

Lab 9. Hess's Law

Prelab Assignment

Before coming to lab:

- This exercise does *not* require a report in your lab notebook. Use a *pen* to record your data, observations, calculations and analysis in the spaces provided on the *report pages* of this handout, [pages 5-10](#). The *report pages* ([pages 5 – 10](#)) will be collected in place of a formal lab report.
- Read the lab thoroughly and answer the pre-lab questions that appear at the end of this lab exercise and hand them in at the start of your lab period.

Purpose

This experiment will give you some practice in calorimetric calculations, using Hess's Law, and computing heats of reaction.

Introduction

In this experiment, you will use a Styrofoam-cup calorimeter to measure the heat released by three different reactions. One of the reactions can be expressed as the combination of the other two reactions. Therefore, the heat of reaction of the one reaction should be equal to the sum of the heats of reaction for the other two. This concept is sometimes referred to as the *Hess's Law* or *additivity of heats of reaction*. The primary objective of this experiment is to confirm this law. The reactions we will use in this experiment are:



You should notice that it is possible to combine reactions A and B to yield reaction C. According to Hess's Law, we should also be able to combine the ΔH values for reactions A and B to give ΔH for the reaction C.

Half of the class will do reaction A and half will do reaction B. All teams will do reaction C. To increase the precision of the experiment, results from the two classes will be pooled. With careful attention to experimental technique, we should be able to get reasonably good results.

Take note that each of the three reactions are carried out in aqueous solution. When a reaction is carried out in an *aqueous solution* the energy given off (or taken up) by the system is assumed to be exchanged with the water (since the solutions used are dilute, we will assume that they have the same thermal properties as water). For example, if a reaction releases heat (i.e. an *exothermic reaction*), then the temperature of the water will increase. On the other hand, an *endothermic reaction* will absorb heat from the water, thus causing a decrease in the temperature of the water. This allows for a simple calculation of the heat of the reaction by first measuring the temperature change for the water, and then using the following equation to calculate the heat, q_{water} , absorbed (or given off) by the water.

$$q_{\text{water}} = m_{\text{water}} \cdot C_{\text{water}} \cdot \Delta t_{\text{water}}$$

In this equation m_{water} is the mass of the water (for simplicity, assume 1.00 mL of water weighs 1.00 g), C_{water} is the specific heat capacity of water, 4.184 J/g°C, and $\Delta t_{\text{water}} = t_{\text{final}} - t_{\text{initial}}$. It is assumed that no heat is lost to the calorimeter, an assumption that is not entirely true. You will need to consider this assumption and the impact it has on the measured value for Δt_{water} when analyzing the data from this experiment.

We will also assume that the coffee cup calorimeter is a perfect insulator (i.e. $q_{\text{calorimeter}} = 0$). In other words, we will assume that the water in the solution gains all of the heat lost by the reaction and none is lost to the surrounding or to the cup itself. Again, since these assumptions are not entirely true you will need to evaluate the impact it has on the measured value for Δt_{water} when analyzing the data from this experiment.

$$q_{\text{rxn}} + q_{\text{water}} = 0$$

OR

$$+ q_{\text{water}} = -q_{\text{rxn}}$$

(i.e. the heat gained by the water = heat lost by the reaction)

and the enthalpy change for the reaction, ΔH_{rxn} , is then

$$\Delta H_{\text{rxn}} = \frac{q_{\text{reaction}}}{\text{moles of limiting reactant}}$$

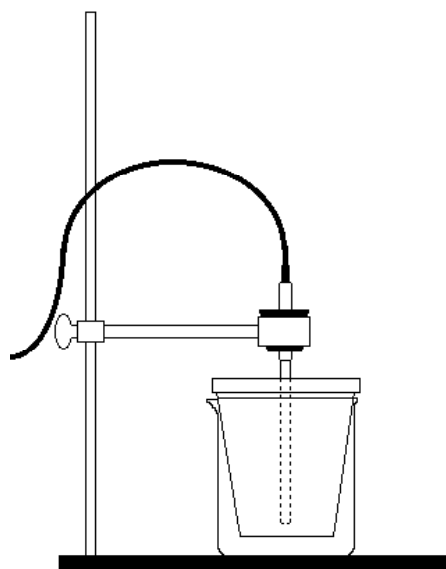


Figure 1. A coffee cup calorimeter in a beaker with a Vernier temperature probe will be used to measure the temperature changes for each of the three reactions. We will assume that the heat lost by the system (i.e. the reacting solutions in the calorimeter) to the calorimeter and the surrounding air is negligible, assumptions that are not entirely true!

Procedure (work in teams of 2)

Before starting, get your assigned reaction from the instructor.

Reaction A (performed by half of the class—your instructor will assign which half of the class!)

1. Obtain and wear goggles.
2. Use a utility clamp and a slit stopper to suspend a temperature probe from a ring stand as shown in Figure 1.
3. Place a Styrofoam cup into a 250-mL beaker as shown in Figure 1. Measure out 100.0 mL of water into the Styrofoam cup. Lower the temperature probe into the solution.
4. Prepare the computer for data collection by opening the *Logger Pro* software and then “Experiment 18” from the *Chemistry with Computers* folder.” The vertical axis has temperature scaled from 15°C to 40°C. The horizontal axis has time scaled from 0 to 200 seconds.
5. Weigh out about 2 grams of solid sodium hydroxide, NaOH, and record the mass to the nearest 0.01 g. Since sodium hydroxide readily picks up moisture from the air, it is necessary to weigh it and proceed to the next step without delay. **Caution:** Handle the NaOH and its resulting solution with care—NaOH is caustic and will cause skin burns and will burn wholes in clothes.
6. Click on the Start button to begin data collection and obtain the initial temperature, t_1 . It may take several seconds for the temperature probe to equilibrate at the temperature of the solution. After three or four readings at the same temperature have been obtained, add the solid NaOH to the Styrofoam cup. Using the stirring rod, stir continuously for the remainder of the 200 seconds or until the temperature maximizes. As soon as the temperature has begun to drop after reaching a maximum, you may terminate the trial by clicking on the Stop button.
7. Choose Data A Table from the Windows menu. From this data, determine the initial temperature, t_1 . Record t_1 in your data table and close the Data A Table window. To determine the final temperature, t_2 , choose Stats from the Analyze menu and find the *maximum* temperature. Record t_2 in your data table.
8. Rinse and dry the temperature probe, Styrofoam cup, and stirring rod. Dispose of the solution as directed by your instructor.

Reaction B (performed by half of the class—your instructor will assign which half of the class!)

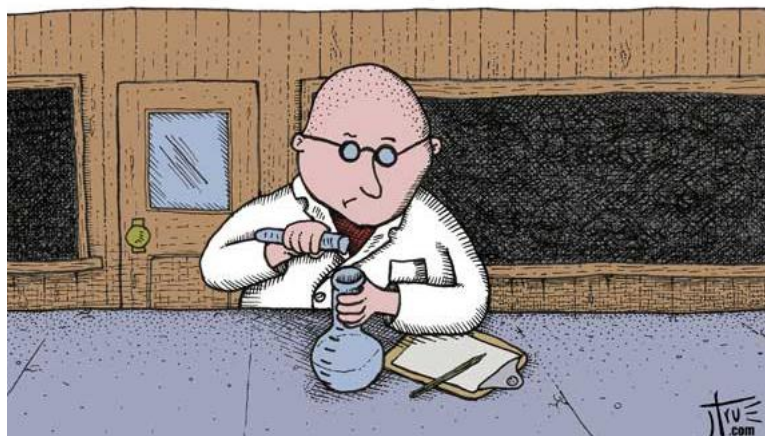
9. Repeat Steps 2-8 using 100.0 mL of 0.50 M hydrochloric acid, HCl, instead of water. **Caution:** Handle the HCl solution and NaOH solid with care as both damage skin and clothing!

Reaction C (performed by the entire class)

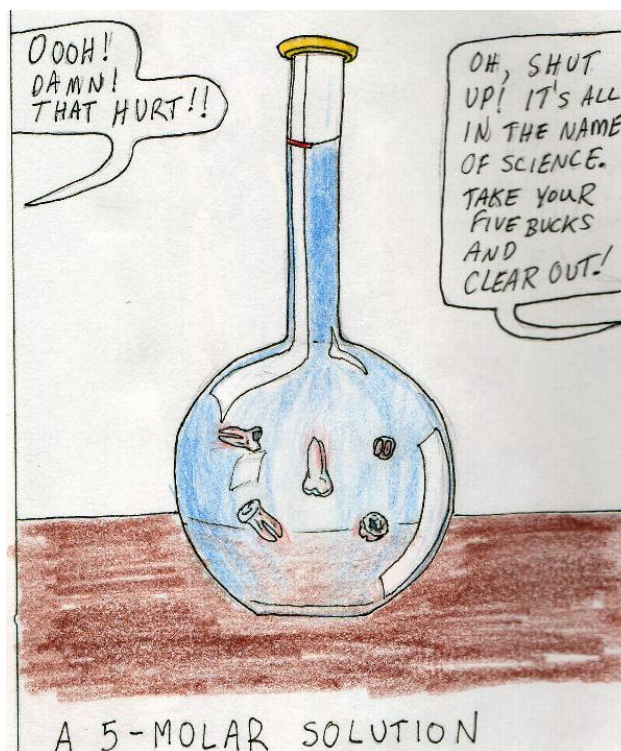
10. Repeat Steps 2-8, initially measuring out 50.0 mL of 1.0 M HCl (instead of water) into the Styrofoam calorimeter. In Step 5, instead of solid NaOH, measure 50.0 mL of 1.0 M NaOH solution into a graduated cylinder. After t_1 has been determined for the 1.0 M HCl, add the 1.0 M NaOH solution to the Styrofoam cup. **Caution:** Handle the HCl and NaOH solutions with care.

On the lighter side...

Hue discovers the element of surprise...



"I can't go out with you again. There just isn't any chemistry between us."



Lab 9 Report Sheet
Hess's Law

Name _____

Team No. _____ Date _____ Section _____

*Experimental Results***Table 1.** Team and class data for Reaction A: $\text{NaOH (s)} \rightarrow \text{NaOH (aq)}$

Mass NaOH (g)	initial t_{water} ($^{\circ}\text{C}$)	final t_{soln} ($^{\circ}\text{C}$)	Δt_{soln} ($^{\circ}\text{C}$)	$\Delta H_{\text{rxn A}}$ (kJ/g NaOH)	$\Delta H_{\text{rxn A}}$ (kJ/mol NaOH)
				Class Average	
				Standard Deviation	
				Accepted Value	
				% Error	

Table 2. Team and class data for Reaction B: $\text{NaOH (s)} + \text{HCl (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)}$

Mass NaOH (g)	initial t_{HCl} ($^{\circ}\text{C}$)	final t_{soln} ($^{\circ}\text{C}$)	Δt_{soln} ($^{\circ}\text{C}$)	$\Delta H_{\text{rxn B}}$ (kJ/g NaOH)	$\Delta H_{\text{rxn B}}$ (kJ/mol NaOH)
				Class Average	
				Standard Deviation	
				Accepted Value	
				% Error	

- 2. Analysis of the class data:** Use *Excel* to determine the average ΔH_{rxn} and standard deviation for each of the three reactions, A, B and C and enter the results in the appropriate table in the results section.
- 3. Demonstration of Hess's Law:** Examine the chemical equations for the three reactions, and determine the relationship between them. Demonstrate this relationship by combining the equations in a manner consistent with Hess's Law by figuring out a way to add equations A and B so they can be added to produce equation C. Show your work below and include the experimentally determined value for the ΔH_{rxn} in each chemical equation.

Demonstration of Hess's Law using the molecular equations:

- 4.** Write the net ionic equation for each of the three reactions A, B, and C, and then again demonstrate the relationship between the three ionic equations by combining them in a manner consistent with Hess's Law. Include the ΔH_{rxn} in each chemical equation.

Net ionic Equation for Reaction A: _____

Net ionic Equation for Reaction B: _____

Net ionic Equation for Reaction C: _____

Demonstration of Hess's Law using the net ionic equations:

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5. Using the class average for the ΔH_{rxn} for each reaction, compare the experimental value for the ΔH_{rxn} for reaction C with the value obtained by applying Hess's Law. Show your work below and discuss the likely reason(s) for any differences.
6. Use the standard enthalpies of formation, ΔH_f° , found in [Appendix B](#) in your text to calculate the "true" (i.e. accepted) ΔH_{rxn} value for all three reactions. Show your work below and enter the ΔH_{rxn} values in the tables in the results section. Note that for aqueous ionic compounds, you must look up each ion separately. i.e. for NaCl(aq), look up Na⁺(aq) and Cl⁻(aq). Circle your answers.
- a. Calculation of ΔH_{rxn} from ΔH_f° values for reaction A:
- b. Calculation of ΔH_{rxn} from ΔH_f° values for reaction B:
- c. Calculation of ΔH_{rxn} from ΔH_f° values for reaction C:

-
7. Calculate the % error for each of the three reactions and enter the result in the tables in the results section. You do not need to show your work.
8. Comment on both the precision (use the standard deviation computed above as support) and the accuracy (% error) of the results. Use specific numerical data from the experiment to support your response. Things to consider in your analysis include but are not limited to the following: What are the most likely sources of errors? What measured value(s) are affected by these errors and how does this affect the calculated values of ΔH_{rxn} for each reaction?
9. a. Which reaction has the largest % Error for the ΔH_{rxn} ? Explain why. Hint: think about the physical states of the reactants in each of the three reactions. Use specific numerical data from the experiment to support your response.

- b. Which reaction has the *smallest* % Error for the ΔH_{rxn} ? *Explain why. Hint: think about the physical states of the reactants in each of the three reactions. Use specific numerical data from the experiment to support your response.*
10. a. Using the theoretical values of ΔH_{rxn} calculated in [question #6](#), how does the sum of the ΔH_{rxn} for reactions A and C compare to the ΔH_{rxn} for reaction B? Use specific numerical data to support your response.
- b. Use your knowledge of chemistry to explain the relationship between the ΔH_{rxn} for reactions A and C vs. that of reaction B.
11. In this experiment we neglected any heat that may have been lost by the system to the calorimeter, *i.e.*, we assumed that $q_{\text{calorimeter}} = 0$. Which measured value is affected by this? Would this tend to make your experimental results (*i.e.* the ΔH_{rxn}) higher or lower than the theoretical values? Explain your reasoning carefully.

Lab 9 Prelab Questions

Name _____

Hess's Law

Team No. _____ Date _____ Section _____

Instructions: Complete the following questions and hand in at the start of your lab period or when instructed by your instructor. Show your work with units and correct significant figures for all questions that involve a calculation.

1. Calculate the heat of solution (ΔH_{soln}) for sodium hydroxide in *kJ/mol NaOH* from the following data:
When 4.160 g of NaOH was added to a coffee cup calorimeter containing 100.0 mL of water at 19.30 °C the highest temperature attained after completely dissolving the NaOH was 29.70 °C. Assume the heat capacity of water = 4.184 J/g °C and the Density of water = 1.000 g/mL. Molar masses in g/mol: Na = 22.990; O = 15.999; H = 1.0079 Circle your answer!

2. The calculation in question 1, above, like today's lab exercise, assumes the heat capacity of the coffee cup calorimeter is negligible. However, the coffee cup calorimeter, the glass thermometer and the stirring rod do absorb heat generated when the sodium hydroxide dissolves in the water inside the coffee cup. How would you expect the value for the heat of solution calculated for sodium hydroxide to be impacted by ignoring the heat capacity of the calorimeter? Hints: What measured value is affected? Is it artificially high or low? How does this impact the calculated value of the ΔH_{soln} for sodium hydroxide?

3. Today's lab exercise uses a coffee cup calorimeter to measure the heat of reaction (enthalpy of reaction) of several different reactions. The coffee cup *minimizes*, but does not eliminate, heat transfer with the surroundings. Small, but significant, amounts of heat may be either gained from or lost to the environment. If the temperature of the room were 22.00 °C on the day of the experiment in question 1, above, how would you expect the value for the heat of solution calculated for sodium hydroxide to be impacted by the poor insulating qualities of the coffee cup calorimeter? Hints: What measured value is affected? Is it artificially high or low? How does this impact the calculated value of the ΔH_{soln} for sodium hydroxide?