

CHEM 231

Experiment 7

Introduction to Calorimetry

Calorimetry is a method of determining heats of reaction. The heat generated by a reaction causes the temperature of a solution to rise. By measuring this temperature change, and relating it to the heat capacity, one can find the amount of heat generated in a reaction.

A “calorimeter” is an insulated device in which the reaction is carried out. In this experiment, the calorimeter is simply a pair of nested Styrofoam cups with loose fitting lids. You will use two such calorimeters, one designated as the “master.” Ideally, the calorimeter itself does not absorb any of the reaction heat. In practice, however, one needs to correct the observed temperature change for the fact that some of the heat is lost in this manner.

In this experiment, you will be studying two reactions by calorimetry. One is the reaction of a strong acid with a strong base and the other is the oxidation of magnesium metal by a strong acid.

The following directions are intended to walk you through the concepts of calorimetry. The calculations are imbedded in the directions and you should do these and understand them as you are doing the experiment, rather than as a post-lab exercise. Do not continue until you understand the calculations.

Procedure

Step 1: Finding the heat capacity of the calorimeter

Begin heating some water (a bit more than 50 mL) in a beaker with a Bunsen burner. Meantime, weigh both calorimeters (without lids) and add 50 mL of distilled water to one (the “master” calorimeter). Adjust the amount of water with a disposable dropper pipette while the calorimeter is on the balance to exactly 50 grams of added water. Set it up with the thermometer in it and the lid closed.

When the water on the burner reaches about 60°C, measure 50 mL of the hot water into calorimeter 2, adjust the amount on the balance and set it up just as you did the first. Wait approximately 10 minutes for the temperatures of both to stabilize.

Meantime, ask yourself: “What do I expect the temperature to be when I mix these two identical volumes of water?” Without doing any complicated calculations, it should be obvious that in the ideal situation, the water should equilibrate to the average of the two temperatures. It will not quite do that because of the fact that the calorimeter absorbs some heat, and the temperature will actually be lower than what we expect. The object here is to determine how much.

While waiting, find the average of the two temperatures so that you know what to expect. (For example, if the cold water is at 25°C and the hot water is at 55°C, average is 40°C and we expect to see something slightly less than this.) By the time you are done thinking this over, it might be time to mix the solutions. Before doing so, write in the space below (and in your notebook, of course) the temperatures of the two solutions and their average:

Cold water _____ Hot water _____ Average _____

Now, carefully and quickly pour the entire contents of the hot calorimeter into the cold one (the master calorimeter). Immediately replace the lid and the thermometer. Wait about 10 minutes for the temperature to stabilize. Record the final temperature and figure the difference between this and the expected average. For example, if the expected average is 40°C and you measure 35°C, the difference is -5°C. Record the results below and in your notebook. If the result is not negative, something is wrong.

Final temperature _____

Change in temperature ($T_{\text{final}} - T_{\text{average}}$) or $\Delta T =$ _____ (must be negative)

The water must have lost a certain amount of heat to account for the difference. This heat must have gone into the calorimeter. The amount can be calculated using the specific heat of water which is 4.18 J/K-gm; multiplying by the mass of 100 mL (100 g) of water, we find its heat capacity to be 418 J/K. The heat lost by the water is, therefore, $\Delta T \times (418 \text{ J/deg})$, a negative number. The heat gained by the calorimeter is simply the opposite of this, a positive number. For example, with a temperature difference of -5°C, the heat absorbed by the calorimeter will be 2090 J. Compute this, and write it below and in your notebook:

Heat absorbed by calorimeter _____ J.

We can now calculate the calorimeter heat capacity. As the hot water was added, the calorimeter temperature increased from that of the cold water to the final temperature. For example, if the cold water was originally at 25°C and the final temperature you recorded above is 35°C, then the change in temperature of the calorimeter was $\Delta T = +10^\circ\text{C}$. The heat capacity of the calorimeter is simply the heat absorbed, recorded above, divided by ΔT (of calorimeter). For example, if the heat recorded above is 2090 J, and the ΔT is 10°C, then the heat capacity of the calorimeter is 209 J/deg. Record these below and in your notebook:

ΔT (of calorimeter) = _____

$C_{\text{cal}} = \text{Heat absorbed} / \Delta T =$ _____ J/deg.

The heat capacity to be used in the subsequent sections is simply the sum of the heat capacity of 100 mL of water and that of the calorimeter, or 418 J/deg + C_{cal} . Record this below and in your notebook:

$C = 418 \text{ J/K} + C_{\text{cal}} = \text{_____ J/deg.}$ (assumes that you are using 100 mL of water)

In your experimental setup, the ΔH of any reaction is calculated using the heat capacity you determined here and the change in temperature of the reaction mixture:

$$\Delta H = -C \times \Delta T$$

The minus sign arises from the fact that an increase in the heat content of the water and the calorimeter results from a decrease in the enthalpies of the reacting substances. It is simply that ΔH for an exothermic reaction is negative.

Step 2: The heat of reaction of a strong acid and a strong base

Set up both calorimeters. Add 50 mL of 1 M sodium hydroxide to one calorimeter and 50 mL of 1 M hydrochloric acid to the other. Use the balance to adjust the amounts to exactly 50 grams just as you did in the first part. Allow both to equilibrate for about 10 minutes. They should both be at about room temperature. Record their temperatures and the average.

After the temperatures are stable, pour the contents of calorimeter 2 into the master calorimeter. Close it and allow 10 minutes for the temperature to stabilize. Record the final temperature and compute ΔT as well as the heat of reaction. As discussed above, the heat of reaction is $\Delta H = -C \times \Delta T$.

The reaction is $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$. Find the number of moles of reactants from the concentrations (1 mole per Liter of each reactant) and the volumes used (50 mL of each). How many moles of water were produced?

Find the standard heat of reaction by dividing your measured ΔH by the number of moles of water. Report in kJ/mol.

Results:

$$\Delta T = \text{_____}$$

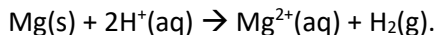
$$\Delta H = -C\Delta T = \text{_____} \text{ (Use } C \text{ determined from the first part.)}$$

$$\text{Moles } \text{H}^+ = \text{Moles } \text{OH}^- = \text{Moles } \text{H}_2\text{O} = (1.000 \text{ mol/L}) \times (0.0500 \text{ L}) = \text{_____}$$

$$\text{Heat of reaction per mole} = \Delta H / (\text{moles } \text{H}_2\text{O}) = \text{_____}$$

Step 3: The heat of reaction of magnesium with strong acid

In this part of the experiment, you will use only one calorimeter, the designated master calorimeter. The reaction is



Begin with 100 mL of 1 M HCl in the calorimeter. Adjust the amount using a balance to be exactly 100 g of solution in the calorimeter. Set it up and allow it to equilibrate for 10 minutes as before.

While waiting, precisely weigh a sample of magnesium on the analytical balance and record the mass. 0.5 grams is an appropriate amount. In this experiment, the magnesium is the limiting reactant and will determine the stoichiometry of the reaction. The acid will be present in excess.

Quickly add the solid magnesium to the calorimeter and close it. You should see an immediate reaction. Wait for the reaction to run and for the system to reach equilibrium. While you are waiting, calculate the number of moles of magnesium metal used.

When finished, calculate, in this order, the temperature change ΔT , the heat of reaction, $\Delta H = -C \times \Delta T$, using the total heat capacity from the first part of the experiment, and the heat of reaction **per mole**, given by $\Delta H/(\text{moles Mg})$.

Results:

$$\Delta T = \underline{\hspace{2cm}}$$

$$\Delta H = -C\Delta T = \underline{\hspace{2cm}} \text{ (Use C determined from the first part.)}$$

$$\text{Moles Mg} = \underline{\hspace{2cm}}$$

$$\text{Heat of reaction per mole} = \Delta H/(\text{moles Mg}) = \underline{\hspace{2cm}}$$