Introduction

Calorimetry is a tool to measure the change in heat energy of a reaction (Landau). It is a very commonly used technique in thermodynamics, since most chemical reactions consume or release some amount of thermal energy (Landau). Using these techniques, reactions can be utilized to create many different kinds of products that require temperature changes.

The purpose of the experiment was to use the principles of thermodynamics to design a container that, when shaken, would cool an energy drink by 5°C in no more than 5 minutes. Heat energy is sapped from the drink using an endothermic reaction— the dissociation of NH_4Cl into its separate ionic components.

The can itself is designed to hold the drink in the inner compartment, and the components of the endothermic reaction in the outer compartment. When activated, the mass of NH_4Cl will be released into the water and take heat energy from the drink to break its ionic bonds and dissociate into its separate ions, resulting in a cooling effect on the drink. In this experiment, the independent variable was the mass of NH_4Cl, the dependent variable was the temperature, and the reaction time was kept constant.

The relationship between heat energy, mass, specific heat, and the change in temperature was exploited to determine the first mass used for the experiment. These variables are all gathered in a single equation $q = mc\Delta T$, where q is the heat energy in joules, m is the mass, c is the specific heat, and ΔT is the change in temperature. The mass that needed to be used in the equation was the total amount of water, because if the heat energy needed to change the temperature of the water by -5°C is known, then the corresponding negative value is the energy consumed by the dissolution of NH_4Cl. Then, once the amount of energy consumed is known, it can be used in the equation $n = \frac{q}{\Delta H_{sol}}$ to determine the number of moles. Once the number of moles of NH_4Cl is known, it can be converted to grams using the molar mass.

After testing, the results can be graphed on a linear function to extrapolate the actual amount of NH_4Cl needed to complete the -5°C temperature change within 5 minutes. By using the extrapolated value, less funds need to be spent on extra material and thus the design can be more cost efficient. To increase cost efficiency, NH_4Cl is used as an alternative to NH_4NO_3 because it is the cheaper of the two and the solubility is lower, meaning less is required to saturate the water.

$$q = mc\Delta T$$

$$n = \frac{m}{M}$$

$$n = \frac{q}{\Delta H_{sol}}$$

Materials

- Calorimeter
- Ammonium chloride (NH_4Cl)
- Temperature probe

Procedure

1. Fill a drinking can with 100mL of water (do not put the can in the calorimeter yet)
2. Pour 70mL of water into the calorimeter
3. Add the predetermined mass of NH_4Cl into the calorimeter with the water
4. Place the drinking can in the calorimeter and cover it with the lid
5. Insert the temperature probe and record the initial temperature
6. Shake gently for 5 minutes
7. Record the final temperature
8. Repeat experiment twice with two different masses (one higher and one lower mass)

Data, observations, and results

	Mass (g)	Initial Temperature (T_1)	Final Temperature (T_2)
Trial 1	11.18g	23.4°C	20.0°C
Trial 2	11.18g	23.4°C	19.0°C
Average of Trial 1&2	11.18g	23.4°C	19.5°C
Trial 3	16.00g	23.4°C	17.5°C
Trial 4	7.00g	23.4°C	20.7°C

Calculations

$q = (170g)(-5^{o}C)(4.18J/mol)$

$q = \ -3553J$

$n = \frac{3.553kJ}{17kJ/mol}$

$n = 0.209mol$

$0.209mol = \frac{m}{(14.007+4.032+35.45)}$

$m = 11.179g$

$q = (170g)(-3.9°C)(4.18J/mol)$

$q = -2771.34J$

$$\frac{|(-3553)-(-2771.34)|}{-3553} \times 100 = \%error$$

$\%error = |-22\%|$

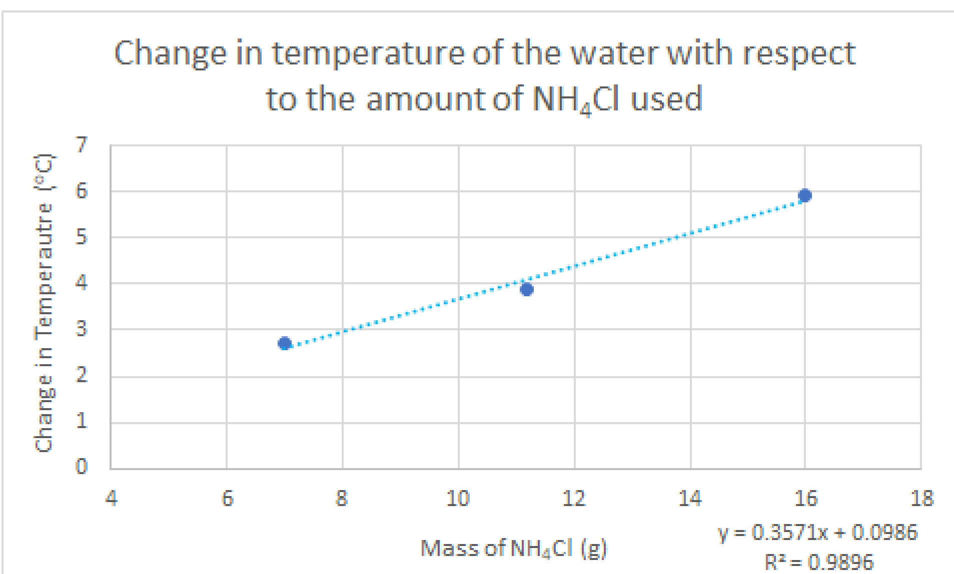
$\%error = 22\%$

<u>Percent Error of Trials</u>

	Percent Error
Average of Trial 1 & 2	22.0%
Trial 3	17.5%
Trial 4	13.7%
Average of Trials	17.7%

<u>Discussion</u>

The results showed that when more solute is put into the water, more heat energy is sapped from the drink to break the ionic bonds. When graphed, the data presented a linear relationship between the mass of ammonium chloride added to the calorimeter and the change in temperature of the water in the drinking can. The water inside the drinking can cools based on the mass of salt added to the calorimeter. With this data, the theoretical mass of salt needed to cool the water in 5 minutes can be extrapolated directly from the graph or calculated using the equation of the trend line.



The data collected varies from its theoretical value due to the inconsistent movement of the calorimeter as well as the structure of the calorimeter. The experiment required the calorimeter to be shaken by hand and, given the nature of human inconsistency, it is likely that the results were affected by the irregular movement of $NH_4Cl_{(aq)}$ over the five minutes. The results may also vary due to the structure of the calorimeter itself. In theory, the calorimeter is a closed system yet the one used in this experiment had flaws such as a hole in its cover and an unsecured lid, both of which allowed for heat to be released from the system. These flaws justify the 17.7% average percentage error calculated from our results.
The percent error of the experiment can be compensated for so that the result is still a -5°C drop in the temperature of the water. Using the trendline of the collected data, 13.73g of ammonium chloride would be needed to produce a -5°C change in temperature with the setup in this experiment.

Conclusion

The results of our experiment had a large percent error because of the use of a partially isolated system that allowed heat energy to escape. The theoretically closed system proved to be much more flawed in practice than it is in the theory and this was proven by our first two trials that had a percent error of 22.0%, where we used the same mass (11.18g), same initial temperature (23.4ºC) and yet had different final temperatures (20.0ºC & 19.0ºC).

References

Landau, "Expanding the role of reaction calorimetry", Thermochimica Acta, 289(2), 101–126, 1996

"What in the World ISN'T Chemistry", General Chemistry Laboratory Manual, Dr. Rashmi Venkateswaran, 2019.