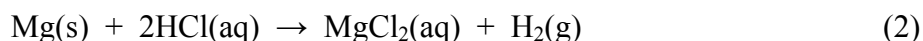


## Enthalpy Change of a Chemical Reaction with Magnesium – Background

To determine the enthalpy change for a reaction, the heat released or absorbed by the reaction can be measured in a container called a calorimeter. When an exothermic reaction occurs in a calorimeter, the heat which is released warms up the calorimeter, the reaction products, and the solvent. The reaction products and solvent (usually H<sub>2</sub>O) can be thought of as a single entity, the solution. The heat released by the reaction is given the symbol  $-q_{\text{rxn}}$ . The letter "q" represents heat, and here it is given a negative (-) sign because the heat is leaving the system (the reactants) and being absorbed by the surroundings (the container and the solution). The amount of heat energy released by the reaction ( $-q_{\text{rxn}}$ ) is equal to the sum of the heat absorbed by the calorimeter ( $+q_{\text{cal}}$ ) and the solution ( $+q_{\text{soln}}$ ). This relationship can be expressed mathematically as

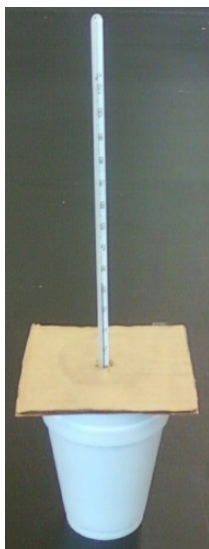
$$-q_{\text{rxn}} = q_{\text{cal}} + q_{\text{soln}} \quad (1)$$

Your goal in today's experiment is to determine the amount of heat released by the following reaction.



You will do this by determining the amount of heat absorbed by the calorimeter ( $q_{\text{cal}}$ ) and absorbed by the solution ( $q_{\text{soln}}$ ).

The reaction shown in equation 2 will be performed in a "coffee cup calorimeter" (Figure 1). The primary requirement of a calorimeter is that it should not transfer heat with the outside environment. So, while the nested coffee cups may look crude, the insulating properties of the styrofoam allow the calorimeter to perform admirably well. The heat absorbed by the calorimeter can be expressed as a function of the heat capacity of the calorimeter ( $C_{\text{cal}}$ ) and the change in temperature ( $\Delta T_{\text{cal}}$ ):



**Figure 1.** Coffee cup calorimeter

$$q_{\text{cal}} = C_{\text{cal}} \times \Delta T_{\text{cal}} \quad (3)$$

where  $C_{\text{cal}}$  has the units J/°C and  $\Delta T_{\text{cal}}$  is measured in °C. Multiplication of these two terms results in a value of heat with units of joules (J). The heat capacity of the calorimeter ( $C_{\text{cal}}$ ) is unique for each calorimeter and must be determined in the laboratory by a separate experiment.

### Determining the Heat Capacity of the Calorimeter ( $C_{\text{cal}}$ )

The heat capacity of the calorimeter can be determined by comparing the heat lost by a sample of hot water to the heat gained by a sample of cold water, once they are mixed together in the calorimeter. When a sample of hot water is mixed with cold water in the calorimeter, a transfer of heat occurs from the hot water ( $-q_{\text{hot}}$ ) into the cold water ( $+q_{\text{cold}}$ ) and into the calorimeter ( $+q_{\text{cal}}$ ).

$$-q_{\text{hot}} = q_{\text{cold}} + q_{\text{cal}} \quad (4)$$

In other words, the cold water and the calorimeter absorb the heat lost by the hot water. Rearranging equation 4, solving for  $q_{\text{cal}}$ , gives:

$$q_{\text{cal}} = -q_{\text{hot}} - q_{\text{cold}} \quad (5)$$

Plugging  $q_{\text{cal}}$  into equation 3, and rearranging to solve for  $C_{\text{cal}}$  gives:

$$C_{\text{cal}} = \frac{-q_{\text{hot}} - q_{\text{cold}}}{\Delta T_{\text{cal}}} \quad (6)$$

The hot and cold water heat terms of equation 6 can be expressed as follows:

$$q_{\text{hot}} = (c_{\text{H}_2\text{O}})(m_{\text{H}_2\text{O}})(\Delta T_{\text{hot}}) = \left(4.184 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}}\right)(50.0\text{g})(\Delta T_{\text{hot}}) \quad (7)$$

$$q_{\text{cold}} = (c_{\text{H}_2\text{O}})(m_{\text{H}_2\text{O}})(\Delta T_{\text{cold}}) = \left(4.184 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}}\right)(50.0\text{g})(\Delta T_{\text{cold}}) \quad (8)$$

In equations 7 and 8,  $c_{\text{H}_2\text{O}}$  is the *specific* heat capacity of water, given as  $4.184 \text{ J}/(\text{g} \cdot ^\circ\text{C})$ ,  $m_{\text{H}_2\text{O}}$  is the mass of water that will be used, 50.0 grams in each case, and  $\Delta T$  is the change in temperature of the sample of water.

$\Delta T$  is always calculated as  $T_f - T_i$  (final temp – initial temp).

$$\Delta T_{\text{hot}} = T_{f_{\text{hot}}} - T_{i_{\text{hot}}} \quad (9)$$

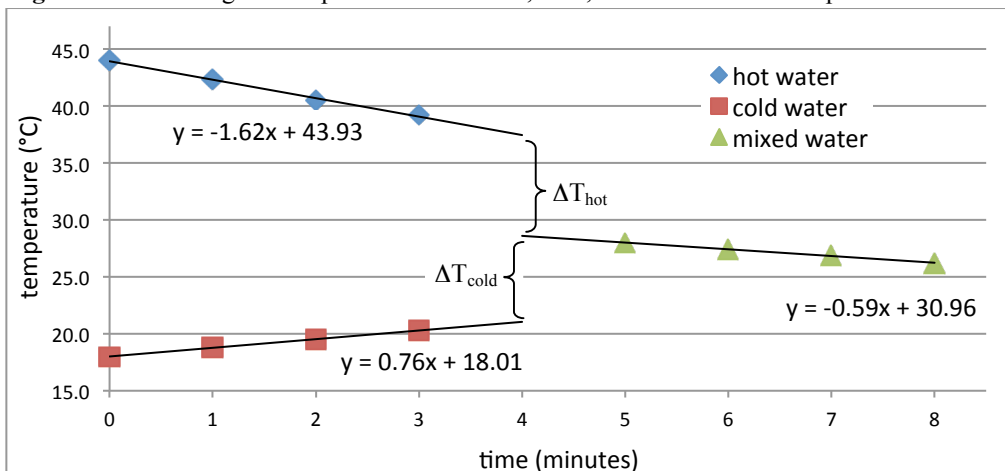
$$\Delta T_{\text{cold}} = T_{f_{\text{cold}}} - T_{i_{\text{cold}}} \quad (10)$$

For the hot and cold water,  $T_{i_{\text{hot}}}$  and  $T_{i_{\text{cold}}}$  will be the respective temperature of each water sample at the mixing time (4.0 minutes). The final temperature of both the hot and cold water,  $T_{f_{\text{hot}}}$  and  $T_{f_{\text{cold}}}$ , will be the temperature of the water at the time you have mixed the hot and cold.

$$T_{f_{\text{hot}}} = T_{f_{\text{cold}}} = T_{\text{mix}} \text{ at mixing time} \quad (11)$$

To determine  $T_{i_{\text{hot}}}$ ,  $T_{i_{\text{cold}}}$  and  $T_{\text{mix}}$  plot the temperatures of the water at each recorded time (Figure 2, for example) and use the appropriate trendline to calculate the temperatures at the mixing time, 4.0 minutes (equations 12, 13, 14). The procedural details of how all of this data will be generated, recorded, and plotted are given in the Procedure.

**Figure 2.** The change in temperature for the hot, cold, and mixed water samples.



$$T_{i_{\text{hot}}} = (-1.62)(4.00) + 43.93 = 37.45 \text{ }^{\circ}\text{C} \quad (12)$$

$$T_{i_{\text{cold}}} = (0.76)(4.00) + 18.01 = 21.1 \text{ }^{\circ}\text{C} \quad (13)$$

$$T_{f_{\text{hot}}} = T_{f_{\text{cold}}} = T_{\text{mix}} = (-0.59)(4.00) + 30.96 = 28.6 \text{ }^{\circ}\text{C} \quad (14)$$

Plugging these values back into equations 9 and 10 gives:

$$\Delta T_{\text{hot}} = 28.6 \text{ }^{\circ}\text{C} - 37.45 \text{ }^{\circ}\text{C} = -8.9 \text{ }^{\circ}\text{C} \quad (9)$$

$$\Delta T_{\text{cold}} = 28.6 \text{ }^{\circ}\text{C} - 21.1 \text{ }^{\circ}\text{C} = 7.5 \text{ }^{\circ}\text{C} \quad (10)$$

$\Delta T_{\text{hot}}$  and  $\Delta T_{\text{cold}}$  can be plugged back into equations 7 and 8, respectively to give:

$$q_{\text{hot}} = (c_{\text{H}_2\text{O}})(m_{\text{H}_2\text{O}})(\Delta T_{\text{hot}}) = \left(4.184 \frac{\text{J}}{\text{g}\cdot^{\circ}\text{C}}\right)(50.0\text{g})(-8.9^{\circ}\text{C}) = -1.9 \times 10^3 \text{ J} \quad (7)$$

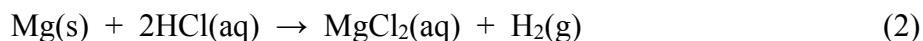
$$q_{\text{cold}} = (c_{\text{H}_2\text{O}})(m_{\text{H}_2\text{O}})(\Delta T_{\text{cold}}) = \left(4.184 \frac{\text{J}}{\text{g}\cdot^{\circ}\text{C}}\right)(50.0\text{g})(7.5^{\circ}\text{C}) = 1.6 \times 10^3 \text{ J} \quad (8)$$

Now *finally*, we can solve for  $C_{\text{cal}}$  by plugging  $q_{\text{hot}}$  and  $q_{\text{cold}}$  into equation 6, along with substituting  $\Delta T_{\text{cold}}$  for  $\Delta T_{\text{cal}}$ .

$$C_{\text{cal}} = \frac{-q_{\text{hot}} - q_{\text{cold}}}{\Delta T_{\text{cold}}} = \frac{-(-1.9 \times 10^3 \text{ J}) - (1.6 \times 10^3 \text{ J})}{7.5^{\circ}\text{C}} = 40 \text{ J}/^{\circ}\text{C} \quad (15)$$

You will perform two trials, generate two plots like Figure 2, and calculate the heat capacity of your calorimeter from your two sets of data, getting an average heat capacity. Use the average value for subsequent calculations in the next part.

### **Enthalpy of the Chemical Reaction: Determining $q_{\text{cal}}$ and $q_{\text{soln}}$**



For each trial, calculate the total heat released during the reaction using equation 1. The term  $q_{\text{cal}}$  will be calculated using equation 3, substituting in the *average*  $C_{\text{cal}}$  determined by equation 15, and  $\Delta T_{\text{cal}} = T_{f_{\text{cal}}} - T_{i_{\text{cal}}}$ .

$$q_{\text{cal}} = \text{average } C_{\text{cal}} \times (T_{f_{\text{cal}}} - T_{i_{\text{cal}}}) \quad (16)$$

Don't worry,  $T_{f_{\text{cal}}}$  and  $T_{i_{\text{cal}}}$  will be measured directly from the thermometer shown in Figure 1.

The term  $q_{\text{soln}}$  of equation 1 can be expressed as:

$$q_{\text{soln}} = (c_{\text{soln}})(m_{\text{soln}})(\Delta T_{\text{soln}}) = \left(3.97 \frac{\text{J}}{\text{g}\cdot^{\circ}\text{C}}\right) \left[ (D_{\text{HCl(aq)}} \times V_{\text{HCl(aq)}}) + m_{\text{Mg(s)}} \right] (\Delta T_{\text{soln}}) \quad (17)$$

where  $c_{\text{soln}}$  is the specific heat capacity of the resulting  $\text{MgCl}_2\text{(aq)}$  solution,  $m_{\text{soln}}$  is the mass of the  $\text{MgCl}_2\text{(aq)}$  solution, and  $\Delta T_{\text{soln}}$  is the change in temperature of the  $\text{MgCl}_2\text{(aq)}$  solution. The mass of the  $\text{MgCl}_2\text{(aq)}$  solution can be estimated by summing the mass of the reagents,  $\text{HCl(aq)}$

and Mg(s). The mass of HCl(aq) can be determined by multiplying the density of 1.0 M HCl(aq) ( $D = 1.015 \text{ g/mL}$ ) by the volume of HCl(aq) used, in mL. Because the calorimeter and the solution will start and end at the same temperature,

$$\Delta T_{\text{soln}} = \Delta T_{\text{cal}} \quad (18)$$

and the value of  $\Delta T$  used in equation 16 will be repeated in equation 17.

The sum of  $q_{\text{cal}}$  and  $q_{\text{soln}}$  will give the magnitude of heat *released* by the chemical reaction ( $-q_{\text{rxn}}$ , equation 1). To determine the "heat of the reaction", take the negative ( $-$ ) of  $-q_{\text{rxn}}$ .

$$q_{\text{rxn}} = -[q_{\text{soln}} + q_{\text{cal}}] \quad (19)$$

To determine the enthalpy of the chemical reaction (in kJ/mole), first express the heat in kJ, then divide that heat by the moles of Mg(s) used.

$$\Delta H_{\text{rxn}} = \frac{q_{\text{rxn}} (\text{kJ})}{\text{moles Mg(s)}} \quad (20)$$

Calculate a  $\Delta H_{\text{rxn}}$  value for each of your two trials of the chemical reaction, and report the average  $\Delta H_{\text{rxn}}$  and the percent error in your value. The accepted value for the  $\Delta H$  of the reaction in equation 2 is  $-470 \text{ kJ/mole}$ .

### **Calculation Summary**

Are you thoroughly confused and overwhelmed? Don't despair...too much. Follow the procedure, perform the experiment, and collect all of the required data. Then come back here and follow these 10 steps to calculate your value of  $\Delta H_{\text{rxn}}$ .

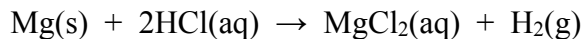
1. Prepare two charts (one for each trial) like Figure 2, using your own data. Follow the step-by-step instructions given at the end of the Procedure.
2. Calculate  $T_{i_{\text{hot}}}$ ,  $T_{i_{\text{cold}}}$  and  $T_{\text{mix}}$  by plugging in 4.00 minutes for  $x$  into the appropriate trendline.
3. Calculate  $\Delta T_{\text{hot}}$  and  $\Delta T_{\text{cold}}$  with equations 9 and 10. Remember that  $T_f$  for both hot and cold =  $T_{\text{mix}}$  and you calculated  $T_{\text{mix}}$  in step 2 above.
4. Plug  $\Delta T_{\text{hot}}$  and  $\Delta T_{\text{cold}}$  into equations 7 and 8, respectively, to calculate  $q_{\text{hot}}$  and  $q_{\text{cold}}$ .
5. Plug  $q_{\text{hot}}$ ,  $q_{\text{cold}}$ , and  $\Delta T_{\text{cold}}$  into equation 15 to calculate  $C_{\text{cal}}$ . Get an average  $C_{\text{cal}}$  from both of your trials.
6. Using your temperature data from the chemical reaction, determine  $\Delta T_{\text{rxn}}$ .  

$$\Delta T_{\text{rxn}} = (T_f - T_i) = \Delta T_{\text{soln}} = \Delta T_{\text{cal}}$$
7. Plug your average  $C_{\text{cal}}$  and  $\Delta T_{\text{cal}}$  (from trial 1 of the reaction) into equation 16 to solve for  $q_{\text{cal}}$  (for trial 1). Repeat this for your second trial, using the average  $C_{\text{cal}}$  and  $\Delta T_{\text{cal}}$  from trial 2.

8. Solve for  $q_{\text{soln}}$  by using equation 17, plugging in all known values and data. Again, solve for  $q_{\text{soln}}$  separately for each trial, using the appropriate  $\Delta T_{\text{soln}}$  value for that trial.
9. Solve for  $q_{\text{rxn}}$  (for each trial) using equation 19. Convert  $q_{\text{rxn}}$  in J to  $q_{\text{rxn}}$  in kJ.
10. Calculate the moles of Mg used in each trial and then calculate  $\Delta H_{\text{rxn}}$  for each trial using equation 20. Calculate an average  $\Delta H_{\text{rxn}}$  for your two trials.

## The Enthalpy Change of a Chemical Reaction with Magnesium - Procedure

**Purpose:** Determine the change in enthalpy ( $\Delta H_{\text{rxn}}$ ) for the reaction of magnesium metal with aqueous hydrochloric acid:



### **Procedure:**

Please work in pairs for this experiment. You will generate a lot of data in this experiment, so it will be most helpful to you if you are well prepared. Within the procedure for each part of today's experiment, there are sample data tables given. You are encouraged to generate tables in your notebook like those. ~~You may be allowed to cut and paste, into your notebook, the tables given, but that is at the discretion of your individual lab instructor.~~

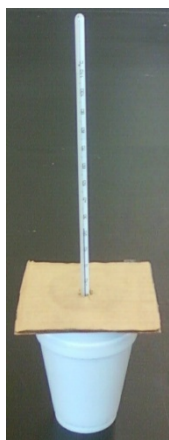
### **Part 1a. Thermometer Correction**

Obtain two thermometers and label them A and B. Immerse both thermometers into a beaker of water and read the temperature on each to 0.1 °C. If they do not agree, the difference in the readings is the correction which must be applied to one of the thermometers each time a reading is recorded. For example, if both thermometers are placed in the same water and thermometer A reads 25.1 °C, and thermometer B reads 25.3 °C, then to each reading of thermometer A add 0.2 °C as a correction. Alternatively, you could subtract 0.2 °C from each reading of thermometer B.

Part 1a. Thermometer Correction - Data	
	temperature (°C)
thermometer A	
thermometer B	

correction: thermometer \_\_\_\_\_ +/- \_\_\_\_\_ °C

### **Part 1b. Heat Capacity of Calorimeter**



**Figure 1.**  
Coffee cup calorimeter

For this procedure, two calorimeters are used. Each consists of *two nested* styrofoam cups, a cardboard cover, and your thermometers from Part 1a (Figure 1). One calorimeter will contain the hot water, and one will contain the cold water. Be consistent for each trial, so label the calorimeters "hot" and "cold". You will only use the "cold" calorimeter for Part 2, so it is critical that you do not mix them up.

Using a graduated cylinder, measure and dispense 50.0 mL of room temperature tap water into the "cold" calorimeter, and 50.0 mL of water, which has been heated to about 45 °C, into the "hot" calorimeter. At one minute intervals measure and record the temperatures of both calorimeters. At the fourth minute pour the warm water into the calorimeter containing the cold water. Read and record the temperature of the mixed water, in the "cold" calorimeter, at minutes 5, 6, 7, and 8. Make sure to record each temperature to one decimal place (0.1 °C). Repeat this entire procedure a second time. A sample table to organize your data is given in the next page.

**Part 1b. Heat Capacity of the Calorimeter - Data**

		0 min.	1 min.	2 min.	3 min.	4 min.	5 min.	6 min.	7 min.	8 min.
trial 1	cold H <sub>2</sub> O temp. (°C)					mix H <sub>2</sub> O temp. (°C)				
	hot H <sub>2</sub> O temp. (°C)									
trial 2	cold H <sub>2</sub> O temp. (°C)					mix H <sub>2</sub> O temp. (°C)				
	hot H <sub>2</sub> O temp. (°C)									

**Part 2. Enthalpy of the Chemical Reaction**

Use a balance to weigh 0.500 g of magnesium turnings (Mg) and then place them into your "cold" calorimeter. Use a spatula to weight the magnesium so that you do not touch the magnesium. Oily residue from your hands will inhibit the reaction with acid. Using a graduated cylinder, measure out 100.0 mL of 1.0 M HCl. Remember to pour the acid into a clean beaker first, and then pour it into a graduated cylinder to accurately and safely measure the correct volume. Pour any excess 1.0 M HCl in your beaker into the waste container. Using one of your labeled thermometers (preferably the same thermometer used in the "cold" calorimeter), measure the temperature of the 1.0 M HCl in the graduated cylinder. Assume that the "cold" calorimeter and the Mg(s) are at the same initial temperature as the 1.0 M HCl solution.

Pour the 1.0 M HCl solution into the "cold" calorimeter which contains the Mg(s). While stirring gently with your thermometer, measure and record the temperature immediately after mixing and then every 15 seconds until the temperature reaches a maximum and declines for two consecutive time intervals. Be certain that you record the maximum temperature even if it occurs at mid-interval. Also, once again, be sure to record your temperatures to one decimal place.

Once you have finished collecting temperature data for trial 1, dispose of the resulting MgCl<sub>2</sub>(aq) reaction solution in your calorimeter into the waste container. Clean and dry your calorimeter and repeat the entire Part 2 experiment a second time for trial 2. A sample table to organize your data is given in the next page. You do not have to record temperatures for all 390 seconds in the table, but only until the temperature reaches a maximum and declines for two consecutive time intervals.

**Waste disposal:**

- Dispose of excess 1.0 M HCl into the waste container.
- Dispose of the completed reaction solutions, MgCl<sub>2</sub>(aq) into the waste container.
- Clean and dry the styrofoam cups and return them to the supply cart.

**Conclusion:**

Report the average  $q_{\text{rxn}}$ , the average  $\Delta H_{\text{rxn}}$ , and your percent error in  $\Delta H_{\text{rxn}}$ . Comment on the accuracy of your  $\Delta H_{\text{rxn}}$  value by discussing the percent error. What are possible sources of error that may have affected, or could affect your accuracy? In your opinion, do you trust your calculations and results, and was the lab successful? What did you like/not like about this lab?

## Part 2. Enthalpy of the Reaction

trial 1

mass of Mg (g):	
vol. of HCl(aq) (mL):	
molarity of HCl(aq):	

time (s)	temp. (°C)
0	
15	
30	
45	
60	
75	
90	
105	
120	
135	
150	
165	
180	
195	
210	
225	
240	
255	
270	
285	
300	
315	
330	
345	
360	
375	
390	

trial 2

mass of Mg (g):	
vol. of HCl(aq) (mL):	
molarity of HCl(aq):	

time (s)	temp. (°C)
0	
15	
30	
45	
60	
75	
90	
105	
120	
135	
150	
165	
180	
195	
210	
225	
240	
255	
270	
285	
300	
315	
330	
345	
360	
375	
390	



## Plotting Data to Determine the Heat Capacity of the Calorimeter (MS Excel 2010)

### I. Enter Data

1. Enter time (x-axis) vs. temperature (y-axis) data for the hot water, cold water, and mixed water similarly to the data on the next page.

### II. Plotting the Three Data Sets

1. Select the hot water data and plot as you would normally.
2. Right click anywhere on the graph → **Select Data...**  
("Select Data Source" window opens)
3. In "Legend Entries (Series)" sub-window, click on **Series 1**
4. Click **Edit** button directly above – "Edit Series" window opens
  - A. In **Series Name:** enter "hot water"
  - B. Click **OK** – takes you back to "Select Data Source" window
5. In "Legend Entries (Series)" sub-window, click **Add** button
  - A. In **Series Name:** enter "cold water"
  - B. In **Series X values:**
    - in Excel sheet, highlight only the X-data (time) for cold water
  - C. In **Series Y values:**
    - highlight and delete current entry
    - in Excel sheet, highlight only the Y-data (temperature) for cold water
  - D. Click **OK** – takes you back to "Select Data Source" window
6. In "Legend Entries (Series)" sub-window, click **Add** button
  - A. In **Series Name:** enter "mixed water"
  - B. In **Series X values:**
    - in Excel sheet, highlight only the X-data (time) for mixed water
  - C. In **Series Y values:**
    - highlight and delete current entry
    - in Excel sheet, highlight only the Y-data (temperature) for mixed water
  - D. Click **OK** – takes you back to "Select Data Source" window, click **OK**

At this point, you should have three differently colored, and shaped, sets of data markers on your graph. Do not delete the color and shape legend on the graph!

### III. Adding Trendlines to Each Data Set

1. Right click on any data point of the "hot water" series
  - add linear trendline with equation and R-squared
2. Repeat for the "cold water" and "mixed water" series
3. You may delete the newly added legend items "Linear (hot water)", "Linear (cold water)" and "Linear (mixed water)".

### IV. Add a Chart Title and Axes Titles

- When complete, your graph should look like the one on the next page.