Measuring the Enthalpy of Reaction with Calorimetry

GHP

Introduction: Calorimetry is a measurement process that takes advantage of the Law of Conservation of Energy to measure the change in energy in a physical or chemical process by observing the temperature change of water in the calorimeter. Since we know that all the heat energy from the process MUST go into or come out of the water in the calorimeter.

In our previous experiment we used this concept to measure the specific heat of an unknown metal using the specific heat of water (in the calorimeter) as well we the change in temperature of the water. All this was entered into the equation below, allowing us to calculate C_m, or the specific heat of the metal:

$$-m_m C_m \Delta T_m = m_w C_w \Delta T_w$$

Above, the changes in temperature (ΔT, blue) were measured with a temperature probe, the masses (m_m and m_w, green) were measured with a balance, and the specific heat of water (C_w, purple) was looked up from a data table. This allows us to then calculate the missing value, Cm.

In this experiment we will use the same concept, but will use the results to calculate the enthalpy, or ΔH of a reaction instead. This reaction will be performed in a calorimeter with a known mass of water, causing a certain change in temperature we can again measure with a temperature probe:

$$q_{rxn} \simeq q_w = m_w C_w \Delta T_w$$

Where q_{rxn} is the total amount of heat absorbed or released from our reaction. TO turn this into an enthalpy we need to relate it to the amount of chemical used in the process:

$$\Delta H_{rxn} = \frac{-q_{rxn}}{mols\ reactant}$$

The added negative sign compensates for the differences in perspective on energy movement. The increase of energy in a calorimeter represents a loss of energy in the reaction, and that difference in direction is denoted by the sign of ΔH.

Procedure: You will need to setup a calorimeter like in the previous experiment: 250 mL beaker nested inside a styrofoam jacket with a styrofoam lid. You will need a temperature probe connected to a Chromebook, with the Vernier Graphical Analysis app open and ready to go with the bluetooth probe connected. See Mr. Kiefer if you need help with this setup.

On the following pages you will have 2 different options of experiments to choose from:

Enthalpy of Neutralization

This records the ΔH_{rxn} of a neutralization reaction. Since neutralization reactions happen in solution we can easily do these in a water filled calorimeter without special apparatus like a bomb calorimeter.

You will mix the acid and base in the calorimeter, causing an exothermic reaction to occur. The temperature of the water in the calorimeter will increase, and the size of ΔT must correspond to the amount of heat released from your reaction.

Enthalpy of Solution

Here you will measure the enthalpy of a solute while it dissolves in solution. While not a chemical reaction, the act of dissolving breaks and forms new connections between atoms and polyatomic ions, and therefore absorbs and releases heat energy.

Like the Enthalpy of Neutralization process, this all happens in the water of the calorimeter, so no special equipment needed. We will simply dissolve a fixed amount of one of many solutes in water, and then calculate the heat absorbed or released from the ΔT of the water in the calorimeter.

Enthalpy of Neutralization

Enthalpy of Neutralization (ΔH_{neu}) is the heat energy absorbed or released during the neutralization of an acid with a base. These reactions are typically exothermic, and in some cases *very* exothermic.

In our experiment you will react **citric acid** ($C_6H_8O_7$) with sodium hydroxide as your base. You will need to record the ΔH_{rxn} twice, getting an average of the two results.

To design your experiment draw on your experiences with calorimetry, and a reminder of the needed calculations to determine the ΔH of a reaction:

$$q_{rxn} \simeq q_w = m_w C_w \Delta T_w$$

$$\Delta H_{rxn} = \frac{-q_{rxn}}{n}$$

Where n is the number of mols reactant that generated or absorbed the heat in the reaction.

You will need to collect data from your calorimeter in order to first calculate q_{rxn}. This data should be organized into a data table that can hold information from both of your trials.

You will need to use data from the acid and base solutions that you used in order to calculate the number of mols of reactant used.

For your trials you will use approximately 50 mL of the base solution, NaOH. This chemical will be in excess so that your citric acid will limit the reaction and determine how much heat is released. You will add the solid citric acid directly to your calorimeter.

The 'water' in your calorimeter will be the mass of the sodium hydroxide solution. While the specific heat of a solutions is not the same as that of water, they should be close. We can still assume the specific heat of the solution in our calorimeter is about 4.184 J/g°C. If you want to look it up, the specific heat of a 0.5 M solution of sodium hydroxide is likely easy data to find, and it would mean more accurate results.

You should use between 2-4 g of the citric acid. How much you actually use of the acid is not important, but you do need to know that exact mass to use for your value of n in the equation above.

Enthalpy of Solution

This process will involve dissolving a certain quantity of a solute, either calcium chloride, or ammonium chloride, into the water of your calorimeter. As mentioned above, as well as in Unit 6 Section 1, dissolving a solute into a solution means breaking the ionic bond holding the cation and anion together, and then creating new attractive forces between the solute ions and water in the solution. Since the amount of energy for each of these steps is not the same, there is an enthalpy of solution (ΔH_{sol}).

You will be measuring the enthalpy of solution of either calcium chloride or ammonium nitrate. You will use a specific mass of either solute and add them while stirring to your calorimeter to create a change in temperature you can use to determine the enthalpy of solution.

To design your experiment draw on your experiences with calorimetry, and a reminder of the needed calculations to determine the ΔH of a reaction:

$$q_{sol} \simeq q_w = m_w C_w \Delta T_w$$

$$\Delta H_{sol} = \frac{-q_{sol}}{n}$$

Where n is the number of mols reactant that generated or absorbed the heat in the reaction.

You will need to collect data from your calorimeter in order to first calculate q_{rxn}. This data should be organized into a data table that can hold information from both of your trials.

You will need an accurate mass of your solute used in order to calculate the number of mols of solute used.

You should use approximately 5 g of either the calcium chloride or the ammonium nitrate. Many procedure for similar experiments suggest using 50 mL of water, but that would be a very small amount in our larger calorimeters. Start trials with 50 mL of water, but increase the volume to 75 mL or 100 mL if temperature reading are erratic.

What to turn in:

1. **A shared Google Doc with the names of all group members at the top.**
2. **A title for your experiment that is specific to what you did in your experiment**. There are a few choices!
3. **A data table containing all of the data recorded in your experiment**. This table should have headings, units, and easily organize data from both of your trials.
4. **ALL of the calculations that you did to determine your value for ΔH**. You can type these out using the equation tool (time consuming) or do the work on paper and embed a picture of your work. Whichever you choose, make sure the calcs are clear and easy to follow.
5. **A comparison of your value of ΔH to the expected values available online.** Search for what you measured and the chemical(s) used and you should be able to find data easily. Check with me if you are uncertain.
6. **A calculation showing the percent error in your experiment.** If you do not know how to calculate percent error, again search for an equation online. It is a simple formula that measures how fara apart your data is from the expected result.
7. **Explanations for why your value of ΔH might vary from the expected value**. Here are some ideas to consider. Pick at least 2:
    a. Heat loss from the calorimeter - be specific on how this would impact your answer. Does your data vary in a way that matches this impact?
    b. Heat absorbed by the glass and styrofoam of the calorimeter. This heat energy would not be accounted for with the thermometer, and many procedures available online offer a method for correcting for this. This correction is often described as the heat capacity of the calorimeter, and is usually added to q_{rxn} before calculating $\Delta H_{rxn.}$
    c. The value used for the specific heat of water. We use the number 4.184 j/g °C as our specific heat in the calculations, but that is for pure water. Look up the specific heats of the actual solutions used in our experiment. How would those differences impact our results if a more accurate C was used?