

# Hess's law

Law



**ASSIGN  
BUSTER**

## Enthalpy and Hess's Law Lab

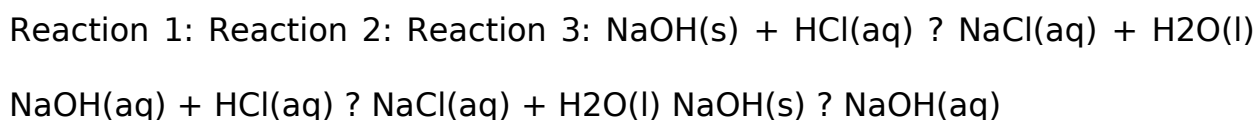
### Introduction

Every chemical reaction is accompanied by a change in heat.

Thermochemical reactions include the heat of reaction as part of the equation. The heat released or absorbed in a reaction at constant pressure is the Enthalpy change for the reaction. The enthalpy change for each reaction is unique to that reaction. Many values for  $H_{rxn}$  were experimentally determined, and many were calculated using Hess's law. This lab will illustrate the principle of Hess's law: if a reaction can be carried out in a series of steps, the sum of the enthalpies for each step equals the enthalpy change for the overall reaction.

$$H_{rxn} = H_{step1} + H_{step2}.$$

The three reactions that we will be using are as follows:



Since reaction 1 can be obtained by adding reactions 2 and 3, the  $H_{rxn1}$  should equal  $H_{rxn2} + H_{rxn3}$ .

At constant pressure  $H_{rxn} = q_p$ . We cannot directly measure

$H_{rxn}$  or  $q_p$ , but we can measure the change in temperature for a solution, and using the specific heat of the solution, and the grams of solution, we can find  $q_p$  using the following equation:

Equation 1:  $q = (\text{grams of solution}) \times (\text{specific heat of solution}) \times \Delta T$  The heat that is released by the reaction will be absorbed by both the surroundings; in this case, the water in the solution and the calorimeter itself:

Equation 2:  $q_{\text{rxn}} = -(q_{\text{solution}} + q_{\text{calorimeter}})$

Since every group's calorimeter is slightly different, the heat capacity for the calorimeter will need to be experimentally determined before it is used.

Pre-Lab Questions 1. 2. 3. 4. 5.

Define  $H_{\text{rxn}}$ . Define specific heat and heat capacity. How are these two terms different? The specific heat of a solution is  $4.18 \text{ J/gK}$  and its density is  $1.02 \text{ g/mL}$ . The solution is formed by combining  $25.0 \text{ mL}$  of solution A with  $25.0 \text{ mL}$  of solution B, with each solution initially at  $21.4 \text{ C}$ . The final temperature of the combined solutions is  $25.3 \text{ C}$ . Calculate the heat of reaction,  $q_{\text{rxn}}$ , assuming no heat loss to the calorimeter. If the calorimeter in the reaction above has a Heat Capacity of  $8.20 \text{ J/C}$ , recalculate the  $q_{\text{rxn}}$ , taking the heat loss to the calorimeter into account. If the reaction above between solutions A and B goes as follows:  $\text{A(aq)} + \text{B(aq)} \rightarrow \text{AB(aq)}$ , and the molarity of A in solution A is  $0.60 \text{ M}$ , and the molarity of B in solution B is  $0.60 \text{ M}$ , what is the enthalpy of reaction ( $H_{\text{rxn}}$ ), for the formation of 1 mole of AB in solution. Express  $H_{\text{rxn}}$  in  $\text{kJ/mol}$

Materials NaOH(s) HCl(aq,  $1.0 \text{ M}$ )  
 250mL Beaker Digital Thermometer Graduated Cylinder Balance Distilled Water NaOH(aq,  $1.0 \text{ M}$ ) Calorimeter w/ Lid Magnetic Stirrer and Stir Bar

Procedure

Part 1: Determining the heat capacity of the Calorimeter 1. 2. 3. 4. 5. 6. 7. 8.

Set up a calorimeter of two nested cups with a cover inside a beaker. Measure 50.0 mL of room temp distilled water into the calorimeter. Place the calorimeter on a magnetic stirrer and add a stir bar, set the stir bar to stir slowly. (Alternatively, gently stir the solution with the thermometer. ) Record the temperature of the water in the calorimeter. Heat or obtain roughly 75 mL of 70°C water. Measure out 50.0 mL of this water using a graduated cylinder. Record the temperature of the hot water, and pour the hot water into the room temp water in the calorimeter. Cover the calorimeter and insert the thermometer. Stir and record the temperature every 20 seconds for three minutes. 10. Empty and dry the inside of the calorimeter, thermometer, and stir bar.

Part 2: Determining the heats of Reaction  
Reaction 1:  $\text{NaOH(s)} \rightarrow \text{NaOH(aq)}$

1. Weigh out about 2.0 g of NaOH(s). Record the actual weight.
2. Add 100.0 mL of room temperature water to the calorimeter.
3. Start the stir bar and record the temperature of the water.
4. Add the NaOH(s), and record the temperature once every 20 seconds, until it stops changing.
5. Dump out the NaOH(aq) into the sink, rinse out, and dry the calorimeter, thermometer, and stir bar. Reaction 2:  $\text{NaOH(aq)} + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$
6. Combine 50.0 mL of 1.0 M NaOH and 50.0 mL 1.0 M HCl, in the calorimeter.
7. Record the temperature once every 20 seconds, until it stops changing.

8. Dump out the solution into the sink, rinse and dry the calorimeter, thermometer, and stir bar. Reaction 3:  $\text{NaOH(s)} + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$ .
9. Measure 50.0 mL of Distilled water into the calorimeter.
10. Add 2.0 g of NaOH(s) (Record the actual weight) and 50.0 mL of 1.0 M HCl to your calorimeter.
11. Record the temperature once every 20 seconds, until it stops changing.
12. Dump out the solution into the sink, rinse and dry the calorimeter, thermometer, and stir bar. Data Create a data table to hold all of your data for the 2 parts. Make sure you have a space for the initial and final temperatures, as well as the mass or volume for each reactant.

#### Calculations 1.

Calculate the Heat Capacity of the Calorimeter.

- a. When equal volumes of hot and cold water are combined, if there is no heat loss the new temperature should be the average of the two starting temperatures. In actual practice, the new temperature will be slightly less than the average because of heat loss to the calorimeter assembly. Additionally, when two solutions are mixed the thermometer cannot instantaneously record the temperature of the combined solutions. The solutions require some time to become completely mixed, and the thermometer needs time to come to temperature equilibrium with the solution. The theoretical temperature that the mixture would have if the process occurred instantaneously can be

found from a graph. Plot the data with temperature on the vertical axis versus time on the horizontal axis. The first few points may be erratic because of incomplete mixing and lack of temperature equilibrium with the thermometer. The points that follow should occur in a straight line as the temperature slowly drops while heat is lost to the calorimeter and to the surroundings. Draw a straight line through these points, and extend it back to find the temperature at time zero, the theoretical instantaneous temperature of mixing,  $T_{mix}$ . See Figure 1.

- b. Calculate the average temperature of the hot and cold water,  $T_{avg}$ .
- c. The difference between the average temperature,  $T_{avg}$ , and the instantaneous temperature,  $T_{mix}$ , is due to the fact that some heat was lost by the water and absorbed by the calorimeter. Calculate  $q_{water}$ , the heat lost by the water:  $q_{water} = (\text{grams of water}) \times (\text{specific heat of water}) \times (T_{mix} - T_{avg})$  where  $q_{water}$  = heat lost by water and the specific heat of water is  $4.18 \text{ J}/(\text{g}\cdot^{\circ}\text{C})$ . The heat absorbed by the calorimeter,  $q_{calorimeter}$ , will be equal to that lost by the water but opposite in sign.  $q_{calorimeter} = -q_{water}$
- d. Calculate the heat capacity of the calorimeter,  $C_{calorimeter}$ , which is the heat that the calorimeter absorbs each time the temperature of the solution changes  $1^{\circ}\text{C}$ :  $C_{calorimeter} = q_{water} / (T_{mix} - T_{initial})$  where  $T_{initial}$  is the initial temperature of the cool water. Calculate  $\Delta H$  for Each Reaction.
- e. Graph the temperature versus time for each of the three reactions tested. Extrapolate the line back to find the theoretical instantaneous mixing temperature,  $T_{mix}$ , as you did above.

- f. Calculate the amount of heat evolved in each reaction,  $q_{rxn}$ , by assuming that all of the heat is absorbed by the solutions and the calorimeter:  $q_{rxn} = - [\text{heat absorbed by solution} + \text{heat absorbed by calorimeter}]$   $q_{rxn} = - [( \text{grams of solution} \times \text{specific heat of solution} \times (T_{\text{solution}} - T_{\text{initial}}) ) + (C_{\text{calorimeter}} \times (T_{\text{solution}} - T_{\text{initial}}))]$  where  $T_{\text{solution}} = (T_{\text{mix}} - T_{\text{initial}})$  for each reaction mixture. Assume that the density of the solutions is 1.03 g/mL, and that the specific heat of the solutions is the same as that of water, 4.18 J/(g·°C).
- g. 3. Calculate the value of the enthalpy change,  $\Delta H$ , in terms of kJ/mole for each of the reactions. Verify Hess's Law.

### Conclusion

Write net ionic equations for the three reactions involved. Show how you must arrange the first two equations to algebraically find the third. Calculate the value of  $\Delta H$  for the third reaction from your  $\Delta H$  values for the first two reactions using Hess's law. Find the percent difference between the calculated and measured values.