Lab 1

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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:

$$q_{solution} = m_{solution} * c_{solution} * \Delta T_{solution} \tag{1}$$

q = heat, m = mass, c = specific heat capacity of water and ΔT = change in temperature in ($^\circ C$ 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:

$$q_{reaction} = -q_{solution} \tag{2}$$

If you combine the two above equations you get:

$$q_{reaction} = -m_{solution} * c_{solution} * \Delta T_{solution} \tag{3}$$

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:

$$\Delta H = EnthalpyChange \tag{4}$$

To find the change in Temperature:

$$\Delta T = T_{initial} - T_{final} \tag{5}$$

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***:

$$q_{reaction} = \Delta_{reaction}H \tag{6}$$

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

$$\Delta_{reaction}H = \frac{q_{reaction}}{mol_{substance}} \tag{7}$$

Hess's Law Formula:

$$\Delta H_{net} = \sum \Delta H_r \tag{8}$$

$$\Delta H_{net} = \text{net change in enthalpy} \tag{9}$$

$$\sum \Delta H_r = \text{sum change in enthalpy reactions} \tag{10}$$

2 Flowcharts

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

PART ONE

Obtain LabQuest Device, 2 Styrofoam cups, a cover and a sample of Mg and MgO. → Clean and dry Styrofoam cups. → Place one cup in the other, weigh them on balance and record your data. → Weigh your Mg metal sample and record the exact weight of metal on data sheet.

Prepare a setup for calorimetric measurements. (See Figure 1) ← Fill a 400 mL beaker with 200 mL of deionized water and set aside. ← Cover the dish with metal sample with parafilm. ← Cut Mg metal sample into small pieces.

Prepare your Vernier LabQuest 2 interface for data collection. → Follow instructions 1 – 21 under Part 1: Data Collection → Pour exactly 25.0 mL of 1.0 M HCl into your calorimeter. → Place the temp probe into the solution through the hole in your lid. (Do not submerge insulated portion).

Once data collection is complete, hit STOP. ← Continue recording temp for 2 mins ← Monitor time, at 1 minute add Mg solid into cups quick while stirring. ← Begin data collection – Press Start/Stop Button on LabQuest Interface.

Remove probe from the solution and wash in distilled water and dry with kim-wipe. → Measure the total mass of the cups + the mixture and record your data. → Dispose of your solution in the labeled waste container. → Wash your cups out and prepare for next part.

Complete this entire procedure again using MgO. ← Retrieve data.

PART TWO

Obtain LabQuest Device, 2 Styrofoam cups, a cover, NaOH and Hcl → Clean and dry Styrofoam cups. → Place one cup in the other, weigh them on balance and record your data. → Prepare a setup for calorimetric measurements. (See Figure 1)

↓

Clean graduated cylinder. ← Reaction 1 – Measure 50 mL of 2.0 M HCl in a graduated cylinder and transfer to a beaker for future. ← Follow instructions 1 – 21 under Part 1: Data Collection ← Prepare your Vernier LabQuest 2 interface for data collection.

↓

Measure 50 mL of 2.0 M NaOH in a graduated cylinder and transfer to a beaker for future. → Pour the 50 mL HCl solution into the calorimeter. → Lower temp probe into HCl and click Start button. → Wait 15 sec and add NaOH all at once and continue recording temp for 180 s, then click Red stop button.

↓

Reaction 2 – Repeat the above procedure but with 2.0 M NaOH as the starting solution and 2.0 M NH4Cl as other. ← Retrieve your data and record the initial and maximum temps. ← Reaction 1 - Dispose of solution down the drain for. ← Rinse and dry probe in distilled water for 2 minutes.

↓

Reaction 2 - Dispose of solution in properly labeled waste container. → Retrieve your data and record the initial and maximum temps. → Reaction 3 – Repeat the above procedure but with 2.0 M HCl as the starting solution and 2.0 M NH3 (under fumehood) as other. → Reaction 3 - Dispose of solution in properly labeled waste container.

↓

Turn LabQuest 2 off and clean all equipment. ← Retrieve your data and record the initial and maximum temps.

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

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	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°C + 273 = 296 K	22.5°C + 273 = 295.5 K
T_{final} (from graph)	34.3°C + 273 = 307.3 K	27.4°C + 273 = 300.4 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.

$$\text{Magnesium} = 2.06g - 1.96g = 0.1g \tag{11}$$

$$\text{Magnesium Oxide} = 2.295g - 2.045g = 0.250g \tag{12}$$

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

$$\text{Molecular weight of Magnesium} = 24.305\frac{g}{mol} \tag{13}$$

$$\frac{24.305\frac{g}{mol}}{0.1g} = 0.0041 \text{ mols} \tag{14}$$

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

$$\frac{40.305\frac{g}{mol}}{0.250g} = 0.0062 \text{ mols} \tag{16}$$

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

$$\Delta T_{Mg} = 307.3K - 296K = 11.3K \tag{17}$$

$$\Delta T_{MgO} = 300.4K - 295.5K = 4.9K \tag{18}$$

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

$$\text{Magnesium Reaction Mixture} = 32.180g - 5.340g = 26.84g \tag{19}$$

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

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

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

$$q_{MgReaction} = -q_{MgSolution} = -1,268.973J \tag{22}$$

$$\frac{-1,268.973J}{1000} = -1.27kJ \tag{23}$$

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

$$q_{MgOReaction} = -q_{MgOSolution} = -485.478J \tag{25}$$

$$\frac{-485.478J}{1000} = -0.49kJ \tag{26}$$

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

$$\Delta H_{Mg} = \frac{-q_{MgSolution}}{mol_{Mg}} = \frac{-1,268.97J}{0.0041mol} = -308,751.825\frac{J}{mol} \tag{27}$$

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

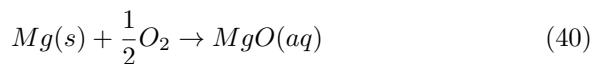$$\Delta H_{MgO} = \frac{-q_{MgOSolution}}{mol_{MgO}} = \frac{-485.478J}{0.0062mol} = -78,302.89\frac{J}{mol} \tag{29}$$

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

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.

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

$$\Delta H_{MgO\,formation} = -308.75\frac{kJ}{mol} + 78.30\frac{kJ}{mol} - 241.8\frac{kJ}{mol} \tag{32}$$

$$\Delta H_{MgO\,formation} = -472.25\frac{kJ}{mol} \tag{33}$$

$$Mg(s) + 2\cancel{HCl(aq)} \rightarrow \cancel{MgCl_2(aq)} + \cancel{H_2(g)} \tag{34}$$

$$+ \tag{35}$$

$$\cancel{H_2O(l)} + \cancel{MgCl_2(aq)} \rightarrow 2\cancel{HCl} + MgO(aq) \tag{36}$$

$$+ \tag{37}$$

$$\cancel{H_2(g)} + \frac{1}{2}O_2 \rightarrow \cancel{H_2O(l)} \tag{38}$$

$$= \tag{39}$$

$$Mg(s) + \frac{1}{2}O_2 \rightarrow MgO(aq) \tag{40}$$

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

$$\Delta T_{R1} = 309.3K - 295.9K = 13.4K \tag{41}$$

$$\Delta T_{R2} = 296.7K - 295.3K = 1.4K \tag{42}$$

$$\Delta T_{R3} = 304.9K - 294.5K = 10.4K \tag{43}$$

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

$$50mL * \frac{1.03g}{1mL} = 51.5g \tag{44}$$

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

$$q_{R1} = 2.89kJ \tag{46}$$

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

$$q_{R2} = 0.307kJ \tag{48}$$

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

$$q_{R3} = 2.24kJ \tag{50}$$

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_Cl; and for Reaction III, you are solving for moles of NH_3).

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

Same for all 3 Reactions

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

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

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

$$\Delta H_{R3} = \frac{-q_{R3}}{mol} = \frac{-2.24kJ}{0.1mol} = -22.4\frac{kJ}{mol} \tag{54}$$
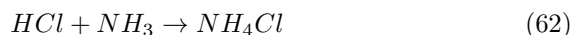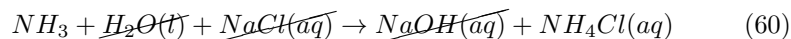
5. Using the data collected for Reactions I and II, apply Hess's law to determine ΔH for NH_4Cl.

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

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

$$\Delta H_{NH_4Cl} = -25.83\frac{kJ}{mol} \tag{57}$$

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.

$$HCl(aq) + \cancel{NaOH(aq)} \rightarrow \cancel{NaCl(aq)} + \cancel{H_2O(l)} \tag{58}$$

$$+ \tag{59}$$

$$NH_3 + \cancel{H_2O(l)} + \cancel{NaCl(aq)} \rightarrow \cancel{NaOH(aq)} + NH_4Cl(aq) \tag{60}$$

$$= \tag{61}$$

$$HCl + NH_3 \rightarrow NH_4Cl \tag{62}$$

7. Record Percent Error:

$$\text{Percent Error} = \frac{\Delta H_{R3} - \Delta H_{R1+R2}}{\Delta H_R 3} \tag{63}$$

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

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

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

$$\text{Percent Error} = -0.4272 * 100 \tag{67}$$

$$\text{Percent Error} = -42.72 \tag{68}$$