



## Experiment 7: Thermochemistry

### Enthalpy of Reaction

#### Safety First!

**Safety Goggles must be worn at all times in the Laboratory.**

#### Potential Hazards

- NaOH (sodium hydroxide) is caustic and can cause chemical burns. These are not painful at first, so you may not realize if you have gotten it on your skin; wash hands after using.
- HCl (hydrochloric acid) is corrosive, can cause painful chemical burns, and will eat holes in your clothes. Rinse hands with water if you get any on you. See your TA for help cleaning up a spill.

#### Waste Disposal

- Neutralized acids and bases may be disposed of in the sink.
- Dissolved magnesium may be disposed of in the sink.
- Magnesium and copper sulfate solutions should be disposed of in the marked chemical waste containers.

#### Experiment Objective(s):

- Determine the heat capacity of a constant pressure (coffee cup) calorimeter, determine the heat of neutralization of an acid/base reaction, determine the heat of reaction of the dissolution of magnesium metal, determine the heat of dissolution of a copper salt and its hydrate, and indirectly determine the heat of hydration.

#### Learning Objectives:

In this experiment you will learn to make accurate temperature measurements. You will also apply your knowledge of thermochemistry to experimentally determine calorimeter constants, and enthalpies of reaction.

#### Background:

In preparation for this experiment you should review chapter 6, sections 6.5 and 6.6, of your textbook, Chemistry (Chang & Goldsby), Custom Edition for IIT, 2015.

Energy changes that occur during chemical or physical changes are commonly measured with a calorimeter. The calorimeter used in this experiment is a simple polystyrene insulated foam cup that will be used as the reaction container. Based on the temperature changes, the number of moles of reactant used, and the capacity of the system, the molar enthalpy change resulting from a chemical or physical process will be measured.

According to the first law of thermodynamics, the energy change in an isolated system is zero. If we mix a known quantity of warm water with a known quantity cold water in the above calorimeter, because no heat being transferred to the surroundings, the heat lost by the warm water ( $q_w$ ) plus the heat gain by the cold water ( $q_c$ ) and the calorimeter ( $q_{cal}$ ) must be zero,

$$q_w + q_c + q_{cal} = 0$$



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Note that  $q_w$  will be negative, since heat is flowing out of the warm water and  $q_c$  and  $q_{cal}$  will be positive, since heat is flowing into the cold water and the calorimeter.

When replacing the expression for heat with a term incorporating mass, heat capacity, and temperature change, the above equation becomes

$$m_w c_w \Delta T_c + m_c c_c \Delta T_c + K \Delta T_c = 0$$

where  $K$  represents the heat capacity of the calorimeter, also called calorimeter constant. The specific heat capacity of warm and cold water is almost the same,  $4.18 \text{ J/g}^\circ\text{C}$ . Thus if the temperatures of the warm water, cold water, and the water after mixing are known, and if the mass of the warm and cold water is known, then  $K$  of the calorimeter may be calculated; determining the calorimeter constant will be done in Part A of the Procedure.

When a chemical reaction takes place in the calorimeter, the change in heat from the reaction ( $q_{\text{reac}}$ ) along with the change in heat of the solution in the calorimeter ( $q_{\text{sol}}$ ) and the calorimeter itself ( $K \Delta T_{\text{sol}}$ ) is again equal to zero:

$$q_{\text{reac}} + q_{\text{sol}} + K \Delta T_{\text{sol}} = 0$$

$$\text{where } q_{\text{sol}} = m_{\text{sol}} c_{\text{sol}} \Delta T_{\text{sol}}$$

In this experiment, it will be assumed that the specific heat capacity of the solutions used will be the same as that of water ( $4.18 \text{ J/g}^\circ\text{C}$ ). If the mass of the solution and the temperature change are known, the heat lost, or gained in a chemical reaction may be calculated.

There are different ways to quantify the change in energy of a system between its initial and final conditions. The change in energy can be determined from the amount of heat,  $q$ , and the amount of work,  $w$ , that the system exchanges with the surroundings.

**The enthalpy change**,  $\Delta H$ , for a chemical reaction is equal to the change in heat energy, measured at *constant pressure*. To determine the molar enthalpy change for the reaction, the number of moles,  $n$ , of the *limiting* reactant is required which may be calculated from the volume and concentration or the mass of the reagent used.

$$\Delta H_{\text{reac}} = q_{\text{reac}} / n$$

In this experiment, you will determine the enthalpy of dissolution of magnesium solid according to the red-ox reaction by measuring the amount of magnesium and using an excess of acid.

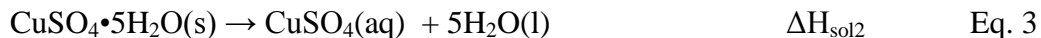


**The enthalpy of solvation** will be measured by examining two different types of copper sulfate salt; copper(II) sulfate anhydrous,  $\text{CuSO}_4\text{(s)}$ , and copper(II) sulfate pentahydrate,  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O(s)}$ , also known as bluestone. The chemical reactions showing the solvation (also called dissolution) of these two compounds are shown in Equations 2 and 3. Despite the similarity between these two reactants, their changes in enthalpy,  $\Delta H_{\text{sol}}$ , are different.



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#### Procedures:

##### Equipment

- 400 mL beakers
- 2 Styrofoam cups
- Thermometer
- Rubber bands
- Cardboard square
- 50 mL graduated cylinder or volumetric pipet

##### Chemicals

- $\text{CuSO}_4$  (s)
- $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  (s)
- 1 M hydrochloric acid (HCl)
- 1 M sodium hydroxide (NaOH)
- Mg (s)

Follow the procedures outlined on the following pages; record data, observations, and any deviations from the prescribed procedure as you work; use blue or black ink. You must have the “landscape pages” with your data and observations signed by your TA before you leave the lab, and you will submit them with your lab report. Be sure to always record the units and the correct number of significant figures for any measured value.



## Experiment 7: Thermochemistry Enthalpy of Reaction

### Pre-lab Questions (10pts)

*Pre-lab questions must be completed and turned in at the beginning of the lab period. If you have not completed the pre-lab questions, you will **not** be allowed to complete the lab.*

1. For an endothermic reaction, what is the sign of  $\Delta H$ ? (1pt)
2. A 225 mL sample of water is cooled from 60.0 °C to 25.5 °C. What is value  $q_{\text{sys}}$  ? (1pt)
3. Define the term “heat capacity”. (1pt)
4. Define the term “specific heat,” also known as “specific heat capacity.” (1pt)
5. Calculate the final temperature when 37 mL of water at 75 °C are added to 49 mL of water at 22 °C. Assume that no heat is lost to the calorimeter or the surroundings. (2pts)

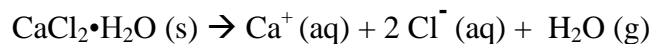


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6. A piece of metal with a mass of 27.0 g at a temperature of 73.2 °C is placed in a calorimeter containing 20.00 mL of water at 22.1 °C. The final temperature of the metal and water at equilibrium is 33.6 °C. Assuming that no heat is transferred to the calorimeter, what is the specific heat of the metal? (2pts)

7. When 5.65 g of solid calcium chloride monohydrate,  $\text{CaCl}_2 \cdot \text{H}_2\text{O}(\text{s})$  is dissolved in 100.00 g of water, the temperature rises from 22.7 °C to 27.8 °C.




Calculate the change in enthalpy for the dissolution of calcium chloride monohydrate (in kJ/mol of  $\text{CaCl}_2 \cdot \text{H}_2\text{O}$ ). Assume that the calorimeter is perfect (no heat is transferred to the calorimeter), and that the specific heat of the solution is the same as that of pure water (4.184 J/g °C). (2pts)

## Experiment 7: Thermochemistry

### Enthalpy of Reaction

Name \_\_\_\_\_

Partner(s) \_\_\_\_\_

Procedure	Data and Observations																																																																				
<p><b>SAFETY:</b> Safety Goggles must be worn at all times in the Laboratory.</p> <ul style="list-style-type: none"> <li>NaOH (sodium hydroxide) is caustic and can cause chemical burns. These are not painful at first, so you may not realize if you have gotten it on your skin; wash hands after using.</li> <li>HCl (hydrochloric acid) is corrosive, can cause painful chemical burns, and will eat holes in your clothes. Rinse hands with water if you get any on you. See your TA for help cleaning up a spill.</li> </ul> <p><b>A. Heat capacity of the Calorimeter</b></p> <ol style="list-style-type: none"> <li>Construct a calorimeter from two Styrofoam cups nested one inside the other. Put them into a 400 mL beaker for added stability. Place the rubber bands around the thermometer so that when inserted through the cardboard and place over the calorimeter, the thermometer is near the bottom without touching the Styrofoam.</li> <li>Measure 50.0 mL of distilled water and put it in the calorimeter. Wait 5 or 10 minutes for the system to reach thermal equilibrium.</li> <li>Record the temperature of the cool water to the nearest 0.1 °C.</li> <li>Measure 50.0 mL of distilled water and put it in a clean, dry 250 mL beaker. On a low setting on the hot plate, heat the water slowly until it is 15 °C – 20 °C above room temperature. Do not over heat the water and do not heat the water for a long time. If the water temperature exceeds 20 °C above room temperature, measure out a new sample of 50.0 mL and start again. If the water remains on the hot plate for more than 10 minutes, discard the sample and start again.</li> <li>When the heated water is in the correct temperature range, remove it from the hot plate and place it on the bench top for 1 minute.</li> </ol>	<div style="border: 1px solid black; padding: 10px; margin-bottom: 10px;">  <p>Why does the procedure call for using EXACTLY 50.0 mL of hot and cool water in this experiment?</p> </div> <p><b>A. Heat capacity of Calorimeter Data</b></p> <table border="1" style="width: 100%;"> <thead> <tr> <th>Initial Temperatures</th> <th>Trial 1</th> <th>Trial 2</th> <th>Trial 3(if needed)</th> </tr> </thead> <tbody> <tr> <td>Cool water</td> <td></td> <td></td> <td></td> </tr> <tr> <td>Hot water</td> <td></td> <td></td> <td></td> </tr> </tbody> </table> <table border="1" style="width: 100%;"> <thead> <tr> <th>Temperature vs time</th> <th>Trial 1</th> <th>Trial 2</th> <th>Trial 3(if needed)</th> </tr> </thead> <tbody> <tr><td>0 seconds</td><td></td><td></td><td></td></tr> <tr><td>15 seconds</td><td></td><td></td><td></td></tr> <tr><td>30 seconds</td><td></td><td></td><td></td></tr> <tr><td>45 seconds</td><td></td><td></td><td></td></tr> <tr><td>60 seconds</td><td></td><td></td><td></td></tr> <tr><td>75 seconds</td><td></td><td></td><td></td></tr> <tr><td>90 seconds</td><td></td><td></td><td></td></tr> <tr><td>105 seconds</td><td></td><td></td><td></td></tr> <tr><td>120 seconds</td><td></td><td></td><td></td></tr> <tr><td>135 seconds</td><td></td><td></td><td></td></tr> <tr><td>150 seconds</td><td></td><td></td><td></td></tr> <tr><td>165 seconds</td><td></td><td></td><td></td></tr> <tr><td>180 seconds</td><td></td><td></td><td></td></tr> </tbody> </table>	Initial Temperatures	Trial 1	Trial 2	Trial 3(if needed)	Cool water				Hot water				Temperature vs time	Trial 1	Trial 2	Trial 3(if needed)	0 seconds				15 seconds				30 seconds				45 seconds				60 seconds				75 seconds				90 seconds				105 seconds				120 seconds				135 seconds				150 seconds				165 seconds				180 seconds			
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6. Record the temperature of the hot water to the nearest 0.1 °C and then quickly add it to the cool water already in the calorimeter.
7. Replace the lid and thermometer of the calorimeter and stir gently.
8. Observe the temperature of the contents of the calorimeter for the next 3 minutes; record the temperature every 15 second.
9. Empty and dry the calorimeter. You may pour the lukewarm water down the drain.
10. Repeat steps 2-9 one more time.

#### B. Enthalpy of Neutralization – NaOH(aq) and HCl(aq)

1. Measure 50.0 mL of 1.0 M NaOH(aq) solution and add it to dry the calorimeter.
2. Measure 50.0 mL of 1.0 M HCl(aq) solution and add it to the clean, dry beaker. Place the HCl filled beaker on the benchtop near the calorimeter and let it stand 3 to 4 minutes.
3. Using a clean and dry thermometer, measure and record the temperature of the HCl solution to the nearest 0.1 °C.
4. Rinse the thermometer with distilled water and dry it with paper towel. Return the thermometer to the calorimeter; measure and record the temperature of the NaOH solution.
5. If the temperature difference is greater than 0.5°C, you will need to adjust the temperature of the HCl solution in the beaker. You will do this EITHER warming or cooling using a warm or cold water bath (be careful not to get any water in the beaker or else you will need to return to Step 3).
6. When the temperatures of the NaOH and HCl are within 0.5 °C, remove the lid of the calorimeter and carefully transfer the HCl solution to the calorimeter.
7. Replace the lid and thermometer. Observe the temperature and record every 15 second for 3 minutes.
8. Discard the neutralized solution down the drain, rinse and dry the calorimeter.
9. Repeat Steps 2-10 one more time.

#### B. Enthalpy of Neutralization Data

Initial Temperatures	Trial 1	Trial 2	Trial 3(if needed)
NaOH solution			
HCl solution			

Temperature vs time	Trial 1	Trial 2	Trial 3(if needed)
0 seconds			
15 seconds			
30 seconds			
45 seconds			
60 seconds			
75 seconds			
90 seconds			
105 seconds			
120 seconds			
135 seconds			
150 seconds			
165 seconds			
180 seconds			

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## Experiment 7: Thermochemistry

### Enthalpy of Reaction

Name \_\_\_\_\_

Partner(s) \_\_\_\_\_

#### C. Enthalpy of Reaction – Mg(s) and HCl(aq)

1. Measure 50.0 mL of 1.0 M HCl(aq) solution and add it to the dry calorimeter.
2. Weigh out between 0.050 g and 0.070 g of magnesium. Record the mass.
3. Measure the initial temperature of the HCl solution in the calorimeter to the nearest 0.1 °C.
4. Carefully add the magnesium to the calorimeter and gently stir with the thermometer. Record any signs you observe that a chemical reaction is taking place.
5. Record the temperature every 15 seconds for 3 minutes OR until one minute after signs of the reaction have stopped.
6. Transfer the reaction contents to the labeled waste container, rinse the calorimeter with distilled water and dry the calorimeter.
7. Repeat Steps 2-7 one more time.

#### C. Enthalpy of Reaction Data

Initial Temperatures	Trial 1	Trial 2	Trial 3
HCl solution			

Temperature versus time	Trial 1	Trial 2	Trial 3
0 seconds			
15 seconds			
30 seconds			
45 seconds			
60 seconds			
75 seconds			
90 seconds			
105 seconds			
120 seconds			
135 seconds			
150 seconds			
165 seconds			
180 seconds			
Final temperature one minute after reaction.			

Observations:

TA Signature: \_\_\_\_\_



## Experiment 7: Thermochemistry

### Enthalpy of Reaction

Name \_\_\_\_\_

Partner(s) \_\_\_\_\_

#### D. Enthalpy of Solvation – $\text{CuSO}_4(\text{s})$ and $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}(\text{s})$

1. Measure 50.0 mL of distilled water and add to the dry calorimeter.
2. Weigh out 2.00 g – 2.50 g of anhydrous copper sulfate,  $\text{CuSO}_4(\text{s})$ , and record the precise mass.
3. Measure the initial temperature of the water in the calorimeter.
4. Add the anhydrous copper sulfate to the calorimeter and gently stir to dissolve. When all the solid is dissolved, record the final temperature.
5. Transfer the reaction contents to the labeled waste container, rinse the calorimeter with distilled water and dry the calorimeter.
6. Repeat steps 2-6 one more time.
7. Now you will do the same procedure for the copper sulfate pentahydrate,  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}(\text{s})$ .
8. Measure 50.0 mL of distilled water and add to the empty and dry calorimeter.
9. Weigh out 2.00 g – 2.50 g of copper sulfate pentahydrate,  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}(\text{s})$ , and record the precise mass.
10. Measure the initial temperature of the water in the calorimeter.
11. Add the copper sulfate pentahydrate to the calorimeter and gently stir to dissolve. When all the solid is dissolved, record the final temperature.
12. Transfer the reaction contents to the labeled waste container, rinse the calorimeter with distilled water and dry the calorimeter.
13. Repeat steps 2-6 one more time.

Clean all glassware and let them dry at your station.

#### D. Enthalpy of Solvation Data

##### -Anhydrous copper sulfate

$\text{CuSO}_4(\text{s})$ Data	Trial 1	Trial 2	Trial 3
Initial water temp.			
Final solution temp.			

##### -Copper sulfate pentahydrate

$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}(\text{s})$ Data	Trial 1	Trial 2	Trial 3
Initial water temp.			
Final solution temp.			

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### Enthalpy of Reaction

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#### Calculations

Table 1: Calorimeter Constant Calculations

	Trial 1	Trial 2	Trial 3
Temperature change of cool water ( $\Delta T_c$ )			
Temperature change of warm water ( $\Delta T_w$ )			
Heat flow of cool water ( $q_c$ )			
Heat flow of warm water ( $q_w$ )			
Heat flow of calorimeter ( $q_{cal}$ )			
Calorimeter Constant (K)			
Average value of K			

#### Sample Calculations

*For each type of calculation, show one sample calculation with words, then substitute in numbers. Be sure to include the units in your calculation examples and in the data tables.*

For all these calculations, assume that the heat capacity of the solutions is the same as for pure water and is constant in this temperature range; 4.184 J/g °C.

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### Enthalpy of Reaction

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Table 2: Enthalpy of Neutralization Calculations

	Trial 1	Trial 2	Trial 3
Temperature change of acid/base solution ( $\Delta T$ )			
Heat flow of solution ( $q_{\text{sol}}$ )			
Heat flow of reaction ( $q_{\text{reac}}$ )			
Moles of reaction (n)			
Enthalpy of reaction ( $\Delta H_{\text{reac}}$ )			
Average value of $\Delta H_{\text{reac}}$			

Table 3: Enthalpy of Reaction Calculations

	Trial 1	Trial 2	Trial 3
Temperature change of acid/Mg solution ( $\Delta T$ )			
Heat flow of solution ( $q_{\text{sol}}$ )			
Heat flow of reaction ( $q_{\text{reac}}$ )			
Moles of magnesium (n)			
Enthalpy of reaction ( $\Delta H_{\text{reac}}$ )			
Average value of $\Delta H_{\text{reac}}$			

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### Enthalpy of Reaction

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Table 4.1: Enthalpy of Solvation anhydrous copper sulfate

	Trial 1	Trial 2	Trial 3
Temperature change of water/ $\text{CuSO}_4$ solution ( $\Delta T$ )			
Heat flow of solution ( $q_{\text{sol}}$ )			
Heat flow of reaction ( $q_{\text{reac}}$ )			
Moles of $\text{CuSO}_4(\text{s})$			
Enthalpy of reaction ( $\Delta H_{\text{reac}}$ )			
Average value of $\Delta H_{\text{reac}}$			

Table 4.2: Enthalpy of Solvation copper sulfate pentahydrate

	Trial 1	Trial 2	Trial 3
Temperature change ( $\Delta T$ ) of water/ $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ solution			
Heat flow of solution ( $q_{\text{sol}}$ )			
Heat flow of reaction ( $q_{\text{reac}}$ )			
Moles of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O} (\text{s})$			
Enthalpy of reaction ( $\Delta H_{\text{reac}}$ )			
Average value of $\Delta H_{\text{reac}}$			

TA Signature: \_\_\_\_\_



## Experiment 7: Thermochemistry

### Enthalpy of Reaction

#### Post-Lab Questions (10pts)

1. The enthalpy of hydration is given by the reaction



Measuring this value directly by reacting anhydrous copper sulfate with water and measuring the heat transferred during this process is very difficult.

Use Hess's Law and your experimental data to indirectly find the enthalpy of hydration. (You will need to combine equations describing the reactions conducted in the experiment in a way to create the hydration reaction.)

## Experiment 7: Thermochemistry

### Enthalpy of Reaction

2. Using Hess's Law and the data table for the enthalpies of formation given below, calculate the expected values for the reactions in this experiment. Compare these calculate values to your experimental results.

Reaction	$\Delta H$ (Experimental)	$\Delta H$ (Calculated)	Percent error
$\text{NaOH(aq)} + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(g)}$			
$\text{Mg(s)} + 2 \text{HCl(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)}$			
$\text{CuSO}_4\text{(s)} + 5\text{H}_2\text{O(l)} \rightarrow \text{CuSO}_4 \cdot 5\text{H}_2\text{O(s)}$			

Enthalpies of Formation Data Table

Compound	$\Delta H_f$ (solid)	$\Delta H_f$ (liquid)	$\Delta H_f$ (gas)	$\Delta H_f$ (aqueous)
$\text{MgCl}_2$	-641.3 kJ/mol			-801.2 kJ/mol
HCl			-92.3 kJ/mol	-167.2 kJ/mol
$\text{H}_2$			0 kJ/mol	
$\text{CuSO}_4$	-771.4 kJ/mol			-844.5 kJ/mol
$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$	-2279.7 kJ/mol			
$\text{H}_2\text{O}$		-285.8 kJ/mol	-241.8 kJ/mol	
NaOH	-425.6 kJ/mol			-469.6 kJ/mol
NaCl	-411.0 kJ/mol			-406.9 kJ/mol