

Experiment 14 - Heats of Reactions

If a chemical reaction is carried out inside a calorimeter, the heat evolved or absorbed by the reaction can be determined. A calorimeter is an insulated container, and we assume that there is no heat exchange between the calorimeter and its surroundings. The reactions we will be performing in this experiment will occur in aqueous solution. The amount of heat absorbed or given off by the reactions can be determined indirectly by measuring the temperature change of the contents of the calorimeter in each reaction. We will use the relationship

$$q = Cm\Delta T$$

where q represents the amount of heat given off or absorbed by the solution, C is the heat capacity of the solution, m is the mass of the solution, and ΔT is the temperature change of the solution ($\Delta T = T_{\text{final}} - T_{\text{initial}}$). Since the solution is mostly water, the heat capacity will be close to that of water.

To determine the enthalpy of the reaction, we must realize that if the solution temperature increased, the reaction itself must have given off heat. Therefore, if q_{solution} is positive, then q_{reaction} must be negative, and vice versa. Additionally, we must realize that enthalpies of reaction are generally listed in units of kilojoules per mole. However, q_{reaction} is the amount of energy absorbed/released by the reaction for the specific amounts that we happened to use in our experiment. If we had used more or less of the reactants, a different amount of heat would be released. Therefore, in order to calculate ΔH_{rxn} for our reactions, we must divide the amount of energy released or absorbed (q_{reaction}) by the number of moles of the reactant used.

In this experiment, two reactions will each be conducted in a calorimeter. The data obtained will allow ΔH_{rxn} (enthalpy of reaction) values to be calculated. After the ΔH values have been calculated for the two observed reactions, the experimental information can be used to predict the ΔH value of a third reaction. The third reaction will not be actually observed, but its ΔH can be calculated, using the principles of Hess's Law.

Safety Precautions:

- Wear your safety goggles.
- Solutions of hydrochloric acid (HCl) are corrosive. If HCl splashes on you, rinse it off immediately with plenty of running water.

Waste Disposal:

- The waste from this experiment may be safely disposed of in the sink with plenty of running water.

Part 1

The first reaction to be studied is the reaction of solid magnesium with aqueous hydrochloric acid. The reaction is given below:



A measured amount of hydrochloric acid solution will be placed in a calorimeter. Magnesium ribbon will be added, and the heat produced by the reaction will be measured by observing the temperature change (ΔT) of the HCl solution.

Procedure – Part 1

1. Weigh a dry, empty calorimeter (two nested Styrofoam cups).
2. Measure out about 100 mL of 1 M HCl_(aq) in a graduated cylinder. Record the precise volume used. (If the temperature of the 1 M HCl is adjusted to about 15 °C before it is placed in the calorimeter, more accurate results may be obtained. The temperature may be adjusted by placing a flask containing the HCl into an ice/water bath for a while.)
3. Add the 1 M HCl solution to the calorimeter and weigh it again. Subtraction gives the mass of the HCl solution. The specific heat of 1 M HCl solution is 4.04 J/g•°C.
4. Allow the temperature of the acid and the calorimeter to equilibrate for a few minutes. While you are waiting, proceed with step 5.
5. Obtain a strip of magnesium ribbon that weighs between 0.2 and 0.3 grams. If the magnesium doesn't look shiny, clean it with steel wool before weighing. Weigh the magnesium strip to the nearest ± 0.001 or ± 0.0001 grams (use a digital balance that has doors), and record the mass.
6. Wind the magnesium ribbon around a pencil to give it a compact, spiral shape.
7. Set up a ring stand with an iron ring, and set the calorimeter under the ring. (This ring will be used to help support the thermometer.) Read and record the temperature of the HCl in the calorimeter. The temperature should be read to the nearest ± 0.1 °C. Leave the thermometer in the calorimeter for the rest of the experiment, supported inside the ring on the ring stand so that it doesn't fall over and break.
8. Drop the magnesium into the acid solution in the calorimeter. If necessary, use the thermometer to hold the ribbon under the surface of the acid.
9. Stir or swirl the calorimeter frequently. As the magnesium reacts, observe the temperature. Record the maximum temperature attained (again, to the nearest 0.1°C). This should be reached as the reaction is ending. Wait for a while after the magnesium has completely reacted to be sure you have observed the maximum temperature.
10. Pour the contents of the calorimeter into a large beaker. Neutralize the solution with sodium bicarbonate, if available, and discard the solution in the sink. Repeat the entire procedure so that your final answer can be a more accurate average result.

Calculations – Part 1

Use your experimental data to calculate the energy absorbed by the solution. Recall the equation $q = Cm\Delta T$, where m is the mass of the entire solution (the water and

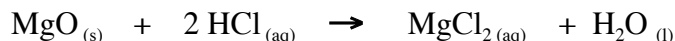
all of the reactants), C is the heat capacity of the solution ($4.04 \text{ J/g}\cdot^\circ\text{C}$), and ΔT is the temperature change of the solution ($\Delta T = T_{\text{final}} - T_{\text{initial}}$). Since we are assuming that heat is not lost to or gained from the surroundings in this experiment (we assume that the calorimeter is perfectly insulated), $q_{\text{solution}} = -q_{\text{reaction}}$. If heat is given off by the reaction, then heat is absorbed by the solution, and vice versa. You therefore only need to change the sign to obtain the heat involved for the reaction. However, this is the amount of energy involved for the specific amounts that you used. We would like to know the amount of energy involved per mole of the reactant. In this experiment, magnesium is the limiting reactant. In order to calculate the ΔH_{rxn} for this reaction, divide q_{reaction} by the number of moles of magnesium used in the experiment. This is ΔH for the reaction in kJ/mole. Do this calculation separately for each trial, and calculate the average value of ΔH_{rxn} and the percent difference between the two values.

Questions – Part 1

1. Does the reaction in part 1 release heat or absorb heat? How can you tell?
2. Is this reaction exothermic or endothermic? What is the sign of ΔH_{rxn} ?
3. Given that $\Delta H = H_{\text{products}} - H_{\text{reactants}}$, is the enthalpy of the reactants greater than or less than the enthalpy of the products? [Enthalpy is abbreviated as H , and is also called heat content. It is a form of potential (stored) energy, which can be converted to, or created from, kinetic energy (energy of motion). Kinetic energy of molecules is perceived by us as heat.]
4. Determine the limiting reactant in the reaction that you performed in part 1. Show your work.

Part 2

The second reaction that will be studied is the reaction of magnesium oxide with aqueous hydrochloric acid. The reaction is given below:



The procedure for studying this reaction is essentially the same as that used in part 1. Magnesium oxide (MgO) is a white powder that can be accurately weighed on a piece of smooth weighing paper. The MgO will be poured into an excess of HCl in a calorimeter and the temperature change will be measured. In this part of the experiment, 2 M HCl (specific heat = $3.89 \text{ J/g}\cdot^\circ\text{C}$) will be used instead of 1 M HCl .

The reaction can be performed on any scale, but it is most convenient to use an amount of MgO that will cause a temperature change of 20 to 25 $^\circ\text{C}$. With this as a guide, you can plan the quantities of reagents that should be used. In doing this, assume that the density of the solution is close to the density of water and assume that the reaction has a ΔH close to - 130 kJ/mole of MgO .

In estimating quantities of reagents to use,

1. Start by choosing some arbitrary but convenient volume of 2 M HCl solution to use.

2. Assuming the density of the solution is close to 1 g/mL, use $q = Cm\Delta T$ to determine the amount of energy needed to raise the temperature by the desired amount (choose a ΔT anywhere between 20 and 25°C).
3. Determine how much MgO must react in this solution to evolve the amount of energy you determined in #2. Assume ΔH is approximately -130 kJ/mol MgO.
4. Check the stoichiometry of your proposed amounts to see whether there will be enough HCl to actually react with all of the MgO. The HCl solution that will be used is 2 M HCl - that is, it contains two moles of HCl per liter of solution. MgO must be the limiting reagent in the reaction. In other words, all of the MgO should be used up in the reaction. If all the MgO is used up, the calorimeter will contain a clear solution at the end of the reaction. If MgO is not all reacted, the liquid in the calorimeter will be cloudy. The experiment will have to be repeated if a cloudy product is found.

Note: when you perform this part of the experiment, you don't need to weigh out the exact amount of MgO that you calculated. Any mass of MgO that is reasonably close to the calculated value will work, but you need to know precisely what mass you actually used. (MgO should be the limiting reagent.)

Procedure – Part 2

The experiment is performed by the same procedure that was used in Part 1. The calorimeter is weighed both empty and with 100 mL (or whatever volume you decided on) of 2M HCl present (record the precise volume of HCl used). The 2 M HCl is again allowed to equilibrate in the calorimeter, and the temperature is measured (± 0.1 °C). (As in Part 1, it is a good idea to adjust the temperature of the 2 M HCl solution to about 15 °C before putting it into the calorimeter.) The MgO is then accurately weighed on a piece of weighing paper, and is transferred directly from the paper to the calorimeter. Add the MgO powder slowly to the calorimeter, and then stir vigorously. (Note: it is best to use powdery MgO rather than chunky MgO; it is more likely to be pure.)

Again, be sure to observe the maximum temperature reached by the system (± 0.1 °C). When the reaction is complete, check to see that a clear solution remains in the calorimeter.

The reaction should be carried out a second time, so that the answer can be reported as an average value of two determinations. The amount of MgO can be changed, if desired, for the second run.

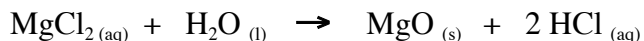
Calculations – Part 2

ΔH for the reaction of MgO with HCl can be calculated in kJ/mole of MgO. This is done as in Part 1. Again, calculate ΔH separately for each trial, and report the average and the percent difference of the two trials.

Hess's Law can then be used to find ΔH for the reaction:

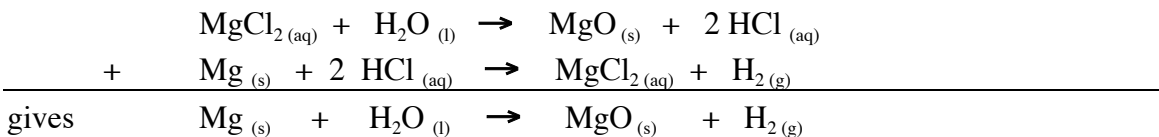


This is done in two steps. First find ΔH for the reaction:



Note that this reaction is the reverse of the reaction performed in Part 2. Its ΔH has the same magnitude as that reaction, but has the opposite sign.

The second step is to use Hess's Law to add two equations together. Adding:



ΔH for this last equation will be the sum of the ΔH values of the first two equations. Since ΔH values have already been found in this experiment for the first two equations, these values can be added to obtain ΔH for the final equation.

Questions – Part 2

5. Describe two sources of error in these experiments. Do not just list possible mistakes one could make; focus instead on systematic errors or assumptions made in the calculations that may not be totally accurate.
6. Calculate the ΔH of the reaction:

$$\text{Mg}_{(s)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{MgO}_{(s)} + \text{H}_{2(g)}$$
 Base this calculation on a combination of your results from parts 1 and 2. Show your work.
7. It has been established by experiment that the reaction:

$$\text{H}_{2(g)} + 1/2 \text{O}_{2(g)} \rightarrow \text{H}_2\text{O}_{(l)}$$
 has $\Delta H_{\text{rxn}} = -286 \text{ kJ/mole of H}_2\text{O formed}$. Use this information and your experimental data to find ΔH for the reaction:

$$\text{Mg}_{(s)} + 1/2 \text{O}_{2(g)} \rightarrow \text{MgO}_{(s)}$$
8. How many milliliters of 2.0 M HCl are needed to react with 3.0 grams of MgO? (Use the balanced equation for the reaction you performed in part 2.)