

Experiment 15

Determination of the Enthalpy of a Reaction using Hess's Law

Pre-Lab Assignment

Before coming to lab:

- Read the lab thoroughly.
- Answer the pre-lab questions that appear at the end of this lab exercise.

Purpose

The heat of formation for magnesium oxide (MgO) will be measured indirectly by calculating the enthalpies for three other reactions and summing them via Hess's Law.

Background

The standard heat of formation, ΔH_f° , of a compound is defined to be the heat evolved or absorbed when one mole of the compound is formed from its constituent elements in their standard, or most stable, states. The standard state of a compound or element is taken to be the most stable form of the substance at one atmosphere of pressure and 25°C. Naturally the heat of formation of an element in its most stable form is defined to be zero since an element has no constituents.

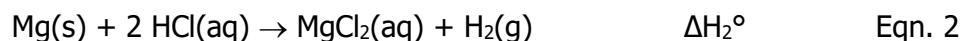
In this experiment, the heat of formation for magnesium oxide (MgO) will be determined by the following balanced reaction (Eqn. 1).



Note that the equation is balanced to form only one mole of MgO(s) which requires the use of the fractional coefficient for O₂. All physical states for each element or compound must be specified since enthalpy depends on the physical state of the substance.

Experimentally it would be difficult to measure the heat of formation for MgO directly, especially in a Styrofoam calorimeter, due to the chemicals involved. Fortunately, we can employ a law formulated by Germain Hess to evaluate the heat of formation for MgO indirectly by using a series of reactions that are easier to measure or whose enthalpies of reactions are available in standardized tables. Hess's Law states that if a series of reactions can be added together to make a certain overall reaction, then their enthalpies can also be added together in the same manner to find the enthalpy for the reaction of interest.

For this experiment, the reactions used are shown in Eqn. 2-4:



Summing these reactions and cancelling any compound that appears on both the reactants and products side results in Eqn. 1. Thus, summing their enthalpies also gives the equation for finding the heat of formation for MgO by Eqn. 5

$$\Delta H_f^\circ \text{ of MgO} = \Delta H_2^\circ + \Delta H_3^\circ + \Delta H_4^\circ \quad \text{Eqn. 5}$$

Example Exercise: Using Hess's Law to Find ΔH

Find ΔH of reaction for: $\text{N}_2(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow 2 \text{NO}_2(\text{g})$ in kJ using reactions A and B below. Clearly indicate any changes needed for each reaction A and B to give the desired reaction.

	Reaction	ΔH_{rxn}	Change/Multiplier
A	$4 \text{NO}(\text{g}) \rightarrow 2 \text{N}_2(\text{g}) + 2 \text{O}_2(\text{g})$	$\Delta H_A = -361.0 \text{ kJ}$?
B	$\text{NO}(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g})$	$\Delta H_B = -57.06 \text{ kJ}$?

Step 1: Reverse Reaction A and multiply all coefficients by $\frac{1}{2}$. Make the same changes to the enthalpy of Reaction A by multiplying the value by $-\frac{1}{2}$.



The change or multiplier needed for Reaction A is $-1/2$.

Step 2: Multiply all coefficients in Reaction B by 2. Make the same change to the enthalpy of Reaction B by multiplying the value by 2.



The change or multiplier needed for Reaction B is 2.

Step 3: Sum all the new ΔH values together

$$\Delta H = 180.5 \text{ kJ} + (-114.12 \text{ kJ}) = 66.4 \text{ kJ}$$

To measure each heat of reaction for Eqn. 2 and 3, coffee-cup calorimetry will be used. For Eqn. 2, solid Mg will be dissolved in $\text{HCl}(\text{aq})$ and the resulting change in temperature will be recorded. The heat released by the reaction will be absorbed by the surrounding solution, which is assumed to have a heat capacity of $4.184 \text{ J/g } ^\circ\text{C}$, the same as water. Assuming no heat is lost to the coffee-cup calorimeter, we can calculate the heat produced by the reaction using Eqn. 8:

$$q_{\text{rxn}} = -q_{\text{soln}} \quad \text{Eqn. 5}$$

The experiment has been designed so the solid Mg will be the limiting reactant. Once the heat produced by the reaction has been calculated, it can be divided by the moles of limiting reactant to give the enthalpy for Eqn. 2.

For Eqn. 3, solid MgO(s) will be dissolved in HCl(aq) which is the *reverse* of Eqn. 3. Solid MgO will be the limiting reactant. Remember that reversing reactions has the effect of changing the sign (positive or negative) of the reaction's enthalpy. Otherwise, the process to measure and calculate the enthalpies for Eqn 2 and 3 will be identical.

Example Problem: Finding ΔH_2

Using the same coffee-cup as the example above, 0.55 g of Mg(s) is dissolved in 61.050 g of HCl(aq) held initially at 20.5°C. The resulting reaction causes the temperature to rise to 61.2°C. Calculate ΔH for this reaction in kJ/mol. Assume the heat capacity for the solution is the same as water, 4.184 J/g °C.

Step 1: Find the ΔT for the solution

$$61.2^\circ\text{C} - 20.5^\circ\text{C} = 40.7^\circ\text{C}$$

Step 2: Find the heat absorbed by the solution, q_{soln}

$$q_{\text{soln}} = (61.050 \text{ g} + 0.55 \text{ g})(4.184 \text{ J/g } ^\circ\text{C})(40.7^\circ\text{C}) = 10500 \text{ J}$$

Step 4: Find the heat released by the reaction, q_{rxn}

$$q_{\text{rxn}} = -q_{\text{soln}} = -10500 \text{ J}$$

Step 5: Find the moles Mg

$$0.55 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} = 0.023 \text{ mols Mg}$$

Step 6: Find ΔH_2°

$$\Delta H_2^\circ = \frac{q_{\text{rxn}}}{\text{mols Mg}} = \frac{-10500 \text{ J}}{0.023 \text{ mols Mg}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = -456 \text{ kJ/mol Mg}$$

Eqn. 4 is the formation of liquid water which has a known, tabulated value of -285.85 kJ/mol that can be used in lieu of measuring the reaction experimentally.

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Procedure

Part I: Determination of ΔH_2°

1. Gather two Styrofoam coffee-cups and stack them together. Collect a cover, a split stopper, a thermometer, and a clamp. Use any glass beaker that will fit the cups and prevent tipping. Assemble the apparatus as shown in Fig. 1.

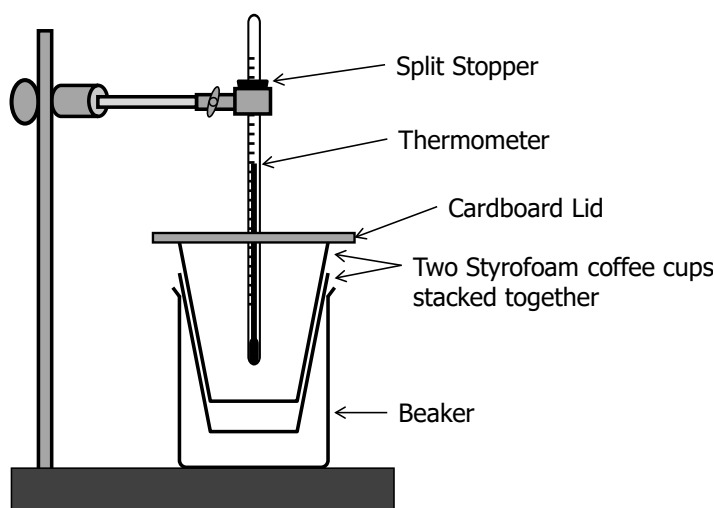


Figure 1: Coffee-cup calorimeter apparatus

2. Remove the coffee cups and cover from the calorimeter apparatus. Weigh the coffee cups and cover. Record this weight in your data sheet.
 3. Add approximately 60 mL of 1.0 M HCl(aq) in the cups with cover and reweigh. Record this weight in your data sheet.
 4. Using a weigh boat, measure out approximately 0.55–0.60 g Mg solid. Record the exact weight in your data sheet.
 5. Measure the temperature of the HCl solution (T_i).
 6. Transfer the Mg solid to the HCl solution in the cups and **immediately** cover. Use the thermometer to gently stir and record the **maximum** temperature reached (T_f).
- Caution:** $H_2(g)$ is being produced. Avoid open flame.
7. Quickly check inside the cups to ensure all of the Mg solid has dissolved. If any pieces of Mg solid remain, place the cover back on the cups and monitor the temperature. Record the overall **maximum** temperature reached as T_f .
 8. Neutralize the excess acid by slowly adding small scoops of baking soda to the flask. Swirl the contents of the cups to mix. Be careful: the mixture will foam up as it reacts. Continue until the mixture stops fizzing when more baking soda is added.

- Dispose of the neutralized solution in the sink with running water. Rinse out the cup with deionized water. Dry the inside and outside of the cups with a paper towel. Repeat Steps 2–8 for a second trial.
- Calculate ΔH_2° in kJ/mol for each trial.

Part II: Determination of ΔH_3°

- Repeat Steps 1–9 from Part I, using approximately 60 mL of 1.0 M HCl(aq) and 0.90–1.0 g MgO(s).
- Calculate ΔH_{rxn} in kJ/mol and then use it to find ΔH_3° for each trial.

Note: ΔH_{rxn} will be the reverse reaction of ΔH_3° .

Part III: Determination of ΔH_f° for MgO

- Using your results from Parts I and II and the tabulated value for ΔH_4° , solve for ΔH_f° of MgO using all values from Trial 1 then all values from Trial 2. Report your average.

Experiment 15—Data Sheet

Name: _____

Part I: Determination of ΔH_2°

	Trial One	Trial Two
1. Mass of calorimeter (g)	_____	_____
2. Mass of calorimeter and HCl (g)	_____	_____
3. Mass of Mg (g)	_____	_____
4. Initial T of solution (T_i soln, °C)	_____	_____
5. Final T of solution (T_f soln, °C)	_____	_____
6. Mass of HCl (g) <i>show calculation:</i>	_____	_____
7. Total Mass of Solution (g) <i>show calculation:</i>	_____	_____
8. Mols of Mg (mols) <i>show calculation:</i>	_____	_____
9. ΔT of solution (ΔT_{soln} , °C) <i>show calculation:</i>	_____	_____
10. Heat absorbed by the soln (q_{soln} , J) <i>show calculation:</i>	_____	_____

Trial One

Trial Two

11. Heat released by the reaction (q_{rxn} , J)
show calculation:

12. Heat of Reaction for Eqn. 2 (ΔH_2° , kJ/mol)
show calculation:

Part II: Determination of ΔH_3°

	Trial One	Trial Two
1. Mass of calorimeter (g)	_____	_____
2. Mass of calorimeter and HCl (g)	_____	_____
3. Mass of MgO (g)	_____	_____
4. Initial T of calorimeter and solution (T_i soln, °C)	_____	_____
5. Final T of calorimeter and solution (T_f soln, °C)	_____	_____
6. Mass of HCl (g) <i>show calculation:</i>	_____	_____
7. Total Mass of Solution (g) <i>show calculation:</i>	_____	_____
8. Mols of MgO (mols) <i>show calculation:</i>	_____	_____
9. ΔT of solution (ΔT_{soln} , °C) <i>show calculation:</i>	_____	_____
10. Heat absorbed by the soln (q_{soln} , J) <i>show calculation:</i>	_____	_____

Trial One

Trial Two

11. Heat released by the reaction (q_{rxn} , J)
show calculation:

12. Heat of Reaction (ΔH , kJ/mol)
show calculation:

13. Heat of Reaction for Eqn. 3 (ΔH_3° , kJ/mol)
show calculation:

Part III: Determination of ΔH_f° for MgO

	Trial One	Trial Two
1. Heat of Reaction for Eqn. 2 (ΔH_2° , kJ/mol)	_____	_____
2. Heat of Reaction for Eqn. 3 (ΔH_3° , kJ/mol)	_____	_____
3. Heat of Reaction for Eqn. 4 (ΔH_4° , kJ/mol)	<u>-285.85</u>	<u>-285.85</u>
4. Heat of Formation for MgO (ΔH_f° , kJ/mol)	_____	_____

show calculation:

5. Average Heat of Formation for MgO (ΔH_f° , kJ/mol)

show calculation:

6. True value for Heat of Formation for MgO (ΔH_f° , kJ/mol)

show calculation:

7. Percent error for the Experiment Heat of Formation for MgO

show calculation:

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3. Calculate ΔH° for the reaction: $2 \text{Al}(s) + 3 \text{Cl}_2(g) \rightarrow 2 \text{AlCl}_3(s)$ using reactions A, B, C, and D below. Clearly indicate any changes needed for each reaction A, B, C, and D to give the desired reaction.

	Reaction	ΔH_{rxn}	Change/Multiplier
A	$2 \text{Al}(s) + 6 \text{HCl}(aq) \rightarrow 2 \text{AlCl}_3(aq) + 3 \text{H}_2(g)$	$\Delta H_A = -1049 \text{ kJ}$	
B	$\text{HCl}(g) \rightarrow \text{HCl}(aq)$	$\Delta H_B = -74.8 \text{ kJ}$	
B	$\text{H}_2(g) + \text{Cl}_2(g) \rightarrow 2 \text{HCl}(g)$	$\Delta H_C = -1845 \text{ kJ}$	
D	$\text{AlCl}_3(s) \rightarrow \text{AlCl}_3(aq)$	$\Delta H_D = -323 \text{ kJ}$	

4. Give **three** experimental errors that may have contributed to any difference between the true value and your average ΔH_f° of $\text{MgO}(s)$.

Experiment 15—Pre-Lab Assignment

Name: _____

For all calculations, show all work and draw a box around the final answers.

1. For what compound will the heat of formation be determined in this lab? _____
2. State Hess's Law in your own words.

3. Calculate ΔH° for the reaction: $\text{C}_2\text{H}_4(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g})$ using reactions A, B, and C below. Clearly indicate any changes needed for each reaction A, B, and C to give the desired reaction.

	Reaction	ΔH_{rxn}	Change/Multiplier
A	$\text{C}_2\text{H}_4(\text{g}) + 3 \text{O}_2(\text{g}) \rightarrow 2 \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{l})$	$\Delta H_{\text{A}} = -1411 \text{ kJ}$	
B	$\text{C}_2\text{H}_6(\text{g}) + 3 \frac{1}{2} \text{O}_2(\text{g}) \rightarrow 2 \text{CO}_2(\text{g}) + 3 \text{H}_2\text{O}(\text{l})$	$\Delta H_{\text{B}} = -1560 \text{ kJ}$	
B	$\text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$	$\Delta H_{\text{C}} = -285.8 \text{ kJ}$	

4. Calculate ΔH° for the reaction: $4 \text{NH}_3(\text{g}) + 5 \text{O}_2(\text{g}) \rightarrow 4 \text{NO}(\text{g}) + 6 \text{H}_2\text{O}(\text{g})$ using reactions A, B, and C below. Clearly indicate any changes needed for each reaction A, B, and C to give the desired reaction.

	Reaction	ΔH_{rxn}	Change/Multiplier
A	$\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{NO}(\text{g})$	$\Delta H_{\text{A}} = +180.5 \text{ kJ}$	
B	$\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightarrow 2 \text{NH}_3(\text{g})$	$\Delta H_{\text{B}} = -91.8 \text{ kJ}$	
B	$2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{g})$	$\Delta H_{\text{C}} = -483.6 \text{ kJ}$	