
Experiment #8: Calorimetry

When a substance undergoes a change in temperature, the quantity (q) of heat lost or gained can be calculated using the mass (m), specific heat (s), and change in temperature ($\Delta T = T_{\text{final}} - T_{\text{initial}}$):

$$q = m s \Delta T \quad (1)$$

The specific heat (s) is the amount of heat required to change the temperature of one gram of a particular substance by one degree Celsius (or 1 K). For water, the specific heat is defined as 1 cal / g °C or measured to be 4.184 J / g °C.

Heat exchange experiments are conducted in a calorimeter which is an insulated container. The calorimeter is constructed such that ideally there will be no heat exchange between the contents of the calorimeter and the surrounding environment. However, heat exchanges can occur between the components within the calorimeter. This can be expressed in the following way:

$$q_{\text{system}} = q_{\text{gain}} + q_{\text{loss}} \quad (2)$$

If the system does not exchange heat with the surroundings, then the $q_{\text{system}} = 0$, and

$$q_{\text{gain}} = -q_{\text{loss}} \quad (3)$$

When a hot substance comes in contact with a cold substance within a calorimeter, heat will spontaneously pass from the hot to the cold substance until they achieve a thermal equilibrium (the same constant temperature). The quantity of heat lost and gained is equal in magnitude, but opposite in sign. For example, if a piece of hot metal is placed in cold water within a calorimeter, the following thermodynamic equation can be used:

$$q_{\text{water}} = -q_{\text{metal}} \quad (4)$$

When a chemical reaction occurs in a water solution, the situation is similar to when a hot piece of metal is placed in water. When a reaction occurs, there is a change in the potential energy of the chemicals, resulting in a heat exchange in the calorimeter. As in the hot metal experiment, the heat flow of the reaction is equal in magnitude but opposite in sign to that of the solution as shown by the following equation:

$$q_{\text{reaction}} = -q_{\text{solution}} = -(m s \Delta T)_{\text{solution}} \quad (5)$$

By measuring the mass of the solution (the water plus the solute) and by observing the temperature change that the solution undergoes, we can calculate q_{solution} and therefore q_{reaction} . Because the experiment is performed under constant pressure conditions, the heat flow of the reaction is also equal to the enthalpy change, ΔH , for the reaction. Since the mass of reactants and products can vary for different experiments, the q_{reaction} will also vary.

Therefore, ΔH is generally reported as the heat of reaction per a fixed amount of one reactant or product, and we will calculate ΔH per gram or per mole of this substance.

$$\Delta H_{\text{rxn}} = \frac{q_{\text{rxn}}}{\text{mass}_{\text{substance}}} \quad \text{or} \quad \Delta H_{\text{rxn}} = \frac{q_{\text{rxn}}}{\text{mole}_{\text{substance}}} \quad (6)$$

If the temperature of the water goes up, heat has been given off by the reaction, so that reaction is exothermic, and ΔH of the reaction is negative. If the temperature of the water goes down, heat is absorbed by the reaction, therefore the reaction is endothermic, and ΔH for the reaction is positive.

Procedure

Construct a calorimeter, where two polystyrene coffee cups with a plastic cover on top are nested together. Weigh and record the mass of the clean, dry calorimeter. Place the calorimeter into a 400 mL glass beaker for stability during the experiment.

Measure 25.0 mL of 1.00 M H_2SO_4 and 75.0 mL of 1.50 M NaOH into separate 100 mL beakers. Record the temperature of each of the solutions to the nearest 0.01 $^\circ\text{C}$ using the same thermometer. Be sure to wash and dry the thermometer after putting it into each solution to avoid initiating the reaction by prematurely mixing the solutions.

Immediately and simultaneously pour both solutions into the calorimeter, replace the lid, insert the thermometer, and stir vigorously. Record the final temperature to the nearest 0.01 $^\circ\text{C}$. Carefully remove the thermometer from the calorimeter, then weigh and record the mass of the calorimeter containing the reaction mixture.

Dispose of any waste down the drain, rinse the cups, lids, and thermometers with DI water, and dry them thoroughly before returning them to where you originally got them.

Name: _____

Section: _____

Data and Calculations

Mass of clean, dry calorimeter: _____ g

Mass of calorimeter with solution after conclusion of reaction: _____ g

Final mass of reaction solution: _____ g

Initial temperature of H₂SO₄ solution: _____ °C

Initial temperature of NaOH solution: _____ °C

Average initial temperature of starting solutions: _____ °C

Final temperature of mixture: _____ °C

Change in temperature of solution (ΔT): _____ °C

1. Write the balanced chemical equation for your acid/base neutralization reaction.

2. Determine the theoretical yield of water (in grams) of the reaction that you carried out.

3. What is the limiting reactant of your reaction?

4. Calculate the mass of each reactant that is theoretically left over at the end of your reaction.

_____ g of NaOH left over _____ g of H₂SO₄ left over

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5. Assuming that the specific heat capacity of the solution that you used is equal to $3.70 \text{ J / g } ^\circ\text{C}$, calculate the heat (q_{solution}) in kJ that was absorbed by the contents of the coffee cup during the reaction.

6. Calculate ΔH of your reaction using the units specified:

_____ kJ / g H_2O formed _____ kJ / mol H_2O formed

7. Rewrite your balanced chemical equation and include the value of ΔH beside it:

_____ \rightarrow _____ $\Delta H =$ _____ kJ

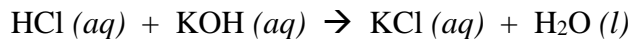
8. Use the ΔH_f° values given to calculate the theoretical value of ΔH for your reaction.

Substance	ΔH_f° (kJ / mol)
$\text{H}_2\text{SO}_4 (aq)$	-909.3
$\text{NaOH} (aq)$	-470.1
$\text{H}_2\text{O} (l)$	-285.8
$\text{Na}_2\text{SO}_4 (aq)$	-1387.1

9. Using your results from questions 7 & 8, calculate the percent error in your determination of ΔH .

Pre-lab Assignment

A student carries out the following reaction in lab by mixing 50.0 mL of a 1.00 M solution of hydrochloric acid with 50.0 mL of a 1.00 M solution of potassium hydroxide:



The average initial temperature of the two solutions was 25.00 °C. After mixing, the final temperature of the solution was 31.43 °C. The mass of the resulting solution was 100.2 g.

1. Determine the theoretical yield of KCl (in moles) of the student's reaction.
2. Assuming that the specific heat capacity of the solutions that the student used are equal to the specific heat capacity of pure water, calculate the heat (q_{solution}) in kJ that was absorbed by the contents of the coffee cup from the reaction.
3. Calculate the value of ΔH_{rxn} in the units kJ / mol KCl.
4. For this reaction, the theoretical value of ΔH_{rxn} is $-55.8 \text{ kJ / mol KCl}$. What was the student's percent error when determining ΔH_{rxn} ?