TOTAL REPORT GRADE 31.75/40

Experiment 2: Thermodynamics

Written by:

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TA: Leah McMunn

Date: 28.09.2018

Introduction:

Introduction 2/8

- good explanation of cost effectiveness, but also need to explain general theory surrounding the concepts involved in the experiment
- things like endothermic vs endothermic, hydration energy vs lattice energy, hess' law and energy conservation
- can also have included some of the explanations behind the formulas used like $q=mc\Delta T$ $n=\Delta H/q$ etc

The purpose of this experiment is to find an efficient way to freeze 100 cm^3 of drinkable water in under 5 minutes using ammonium chloride ($NH_4Cl_{(s)}$).
Ammonium chloride ($NH_4Cl_{(s)}$) and ammonium nitrate ($NH_4NO_{3(s)}$) are provided in a theoretical part of the pre-lab. Both of the compounds can be used to cool down the drink. Nonetheless, the process has to be economically optimized.
The cost of lowering the temperature of 100 cm^3 water by 5°C:

$q = mc\Delta T = 100\ g \times 4.18\ Jg^{-1}K^{-1} \times 5^{o}K = 2090\ J = 2\ kJ$ (endothermic)

For ammonium chloride ($NH_4Cl_{(s)}$):
moles (NH_4Cl) = 2 kJ / 17 kJ/mol = 0.1 mol;
m (NH_4Cl) = nm = 53.49 g/mol x 0.1 mol = 5.3 g;
Price (NH_4Cl) = $/g x g = 0.1248 $/g x 5.3 g = 0.66 $

For ammonium nitrate ($NH_4NO_{3(s)}$):
moles (NH_4NO_3) = 2 kJ / 25 kJ/mol = 0.08 mol;
m (NH_4Cl) = nm = 80.04 g/mol x 0.08 mol = 6.4 g;
Price (NH_4Cl) = $/g x g = 0.1292 $/g x 6.4 g = 0.83 $

0.66 $ < 0.83$

According to the calculations listed above, it is more financially efficient to use ammonium chloride ($NH_4Cl_{(s)}$), than ammonium nitrate ($NH_4NO_{3(s)}$). Furthermore, ammonium nitrate ($NH_4NO_{3(s)}$) is a highly volatile compound that can explode if shacked. Summarizing, during this experiment only ammonium chloride will be tested.

In order to make the cooling process safe, the can of the drink will have a second layer filled with water and an active ionic compound. ✓ Thus, no unwanted chemicals will be consumed by the consumer. In order to simulate this reaction, the drink will be located in a tin can submerged into the styrofoam calorimeter with a distilled water and the ionic compound. The styrofoam calorimeter can be further isolated to reduce the leakage of heat from the outside. However, in a real life situation, the outer layer of the can will not be isolated, so even use of the styrofoam calorimeter is not ideal. For the measurement purposes thermometer will be placed inside the drink.

According to the calculations outlined above, it is required 5.3 g of pure ammonium chloride (NH_4Cl) to lower the temperature of 100 cm^3 of distilled water by 5°C. However, this does not take in account that an ionic compound will be added directly to the drink, but to the liquid that surrounds submerged container with the drink. Thus, the volume of the surrounding liquid and the mass of the ionic compound have to be found experimentally. ✓

Materials:

1. Ammonium chloride ($NH_4Cl_{(s)}$)
2. Tap water ($H_2O_{(l)}$)

Apparatus:

- Tin can
- Styrofoam calorimeter
- LabQuest 2
- Thermometer attachment for the LabQuest 2
- Balance
- Graduated cylinder

- Spatula
- Weighing paper

Procedure: Procedure 3.5/4

1. Measure 100 mL of a tap water ($H_2O_{(l)}$) and pure it into the styrofoam cup using the graduated cylinder.
2. Measure 100 mL of a tap water ($H_2O_{(l)}$) and pure it into the tin can using the graduated cylinder. the can is technically aluminum
3. Place tin can into the styrofoam calorimeter.
4. Place a weighing paper on the balance and set the balance to zero.
5. Take ammonium chloride ($NH_4Cl_{(s)}$) using spatula from the container and place it on the weighing paper.
6. Measure 5.3 g of ammonium chloride ($NH_4Cl_{(s)}$) placing it on the weighing paper.
7. Carefully pour ammonium chloride ($NH_4Cl_{(s)}$) into the styrofoam calorimeter without contaminating the tap water ($H_2O_{(l)}$) in the tin can.
8. Close the styrofoam calorimeter.
9. Insert the connected to labquest 2 thermometer attachment into the styrofoam calorimeter through the whole in the lid.
10. Gently stir the styrofoam calorimeter to make sure that all of the ammonium chloride ($NH_4Cl_{(s)}$) is dissolved.
11. Stop stirring the styrofoam calorimeter after the five (5) minutes from the start of the dissolution reaction.
12. Record the extremums of the temperature graph in the five (5) minutes interval displayed in the LabQuest 2.
13. Dispose the used material according to the TA's instructions and clean the apparatus. discard into aqueous waste
14. Repeat steps 1-13 changing the volume of the tap water ($H_2O_{(l)}$) inside the styrofoam calorimeter and the mass of the ammonium chloride ($NH_4Cl_{(s)}$) according to the temperature change. which masses did you use?

Safety Precautions:

- Protection glasses and a lab coat have to be weared during the lab.
- Very cold objects and compounds should not be touched without the proper equipment.
- Toes covering shoes have to be weared during the lab.
- Long hair should be tied during the lab.
- After completing the practical part of the lab, the used material should be disposed according to the orders of the lab personnel.
- Ammonium chloride (NH_4Cl) may release some chlorine ions during the experiment. It is important to do not inhale or smell any compounds during this lab.

Observations:

- Tap water ($H_2O_{(l)}$) inside the styrofoam calorimeter became colder after the addition of the ammonium chloride ($NH_4Cl_{(s)}$).
- During the first trial, styrofoam calorimeter was not stirred, so the ammonium chloride ($NH_4Cl_{(s)}$) did not fully reacted with the tap water ($H_2O_{(l)}$). Changes to the experimental design were introduced to eliminate this mistake.
- During the last two trials, the masses were similar, but the temperature difference was large.

Observations 1.5/2

- make sure that your observations are also written in your raw data
  - they occur in the lab, therefore they should be in your raw data

- The required amount of ammonium chloride ($NH_4Cl_{(s)}$) occupied most of the surface of the weighing paper. This resulted in a partial spread of ammonium chloride ($NH_4Cl_{(s)}$) on the balance's dish.

Tables 4/4

✓

Volume of the surrounding liquid (H_2O) , mL	Mass of the ammonium chloride (NH_4Cl), g	Resulted temperature of the drink, °C (Tf - Ti = ΔT)
100	10.6	23.1°C - 21.2°C = -1.9°C
100	10.6	18.4°C - 22.4°C = -4.0°C
100	13.3	16.4°C - 22.4°C = -6.0°C
100	11.7	16.9°C - 21.1°C = -4.2°C
100	12.3	18.0°C - 23.4°C = -5.4°C

Calculations:

The initial mass of the ammonium chloride was chosen according to the calculations outlined in the hypothesis section. However, these calculations considered the volume to be 100 cm^3, while on practice, it was 200 cm^3, as it is the total volume of the tap water ($H_2O_{(l)}$) in a system.

5.3g / xg = 100cm^3 / 200cm^3

✓ xg = 10.6g

First trial contained an experimental mistake described in details in the discussion section. Measurements of the second trial showed a temperature change in 4.0°C. The following calculations were conducted to find a proper mass of the ammonium chloride ($NH_4Cl_{(s)}$).

10.6g/xg = -4.0°C/-5.0°C

Calculations 1.75/2

✓ xg = 13.3g

you should have calculated this using the equation of your line in your graph

Nonetheless, an addition of 13.3g of the ammonium chloride ($NH_4Cl_{(s)}$) resulted in the temperature change of the system by 6.0°C. Theoretically, the correct temperature of the drink is the average temperature between the second and third trials. The next mass was found through the calculations of the average mass from the second and third trials.

(10.6g + 13.3g) / 2 = 12.0 g

During the fourth trial, solute was stirred in a solution before the thermometer was inserted. This was done to ensure the dissolution of the ammonium chloride ($NH_4Cl_{(s)}$). However, the dissolution reaction was too fast, so the initial temperature was not determined correctly. Fifth trial excluded this mistake and contained a slightly increased amount of ammonium chloride ($NH_4Cl_{(s)}$), achieving the desired temperature change. Comparing the practical and theoretical results, the percentage error can be calculated.

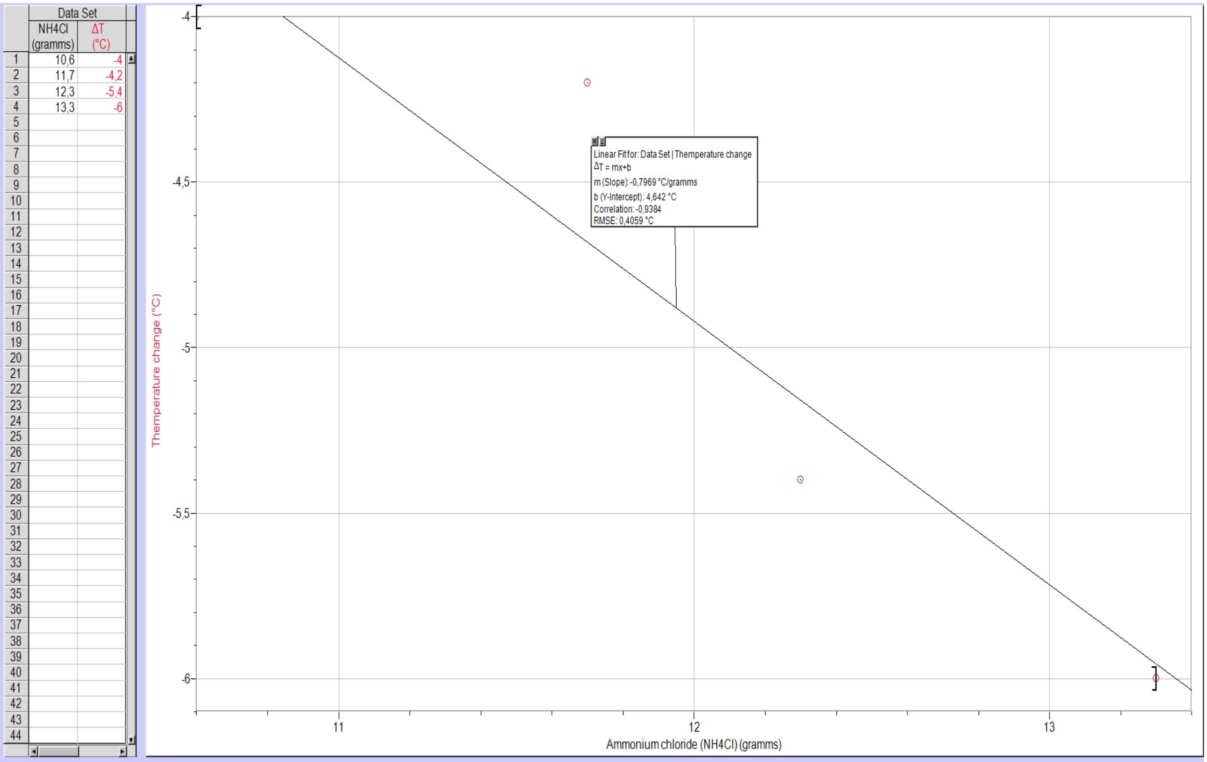
%error = ((|Practical - Theoretical|) / Theoretical) x 100%

%error = ((|12.3g - 10.6g|) / 10.6g) x 100%

%error = 16.0%

Graph: (Data retrieved from the first trial was omitted, as it contains an experimental error.)

Graphs 4/4 + 2



	Data Set	
	NH4Cl (gramms)	ΔT (°C)
1	10,6	-4
2	11,7	-4,2
3	12,3	-5,4
4	13,3	-6
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Linear Fit for: Data Set | Themperature change
ΔT = mx+b
m (Slope): -0,7969 °C/gramms
b (Y-Intercept): 4,642 °C
Correlation: -0,9384
RMSE: 0,4059 °C

~~**Observations:**~~ Discussion

In an overall, the hypothesis appeared to be correct. The drink became colder after the addition of the ammonium chloride ($NH_4Cl_{(s)}$) into the surrounding tap water ($H_2O_{(l)}$) inside the styrofoam calorimeter. The choice of the ammonium chloride ($NH_4Cl_{(s)}$) over the ammonium nitrate ($NH_4NO_{3(s)}$) was proved to be correct, due to the volatile and explosive nature of the ammonium nitrate ($NH_4NO_{3(s)}$). Furthermore, ammonium chloride ($NH_4Cl_{(s)}$) is cheaper than the ammonium nitrate ($NH_4NO_{3(s)}$). The design of the experiment was theoretically accurate. Practically, not all particles of the ammonium chloride ($NH_4Cl_{(s)}$) dissolved. During the first trial, the solvent was not stirred after the addition of the solute. This mistake was taken in account for the second trial. The difference between the first two trials is more than 50% (100% - (-1.9°C / -4.0°C) x 100% = 57.5%). As the styrofoam calorimeter has to be closed after the addition of the solute, it is almost impossible to use a stirring rod or any similar instruments to guarantee the complete dissolution of the ionic compound. The results retrieved during the experiment are distant from the practical application, as the tin can will be used instead of the styrofoam calorimeter. That is, more ammonium chloride ($NH_4Cl_{(s)}$) should be added, as less effective insulation will be in place.

In the dissolution reaction studied during this experiment, solute is not the only controllable variable. It is possible to decrease the amount of the solvent, which will eventually decrease the volume of the whole system. On the other hand, decreasing the volume of the solvent will increase an amount of non-dissolved particles of the ammonium chloride ($NH_4Cl_{(s)}$). Thus, decrease of the volume of the tap water ($H_2O_{(l)}$) will make the dissolution of the same amounts of solute more complicated, which is one of the main issues in this experiment. Practically, an outer layer of the tin can may become too cold for an individual to hold it, if the concentration of the ionic compound momentarily dissolved to the surrounding tap water ($H_2O_{(l)}$) will reach too high levels. Lastly, the solubility of the ammonium chloride ($NH_4Cl_{(s)}$) is 37.2g / 100cm^3 of distilled water ($H_2O_{(l)}$).[1] Therefore, there is a limit to the minimal amount of a solvent present. During this experiment, tap water was

Discussion 6/8
- what was your theoretical mass of salt based on your calculations? what mass actually would have given you a 5°C change based on your graph?
- compare the two
- great error and potential solution discussion

used instead of distilled water. For that reason, the practical solubility was lower compared to the theoretical one.

The inside volume of the styrofoam calorimeter was larger than a filled tin can, ammonium chloride ($NH_4Cl_{(s)}$) and a surrounding tap water ($H_2O_{(l)}$). This allowed a large amount of air to take a part in the studied enthalpy change system, heating up the tin can and the surrounding water. Additionally, the tin can was not fully covered with the tap water around, lowering the heat exchange rate between them.

Multiple improvements can be introduced to improve the experimental design for this experiment. First of all, the water used should not be distilled, but tap. This will ensure that the practical solubility rate of the ammonium chloride ($NH_4Cl_{(s)}$) equals to the theoretical solubility rate. Secondly, the second whole for the stirring rod should be made in the styrofoam calorimeter lid, so there will be a possibility to properly stir the solvent. Styrofoam calorimeter should be replaced with a tin can, as it will be closer to the real life situation. The tin can with the drink should not have a concave inside spherical bottom, as parts of the ammonium chloride ($NH_4Cl_{(s)}$) may accumulate there, without being dissolved. The volume of the styrofoam calorimeter should be reduced, so the tap water ($H_2O_{(l)}$) will reach the edges of the tin can. This is needed to increase the heat conductivity between the two liquids. The volume of the solvent should be decreased for the future experiments, as this will decrease the total volume of the system. Lastly, the thermometer itself has to be minimized, as it becomes a part of the enthalpy change system.

Conclusion:

Conclusion 2/2

The Born-Haber cycle[2] for the ammonium chloride ($NH_4Cl_{(s)}$) outlined in the pre-lab appeared to be correct. The percentage error for the most successful attempt was found to be 16%. The amount of ammonium chloride ($NH_4Cl_{(s)}$) used during the most successful trial is equal to 12.3g. Summarizing, it is possible to create the self cooling drink using the proper ionic compounds.

References:

1. Ammonium chloride. (n.d.). Retrieved from https://pubchem.ncbi.nlm.nih.gov/compound/ammonium_chloride
2. Libretexts. (2017, February 15). Lattice Energy: The Born-Haber cycle. Retrieved from https://chem.libretexts.org/Textbook_Maps/Inorganic_Chemistry/Supplemental_Modules_(Inorganic_Chemistry)/Crystal_Lattices/Thermodynamics_of_Lattices/Lattice_Energy:_The_Born-Haber_cycle

Creative Aspect:

1. Creative name for product.
   a. GJuice.
2. Creative label design for product.
   a. GGs choose GJuice.
3. Creative design for can.

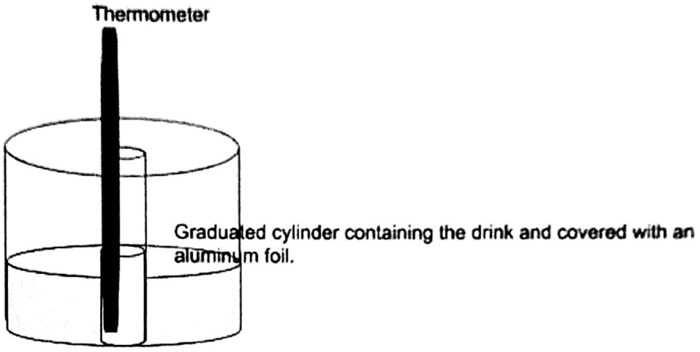

NEW ENERGY DRINK
G
JUICE
FROSTY WHEN YOU ARE THURSTY
GG
uOttawa

Raw Data:

Assessment Criteria for Planning A Tall Cold Drink of Water

TA Name:	Leah McManus	**Names of Students in Group:**	**a.** Alexey
			b. Xinzhu Li
		Date: 25.09.2018	
Criteria:	**Marks Possible**	**Assessment**	
		Self	**TA**
1. Identify the problem and state it clearly in a way that can be tested.	1	1	1
2. Use proper apparatus, techniques and safety precautions.	1	1	1
3. Materials are easily available.	1	1	1
4. Plan to vary only one independent variable at a time.	1	1	1
5. Controls on other variables are clearly stated.	1	1	1
6. Measurement errors are minimized by appropriate procedures or apparatus.	1	1	1
7. The methods are clear enough to be followed by other students.	1	1	1
8. No invalid assumptions are made.	1	1	1
9. Reagents that need accurate measurement are identified.	1	1	1
10. Lab trials are stated.	1	1	1
11. Repeats are stated.	1	1	1
12. Chemistry vocabulary is used correctly.	1	1	1
13. Limitations of the experimental design are described.	1	1	1
TOTAL:	13	13	13

reduce the leakage of heat from the outside. However, in a real life situation, the outer layer of the can will not be isolated. For the measurement purposes thermometer will be placed inside the drink.



Beaker with water and the ionic compound.

According to the calculations outlined in the hypothesis section, it is required 5.3 g of pure ammonium chloride (NH4Cl) to lower the temperature of 100 cm3 of distilled water by 5oC. However, this does not take in account that an ionic compound will be added directly to the drink, but to the liquid that surrounds submerged container with the drink. Thus, the volume of the surrounding liquid and the mass of the ionic compound have found experimentally.

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- Long hair should be tied during the lab.
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- Ammonium chloride (NH_4Cl) may release some chlorine ions during the experiment. It is important to do not inhale or smell any compounds during this lab.

Observations:

Volume of the surrounding liquid (H_2O) , mL	Mass of the ammonium chloride (NH_4Cl), g	Resulted temperature of the drink °C $T_f - T_i = -3°C$
100	10.6	21.1
100	10.6	18.4 − 22.4 = −4°C
200 ml	13.3	16.4 − 22.45 = −6°C
100 ml	11.4	~~21.4~~ − 21.4 = −4°C
100 ml	12.8	17.0 − 22.4 = −5.4°C