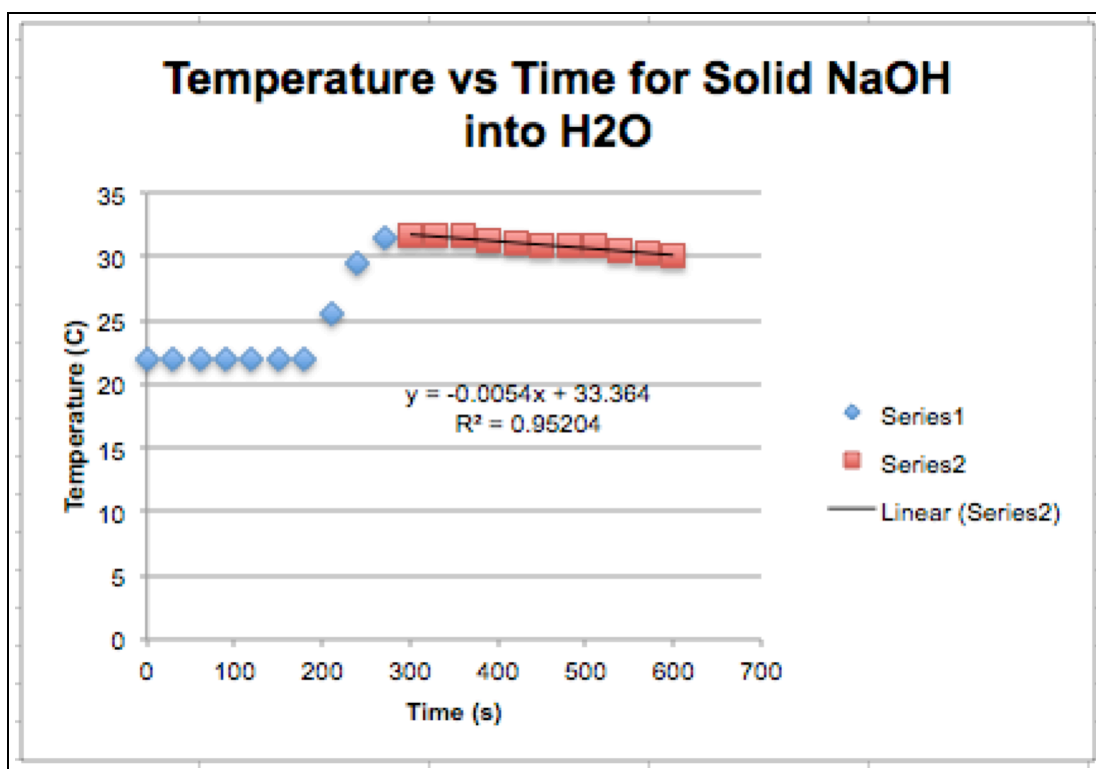


Figure 1. Temperature was measured every 30 seconds before, during, and after the reaction had occurred. We estimate actual heat of the reaction by extrapolating what the heat would have been if the reaction had undergone instantaneously. We can see that there was a relatively slow climb to the max temperature with cooling. This is done by drawing a trendline through the declining temperature (due to cooling) and taking a prediction of the value at the time of addition by evaluating the equation describing the trendline at the time of addition. $34.3522 = -0.0054 \cdot 183 + 33.364$

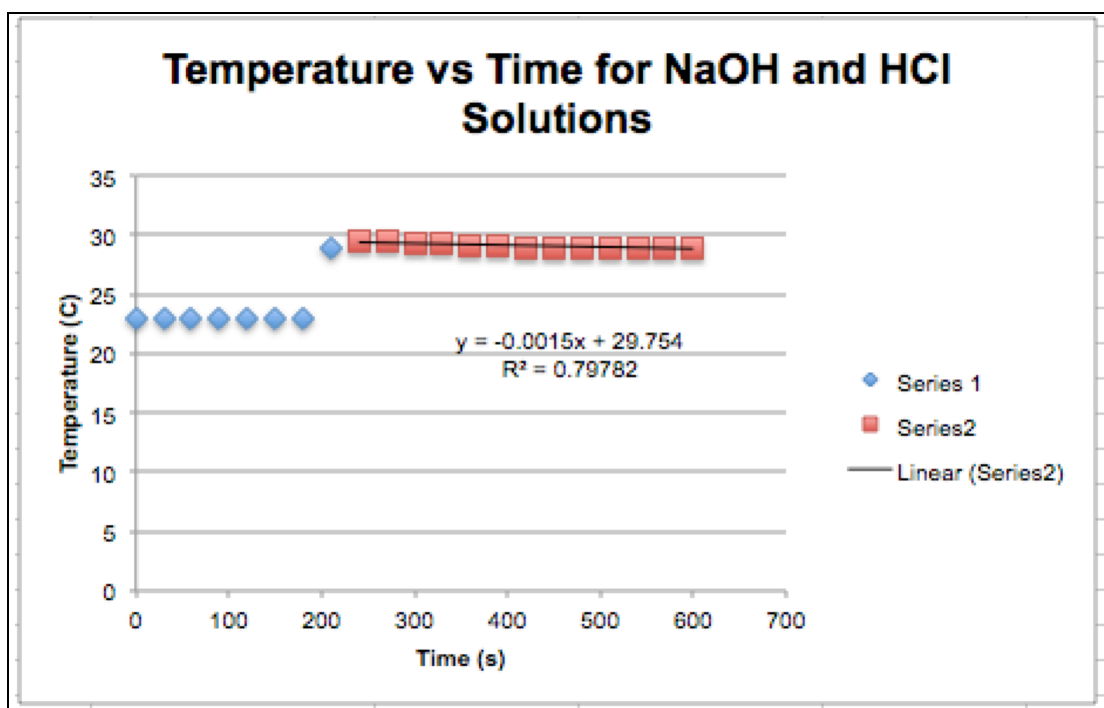


Figure 2. Temperature was measured every 30 seconds before, during, and after the reaction had occurred. We estimate actual heat of the reaction by extrapolating what the heat would have been if the reaction had undergone instantaneously. We can see that there's a very fast jump in temperature, indicating the reaction occurred faster, which means it's likely that our estimated final temperature is much closer to the actual temperature measure during the start of cooling. This is done by drawing a trendline through the declining temperature (due to cooling) and taking a prediction of the value at the time of addition by evaluating the equation describing the trendline at the time of addition. Estimated temperature is 29.4645
 $= -0.0015 \cdot 193 + 29.754$.

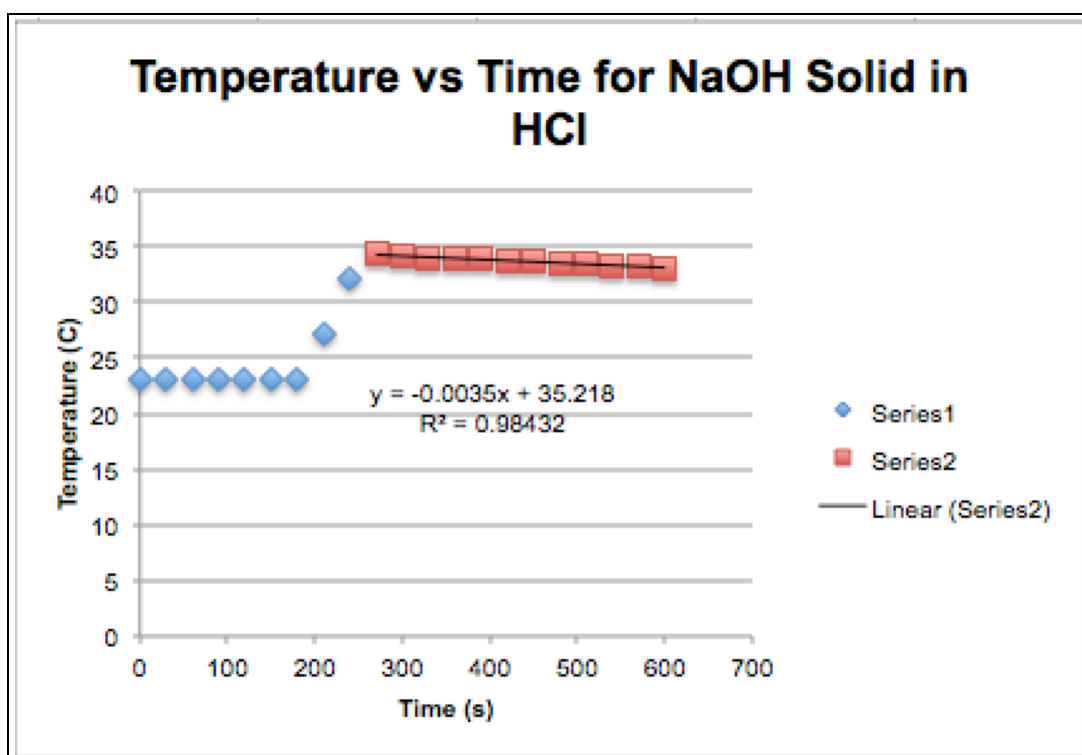


Figure 3. Figure 2. Temperature was measured every 30 seconds before, during, and after the reaction had occurred. We estimate actual heat of the reaction by extrapolating what the heat would have been if the reaction had undergone instantaneously. This is done by drawing a trendline through the declining temperature (due to cooling) and taking a prediction of the value at the time of addition by evaluating the equation describing the trendline at the time of addition. Estimated instantaneous temperature is 34.5425
 $= -0.0035 \times 193 + 35.218$

Table 1. Calculation results with units for heat of reaction and Hess' law analysis

Data Table 3	
Mass from reaction 1 (g)	51.874
Mass from reaction 2 (g)	99.66
Mass from reaction 3 (g)	56.796
Instantaneous estimated temperature for reaction 1 (C)	34.3522
Instantaneous estimated temperature for reaction 2 (C)	29.4645
Instantaneous estimated temperature for reaction 3 (C)	34.5425
Temperature change for reaction 1 (C)	12.3522
Temperature change for reaction 2 (C)	6.4645
Temperature change for reaction 3 (C)	11.5425
Heat gained reaction 1 (J)	2498.956289
Heat gained reaction 2 (J)	2544.795677
Heat gained reaction 3 (J)	2556.714537
Heat lost from calorimeter 1 (J)	221.8002103
Heat lost from calorimeter 2 (J)	116.0787114
Heat lost from calorimeter 3 (J)	207.2609679
Heat of reaction 1 (J)	-2720.756499
Heat of reaction 2 (J)	-2660.874388
Heat of reaction 3 (J)	-2763.975505
Reaction 1 NaOH (moles)	0.0508
Reaction 2 NaOH (moles)	0.05
Reaction 3 NaOH (moles)	0.0495
Ethalpy of reaction 1 (J/mole)	-53558.1988
Ethalpy of reaction 2 (J/mole)	-53217.48776
Ethalpy of reaction 3 (J/mole)	-55837.88899
Percent error 1	670%
Percent error 2	667%
Percent error 3	706%

Discussion

Write a formal conclusion for experiment 6, that includes responses to the bullet points below.

We are applying Hess's Law to our analysis of the chemical reactions because we are conducting two reactions which are each disjoint and fully inclusive subsets of a third reaction. That is, the reactants of two of the reactions combined is the same as the reactants of the third reaction (and thus the products are the same as well). We are to see that the heat of reaction of the first two reactions equal the heat of reaction of the third reaction, as the first reaction had NaOH(s) as a reactant, the second had $\text{Na}^+ + \text{OH}^- + \text{H}^+ + \text{Cl}^-$, and the third had $\text{NaOH} + \text{H}^+ + \text{Cl}^-$ as reactants. Thus we see that the third may be written as a combination of the first two.

The heat of the chemical reaction is negative because the heat (molecular motion) induced by the reaction is given off into the environment, so it is leaving our localized system, thus it is negative. We negate the sign of the sum of the heat of solution and heat of calorimeter to indicate this flow of energy (energy should always be conserved in an closed system, and if a system is open then it indicates giving off energy into its environment by taking the negative of the energy given off). Given constant pressure and volume, the enthalpy of a system is only going to be the internal energy of the system (which is going to vary based on the elements contained in that system and their configuration). Thus, this is relevant as we are declaring to be equivalent to enthalpy because we are assuming that our reaction occurred under constant atmosphere pressure and constant volume.

Our calculated enthalpy of reaction values were very different from the literature values. This likely has to do with a calculation error. However, in general, different values are expected as we were not conducting our experiment at constant atmospheric pressure, for instance, and mainly because our insulator was not nearly industrial enough. This means we are not accounting for all of the heat lost. What was most surprising was the temperature change of reaction 1 (12.3C), which was strangely similar in magnitude to reaction 3 (11.5C), and this is the primary cause for similar changes in heat of reactions, etc. This could be an error due to small perturbations in airflows or malfunctioning/a varying in and out of contact calorimeter, which is a small experimental random error.

The purpose of this experiment was to demonstrate Hess's Law experimentally by showing that two reactions whose sum was identical to a third reaction had equivalent enthalpy.