Experiment 2.

A Tall Cold Drink of Water

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Introduction.

This laboratory session explored enthalpy and the laws associated with it. Enthalpy is a thermochemistry variable that is determined by the change in energy of a chemical reaction in relation to constant pressure.[1] "The standard heat of formation of a substance is defined as the change in enthalpy when one mole of the substance is formed from its elements in their standard states."[2] Hess' Law specifies that not every chemical reaction is simple, and that for some reactions, multiple processes need to be carried out to determine the enthalpy change of the reaction. Specific heat capacity, which varies with temperature, is a measurable quantity that determines the thermal energy of a reaction. The amount of thermal energy that is required to heat one gram of a substance by one degree is known as the specific heat of a substance. To calculate the heat capacity of a substance, the mass of the substance is multiplied by its specific heat capacity.[3] Calorimetry is described as a method of measuring the heat of formation within a chemical reaction with the use of a calorimeter. A calorimeter consists of two vessels, inner and outer. The outer vessel is used to isolate the inner vessel to prevent excess heat loss within a reaction.[4] Water was used for the experiment in this laboratory session and it was predicted that the water would drop 5°C in five minutes with the addition of ammonium chloride, through calorimetry. Water was placed in the outer vessel of the calorimeter, and an aluminum can was filled with water as the inner vessel. To change the temperature of the water, ammonium chloride was placed in the outer vessel in attempt to lower the temperature of the water inside the can of the calorimeter. To determine the energy required to change the temperature of the water, the following formula was used: $Q = mc\Delta t$, Q being the energy added to a system, m being the mass, c the specific heat capacity, Δt being the change in temperature. This experiment used water and ammonium chloride to test, through multiple trials, how the two substances reacted together to change the initial measured temperature of the water.

Procedure.

Using the provided guidelines from the contract (uOttawa Gee Gee's Sport and Health Drinks, Dr. R. Venkateswaran), a volume of 100 grams of water was measured and placed in a calorimeter and an aluminum can filled with water was placed in the center of the calorimeter. The initial temperature of the water was measured and was found to be around room temperature. It was collectively agreed upon that an amount of 13.28 grams of ammonium chloride was used as the initial volume of salt for the experiment. The salt was placed into the water between the calorimeter and the aluminum can. After five minutes, the temperature of the water inside the can was re-measured and recorded. This experiment was repeated for four trials using increasing amounts of ammonium chloride, in hopes that the temperature of the water would drop five degrees exactly in five minutes.

Results.

The initial temperature measurements of the water varied between 21.8°C and 28.2°C. Once the ammonium chloride was placed into the water, it sunk and dissolved quickly. With each trial, the temperature of the solution had a dramatic decrease at the beginning of 2 minutes within the waiting period of 5 minutes. In the following 3 minutes, the temperature of the solution remained almost constant for each trial. The change in the initial and final temperature measurements of the water varied between -3.4°C and -5.6°C. The desired change of -5°C was achieved during the last trial.

Table 1: Data recorded during the experiment to determine the enthalpy of the reaction.

Trial No.	Initial Temp. (°C)	Final Temp. (°C)	ΔT (°C)	Mass of NH_4Cl (g)
1	23.9	20.5	-3.4	13.28
2	28.2	23.8	-4.3	15.28
3	21.8	17.5	-4.3	18.28
4	22.8	17.2	-5.6	22.28

Calculations.

Firstly, the Q values of the water outside the can, the water inside the can, and the Q value of the can itself were calculated.

$Q_{(\text{Water inside can})}=mc\Delta T$

$= 100.\text{g} \times 4.18\ \text{J/g*K} \times 5\text{K}$

$=2090\text{J}$

$Q_{(\text{Water outside can})}=mc\Delta T$

$=100.\text{g} \times 4.18\ \text{J/g*K} \times 5\text{K}$

$=2090\text{J}$

$Q_{(\text{Can})}=mc\Delta T$

$=8.63\text{g} \times 0.902\ \text{J/g*K} \times 5\text{K}$

$=38.9\text{J}$

$Q_{(\text{Total})}=2090\text{J}+2090\text{J}+38.9\text{J}$

$=4218.9\text{J}$

The ΔH value for the reaction, $NH_4Cl_{(s)} \rightarrow NH_4^+{}_{(aq)} + Cl^-{}_{(aq)}$, was calculated in the prelab to have a value of ΔH=17kJ. Using this value and the $\Delta H = q_s/n_{salt}$, the number of moles of salt required to satisfy the reaction were calculated.

17kJ x 1000J/kJ= 17000J

$\Delta H = q_s/n_{salt}$

$N_{salt} = q_s/\Delta H$

=4218.9J/17000J

=0.248 mol NH_4Cl

Lastly, the the number of moles were converted into the needed mass in grams.

Molar mass of NH_4Cl= (14g/mol N) + 4(1g/mol H) + (35.5g/mol Cl)

MM=53.5g/mol NH_4Cl

Needed mass:

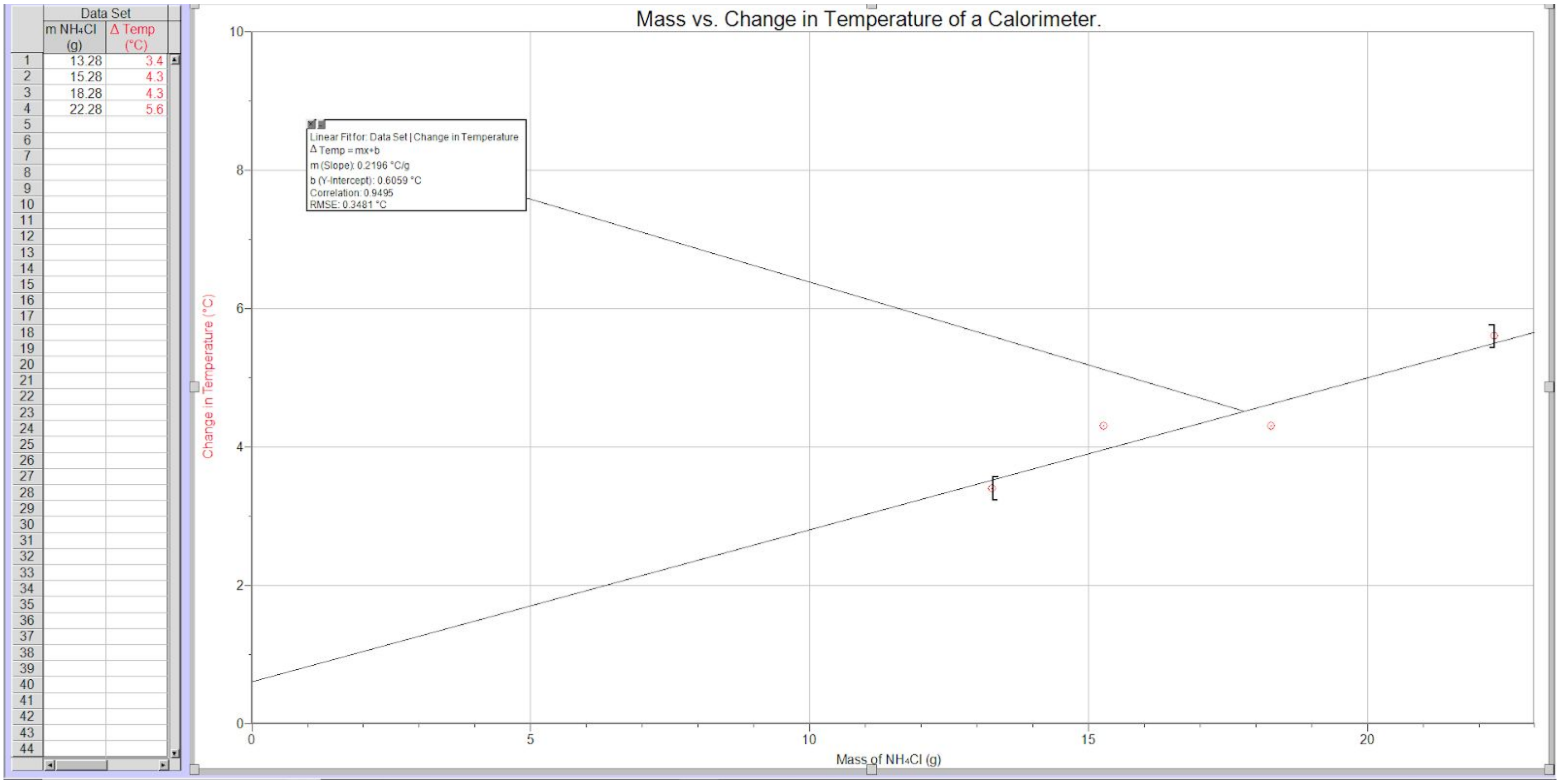
0.248mol NH_4Cl x 53.5g/mol NH_4Cl

=13.3g NH_4Cl

Discussion.

Collectively, NH_4Cl was the chemical that was decided upon to be most suitable for the experiment because the price per 500g was the cheaper and safer to work with when compared to the other chemical, NH_4NO_3. Also, due to the regulations of health and safety imposed by Health

Canada, NH_4Cl was the safer choice. As described in, *Type E High Explosives,* NH_4NO_3 is classified as highly explosive and therefore, not suitable at all to be put into consumers drinks.[5]



	Data Set	
	m NH_4Cl (g)	Δ Temp (°C)
1	13.28	3.4
2	15.28	4.3
3	18.28	4.3
4	22.28	5.6

Linear Fit for: Data Set | Change in Temperature
Δ Temp = mx+b
m (Slope): 0.2196 °C/g
b (Y-Intercept): 0.6059 °C
Correlation: 0.9495
RMSE: 0.3481 °C

Figure 1: Mass (g) vs Change in Temperature of a Calorimeter (°C)

According to the line of best fit generated from the graph of the collected data:

$y=mx+b$

$y=0.2196\ °C/g\ (x)\ +0.6059\ °C$

$5=0.2196x+0.6059$

$x=20.01g$ (This is the actual amount of NH_4Cl that would be needed to get a temperature change of 5°C)

With this value of 20.01g NH_4Cl in mind, the percent error would be:

$$P.E. = \left(\frac{Experimental\ value - Theoretical\ value}{Theoretical\ value}\right) \times 100\%$$

$$= \left(\frac{20.01g - 13.3g}{13.3g}\right) \times 100\%$$

$$= 50.5\%$$

As observed in **figure 1**, the line does not pass through the origin, which could have been caused by a systematic error occurring with the thermometer. Although the result is not accurate, every dot lies evenly close to the line with the correlation number being 0.9495, which is very close to 1. This means that the result was precise, so the mass of ammonium chloride that was added was proportional to the temperature change of the water.

In this experiment, there was a large percentage error, because every time ammonium chloride was placed into the water; there was always a small mass of the salt that remained stuck inside the beaker and unable to enter the solution. As more ammonium chloride was added, the mass of ammonium chloride grew gradually close to the solubility of the compound water and, as a result, the reaction speed would become slower. At the end of the five minutes given for the reaction time, it was observed that a small amount of ammonium chloride would still remain at the bottom of the container. Another possible source of error in the experiment could have been caused by the fact that the second trial had an unusually high initial temperature, exceeding that of room temperature, which was most likely caused by using the hot water faucet to measure the 100g of water.

Conclusion.

In this experiment, it was concluded that the mass of ammonium chloride was proportional to the temperature change inside a calorimeter. It was hypothesized that a minimum of 13.28g of reactant was needed to achieve a temperature change of -5℃ in 5 minutes, however, according to **figure 1**, at least 20.01g of ammonium chloride was needed to decrease the temperature of the water by 5℃ in 5 minutes. The result of 20.01g of ammonium chloride was determined to be most reliable because it was tested in a real environment with many diverse factors, such as, a higher initial temperature of the water, that could have affected the speed and effectiveness of the endothermic reaction.

References.

[1] Silberberg, 2e. "Chemistry: The Molecular Nature of Matter and Change." ch 5. p. 205

[2] Unknown Author. "...But a Hot Temper Leaps O'er a Cold Degree." Experiment 2. p. 28.

[3] Unknown Author. "...But a Hot Temper Leaps O'er a Cold Degree." Experiment 2. p. 29.

[4] vlab.amrita.edu. (2012). *Calorimetry -Water equivalent Calorimetry.* Retrieved from http://vlab.amrita.edu/?sub=2&brch=190&sim=1352&cnt=1

[5] Natural Resources Canada. (2015). Type E High Explosives. Retrieved from https://www.nrcan.gc.ca/explosives/resources/guidelines/16423#int-int

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Exp. 2

Trial 1:

initial temp of water - 23.9°C
final temp of water - 20.5°C $\Delta t = 3.4°C$
mass of NH_4Cl used - 13.28 g

Trial 2:

t_1 of H_2O - ~~27.2°C~~ 28.0°C
t_2 of H_2O - 23.8°C $\Delta t = 4.2°C$
mass of NH_4Cl - 15.28 g

Trail 3:

t_1 of H_2O - 21.8°C
t_2 of H_2O - 17.5°C $\Delta t = 4.3°C$
mass of NH_4Cl - 18.28 g

Trial 4:

t_1 of H_2O - 22.8°C
t_2 of H_2O - 17.2°C $\Delta t = 5.6°C$
mass of NH_4Cl - 22.28 g