

Lab 1

Morgan Benavidez

August 25th, 2022

1 Pre-Lab

1.1 Purpose

1. Understand the process of heat transfer, with respect to enthalpy.
2. Utilize and understand the process of calorimetry to determine heat transfer.
3. Apply Hess's law to calculate enthalpy through direct and indirect processes.
4. Determine the enthalpy of formation and enthalpy of reaction.

1.2 Definitions

Calorimetry - A procedure that allows us to measure the amount of energy (as heat given off or absorbed) under carefully controlled conditions.

Calorimeter - An object used for calorimetry. A simple calorimeter just consists of a thermometer attached to a metal container full of water suspended above a combustion chamber.

Enthalpy - The sum of the internal energy and the product of the pressure and volume of a thermodynamic system. Represented with the letter H. $H = E + PV$

Endothermic - Absorbs heat.

Exothermic - Releases heat.

Enthalpy change - The heat change is denoted by ΔH .

1.3 Equations

To measure the amount of energy (in the form of heat) absorbed or given off, we can use this equation:

$$q_{solution} = m_{solution} * c_{solution} * \Delta T_{solution} \quad (1)$$

q = heat, m = mass, c = specific heat capacity of water and ΔT = change in temperature in ($^{\circ}\text{C}$ or K).

We will assume that no heat is lost to the universe (surrounding area) by the calorimeter. With this assumption, the energy as heat lost by the reaction will equal the energy gained by the solution. We can represent this in the following equation:

$$q_{reaction} = -q_{solution} \quad (2)$$

If you combine the two above equations you get:

$$q_{reaction} = -m_{solution} * c_{solution} * \Delta T_{solution} \quad (3)$$

We determine the heat change for the reaction (unknown) by determining the heat change for the solution (known). The heat change is called *enthalpy change*. This is represented as follows:

$$\Delta H = \text{Enthalpy Change} \quad (4)$$

To find the change in Temperature:

$$\Delta T = T_{initial} - T_{final} \quad (5)$$

If answer (ΔT) is positive, this means that heat was released to the solution and the reaction is *exothermic*. When q is calculated, then the value will be negative. The negative sign simply indicates the direction of heat flow, not the quantity.

If the answer (ΔT) is negative, heat was absorbed from the solution, the reaction is *endothermic* and q will be positive.

Once the heat quantity ($q_{reaction}$) is calculated, **at constant pressure**:

$$q_{reaction} = \Delta_{reaction}H \quad (6)$$

Calculate $\Delta_{reaction}H$ per mole of substance:

$$\Delta_{reaction}H = \frac{q_{reaction}}{mol_{substance}} \quad (7)$$

Hess's Law Formula:

$$\Delta H_{net} = \sum \Delta H_r \quad (8)$$

ΔH_{net} = net change in enthalpy (9)

$\sum \Delta H_r$ = sum change in enthalpy reactions (10)

2 Flowcharts

Please see the following two pages for flowcharts for both Part One and Part Two.

PART ONE



PART TWO



3 Waste Collection

3.1 Part One

3.1.1 Mg metal - Dispose of your solution in the labeled waste container.

3.1.2 MgO - Dispose of your solution in the labeled waste container.

3.2 Part Two

3.2.1 Reaction I - Dispose of your solution down the drain.

3.2.2 Reaction II - Dispose of your solution in the labeled waste container.

3.2.3 Reaction III - Dispose of your solution in the labeled waste container.

4 Declaration of Academic Integrity

I certify that this is my own work and I understand that if I am found to be in violation of the honor code, I will be subject to the highest penalty.

Morgan Benavidez

5 Calculations

5.1 Part One

Part I: Determination of $\Delta H_{formation}$ of MgO

<i>Data</i>	<i>Magnesium</i>	<i>Magnesium Oxide</i>
Mass of weigh dish	1.96 g	2.045 g
Mass of dish + sample	2.06 g	2.295 g
Mass of Sample	0.1 g	0.250 g
Molecular weight of Mg/MgO	24.305 $\frac{g}{mol}$	40.305 $\frac{g}{mol}$
Number of moles of Mg/MgO	0.0041 mols	0.0062 mols
Mass of Styrofoam cups	5.340 g	5.430 g
Mass of cup + reaction mixture	32.180 g	29.110 g
Mass of Reaction Mixture	26.84 g	23.680 g
$T_{initial}$ (from graph)	$23^{\circ}\text{C} + 273 = 296 \text{ K}$	$22.5^{\circ}\text{C} + 273 = 295.5 \text{ K}$
T_{final} (from graph)	$34.3^{\circ}\text{C} + 273 = 307.3 \text{ K}$	$27.4^{\circ}\text{C} + 273 = 300.4 \text{ K}$
ΔT (from graph)	11.3 K	4.9 K
q, heat transferred	-1.27 kJ	-0.49 kJ
$\Delta H_{reaction}$	-308.75 $\frac{kJ}{mol}$	-78.30 $\frac{kJ}{mol}$
$\Delta H_{formation}$ of MgO	-472.25 $\frac{kJ}{mol}$	-472.25 $\frac{kJ}{mol}$

1. Calculate the mass of your metal by difference.

$$\text{Magnesium} = 2.06g - 1.96g = 0.1g \quad (11)$$

$$\text{Magnesium Oxide} = 2.295g - 2.045g = 0.250g \quad (12)$$

2. Calculate the molecular weight and then the number of moles of your metal and metal oxide.

$$\text{Molecular weight of Magnesium} = 24.305 \frac{g}{mol} \quad (13)$$

$$\frac{24.305 \frac{g}{mol}}{0.1g} = 0.0041 \text{ mols} \quad (14)$$

$$\text{Molecular weight of Magnesium Oxide} = 40.305 \frac{g}{mol} \quad (15)$$

$$\frac{40.305 \frac{g}{mol}}{0.250g} = 0.0062 \text{ mols} \quad (16)$$

3. Calculate $\Delta T := T_{final} - T_{initial}$

$$\Delta T_{Mg} = 307.3K - 296K = 11.3K \quad (17)$$

$$\Delta T_{MgO} = 300.4K - 295.5K = 4.9K \quad (18)$$

4. Calculate the mass of the reaction mixture (Solid added to a solution) by difference.

$$\text{Magnesium Reaction Mixture} = 32.180g - 5.340g = 26.84g \quad (19)$$

$$\text{Magnesium Oxide Reaction Mixture} = 29.110g - 5.430g = 23.680g \quad (20)$$

5. Calculate q, the heat transferred with the metal and metal oxide.

$$q_{Mg} = 26.84g * 4.184 \frac{J}{g * K} * 11.3K = 1,268.973J \quad (21)$$

$$q_{Mg\,Reaction} = -q_{Mg\,Solution} = -1,268.973J \quad (22)$$

$$\frac{-1,268.973J}{1000} = -1.27kJ \quad (23)$$

$$q_{MgO} = 23.680g * 4.184 \frac{J}{g * K} * 4.9K = 485.478J \quad (24)$$

$$q_{MgO\,Reaction} = -q_{MgO\,Solution} = -485.478J \quad (25)$$

$$\frac{-485.478J}{1000} = -0.49kJ \quad (26)$$

6. Calculate $\Delta H_{reaction}$ per mole of Mg and MgO

$$\Delta H_{Mg} = \frac{-q_{Mg\,Solution}}{mol_{Mg}} = \frac{-1,268.97J}{0.0041mol} = -308,751.825 \frac{J}{mol} \quad (27)$$

$$\Delta H_{Mg} = \frac{-308,751.825 \frac{J}{mol}}{1000} = -308.75 \frac{kJ}{mol} \quad (28)$$

$$\Delta H_{MgO} = \frac{-q_{MgO\,Solution}}{mol_{MgO}} = \frac{-485.478J}{0.0062mol} = -78,302.89 \frac{J}{mol} \quad (29)$$

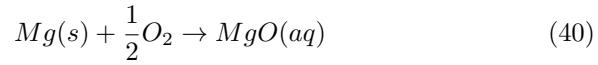
$$\Delta H_{MgO} = \frac{-78,302.89 \frac{J}{mol}}{1000} = -78.30 \frac{kJ}{mol} \quad (30)$$

7. Utilizing the three equations provided and your knowledge of Hess's law, show how they can be manipulated to determine the equation and final value for $\Delta H_{formation}$ of MgO. Show balanced equations, cancelled out compounds, and enthalpies of all reactions as seen in the example of this lab.

$$\Delta H_{MgO formation} = \Delta H_{Mg} + (-\Delta H_{MgO}) + (-241.8 \frac{kJ}{mol}) \quad (31)$$

$$\Delta H_{MgO formation} = -308.75 \frac{kJ}{mol} + 78.30 \frac{kJ}{mol} - 241.8 \frac{kJ}{mol} \quad (32)$$

$$\Delta H_{MgO formation} = -472.25 \frac{kJ}{mol} \quad (33)$$



5.2 Part Two

Part II

Data	Reaction I	Reaction II	Reaction III
$T_{initial}$ (from graph)	295.9 K	295.3 K	294.5 K
T_{final} (from graph)	309.3 K	296.7 K	304.9 K
ΔT (from graph)	13.4 K	1.4 K	10.4 K
q, heat transferred	2.89 kJ	0.307 kJ	2.24 kJ
Number of moles of reactant	0.1 mols	0.1 mols	0.1 mols
$\Delta H_{reaction}$	-28.9 $\frac{kJ}{mol}$	-3.07 $\frac{kJ}{mol}$	-22.4 $\frac{kJ}{mol}$
$\Delta H_{reaction} NH_4Cl$	-25.83 $\frac{kJ}{mol}$	-25.83 $\frac{kJ}{mol}$	-

1. Calculate $\Delta T : \Delta T = T_{final} - T_{initial}$

$$\Delta T_{R1} = 309.3K - 295.9K = 13.4K \quad (41)$$

$$\Delta T_{R2} = 296.7K - 295.3K = 1.4K \quad (42)$$

$$\Delta T_{R3} = 304.9K - 294.5K = 10.4K \quad (43)$$

2. Calculate the heat transferred q for each of the three reactions. Use 1.03 g/mL for the density of all solutions.

$$50mL * \frac{1.03g}{1mL} = 51.5g \quad (44)$$

$$q_{R1} = 51.5g * 4.184 \frac{J}{g * K} * 13.4K = 2,887.38J \quad (45)$$

$$q_{R1} = 2.89kJ \quad (46)$$

$$q_{R2} = 51.5g * 4.184 \frac{J}{g * K} * 1.4K = 301.67J \quad (47)$$

$$q_{R2} = 0.307kJ \quad (48)$$

$$q_{R3} = 51.5g * 4.184 \frac{J}{g * K} * 10.4K = 2,240.95J \quad (49)$$

$$q_{R3} = 2.24kJ \quad (50)$$

3. Calculate the number of moles of each reactant from the provided concentration (for Reaction 1, you are solving for moles of HCl; for Reaction II, you are solving for moles of NH_4Cl ; and for Reaction III, you are solving for moles of NH_3).

$$50mL * \frac{2.0mol}{1L} * \frac{1L}{1000mL} = 0.1mol \quad (51)$$

Same for all 3 Reactions

4. Calculate $\Delta H_{reaction}$ per mole of HCl, NH_4Cl and NH_3 in $\frac{kJ}{mol}$

$$\Delta H_{R1} = \frac{-q_{R1}}{mol} = \frac{-2.89kJ}{0.1mol} = -28.9 \frac{kJ}{mol} \quad (52)$$

$$\Delta H_{R2} = \frac{-q_{R2}}{mol} = \frac{-0.307kJ}{0.1mol} = -3.07 \frac{kJ}{mol} \quad (53)$$

$$\Delta H_{R3} = \frac{-q_{R3}}{mol} = \frac{-2.24kJ}{0.1mol} = -22.4 \frac{kJ}{mol} \quad (54)$$

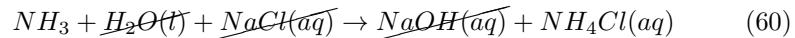
5. Using the data collected for Reactions I and II, apply Hess's law to determine ΔH for NH_4Cl .

$$\Delta H_{NH_4Cl} = \Delta H_{R1} + (-\Delta H_{R2}) \quad (55)$$

$$\Delta H_{NH_4Cl} = -28.9 \frac{kJ}{mol} + 3.07 \frac{kJ}{mol} \quad (56)$$

$$\Delta H_{NH_4Cl} = -25.83 \frac{kJ}{mol} \quad (57)$$

6. Using the data collected for Reactions I and II, and your knowledge of Hess's law, show how they can be manipulated to determine the equation and final value for $\Delta H_{reaction}$ of NH_4Cl . Show balanced equations, cancelled-out compounds, and enthalpies of all reactions as seen in the example in this lab.



7. Record Percent Error:

$$\text{Percent Error} = \frac{\Delta H_{R3} - \Delta H_{R1+R2}}{\Delta H_{R3}} \quad (63)$$

$$\text{Percent Error} = \frac{-22.4 \frac{kJ}{mol} - ((-28.9 \frac{kJ}{mol}) + (-3.07 \frac{kJ}{mol}))}{-22.4 \frac{kJ}{mol}} * 100 \quad (64)$$

$$\text{Percent Error} = \frac{-22.4 \frac{kJ}{mol} - (-31.97 \frac{kJ}{mol})}{-22.4 \frac{kJ}{mol}} * 100 \quad (65)$$

$$\text{Percent Error} = \frac{9.57 \frac{kJ}{mol}}{-22.4 \frac{kJ}{mol}} * 100 \quad (66)$$

$$\text{Percent Error} = -0.4272 * 100 \quad (67)$$

$$\text{Percent Error} = -42.72 \quad (68)$$

Discussion and Conclusion

The major concepts covered in this experiment were learning how to measure and calculate energy changes for reactions using calorimetry. We built a calorimeter using 2 Styrofoam cups. We used two cups to increase the insulation and help reduce the amount of heat lost to the universe. We assumed that no energy is transferred beyond the solution (system) and because of this assumption we can say that the energy lost by the reaction will equal the energy gained by the solution.

Based on the Delta T values derived from our experiments, the solution increase in temperature and gained energy. Which means the reaction itself lost energy/heat, therefore we can say that the reactions were exothermic. A general rule of thumb is that if a chemical reaction has a positive Delta H, the reaction is endothermic. If the opposite is true and Delta H is negative, the reaction is exothermic. The Delta H in our experiments was negative, making our reactions exothermic.

We used Hess's Law in Part I and II by first calculating the Delta H of each individual reaction (heat quantity / mole of the substance). We can then sum the enthalpy change for these individual reactions to find the total enthalpy change. (Please see pages 8, 9 and 11 of my lab report for a detailed analysis).

The Percent Error for Part I is $-601.8 \text{ kJ/mol} - (-472.25 \text{ kJ/mol}) = -129.55 \text{ kJ/mol} / -601.8 \text{ kJ/mol} = 0.215 * 100 = 21.5\%$.

The Delta H reaction I got for Reaction III was -22.4 kJ/mol and the Delta H reaction I got for Reaction III using Hess's law was -25.83 kJ/mol. $-25.83 \text{ kJ/mol} - (-22.4 \text{ kJ/mol}) = -3.43$ kJ/mol / $-25.83 \text{ kJ/mol} = 0.1328 * 100 = 13.28\%$ error. These values are close to each other, and I think that is a good percent error.

In conclusion, the lab results sufficiently demonstrated exothermic reactions, gave me plenty of practice with interpreting, extracting, and calculating the data we produced from our experiments using various chemical formulas including Hess's Law. Our results were also within a reasonable percent error.

6 Signatures from Lab

Please see the next few pages for things that needed to be signed off on.

G. C. 2-1: LABORATORY REPORT

Name: Morgan Benavidez

Lab partners' names: _____

YOU MUST SHOW ALL CALCULATIONS TO RECEIVE CREDIT FOR THEM!

Proper use of significant figures and units is part of the grading.

Part I: Determination of $\Delta H_{\text{formation}}$ of MgO

Data	Magnesium	Magnesium Oxide
Mass of weigh dish	1.96 g	2.045 g
Mass of dish + sample	2.06 g	2.295 g
Mass of sample	$2.06 \text{ g} - 1.96 \text{ g} = .1 \text{ g}$.250 g
Molecular weight of Mg/MgO	24.305 g/mol	40.305 g/mol
Number of moles of Mg/MgO	.00411 mol	.0062 mol
Mass of Styrofoam cups	5.340 g	5.430 g
Mass of cup + reaction mixture	29.180 g	29.110 g
Mass of reaction mixture	23.840 g	23.680 g
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q, heat transferred	-1268.97 J	-0.49 KJ
$\Delta H_{\text{reaction}}$	-308.75 KJ/mol	-78.30 KJ/mol
$\Delta H_{\text{formation}}$ of MgO	-472.25 KJ/mol	

Syrofoam (empty) - 2.150 g
 $+ 2.190 \text{ g}$
 $\hline 5.340 \text{ g}$

My
 $q = 26.84 \text{ g} \times 4.184 \text{ J/g-K} \times 11.3 \text{ K}$
 $q_{\text{rxn}} = 1,268.973 \text{ J}$

$q_{\text{rxn}} = -q_{\text{sol}}$
 $q_{\text{rxn}} = -1,268.97$

$\Delta H_1 = \frac{-1,268.97 \text{ J}}{0.00411 \text{ mol}} = -308,751.825$

MgO
 $q_{\text{sol}} = 23.680 \text{ g} / (4.184 \text{ J/g-K}) \times 4.9 \text{ K}$

$485.478 \text{ J} = -485.478 \text{ J}$

$.0062 \text{ mol} = -78,302.89 = \Delta H_2$

MgO = mass of molar (water) ΔT

G.C. 2-I: LABORATORY REPORTName Morgan Bennidez

Lab partners' names:

YOU MUST SHOW ALL YOUR CALCULATIONS TO RECEIVE CREDIT FOR THEM!**Proper use of significant figures and units is part of the grading.****Part II**

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q, heat transferred	2.89 kJ	(0.307 kJ)	2.24 kJ
Number of moles of reactant	0.1 mol	0.1 mol	0.1 mol
ΔH _{reaction}	-28.9 kJ/mol	-3.07 kJ/mol	-22.9 kJ/mol
ΔH reaction NH ₄ Cl (calculated from Reactions I and II)	$\Delta H_1 + (-\Delta H_2) = -25.83 \frac{\text{kJ}}{\text{mol}}$		

Students must share data with entire group before they leave the lab. All students should have data for both Parts I and II. You will be required to complete ALL calculations associated with both Parts I and II.

$$\frac{\Delta H_3 - \Delta H_{1+2}}{\Delta H_3} \times 100 = -15.38\% \text{ error}$$

3 Waste Collection

3.1 Part One

3.1.1 Mg metal - Dispose of your solution in the labeled waste container.

3.1.2 MgO - Dispose of your solution in the labeled waste container.

3.2 Part Two

3.2.1 Reaction I - Dispose of your solution down the drain.

3.2.2 Reaction II - Dispose of your solution in the labeled waste container.

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Morgan Bearvillez

