EXPERIMENT #12

Specific Heat and Heat Capacity

OBJECTIVES:

- Accurately weigh masses of metal, solute, and water
- Quantitatively measure temperature changes
- Calculate heat exchange
- Calculate the specific heat of an unknown metal; calculate the heat of solution of an unknown solid

BACKGROUND:

Energy is the capacity to do work. Its SI unit of measurement is the Joule. Many chemical reactions and physical processes absorb or give off energy in the form of heat. Heat is a form of energy that moves from one object to another because of a temperature difference. The spontaneous, that is, the natural, direction of heat flow is from an object at high temperature to an object at low temperature. Calorimetry is the measurement of heat flow; the experimental apparatus is called a calorimeter. A calorimeter is designed to prevent the gain or loss of heat from its surroundings; but, within the calorimeter chemical reactions or solution processes may occur and/or heat may pass from one portion of the contents to another. A calorimeter usually contains a known amount of water at a known temperature. The change in water temperature is measured and the heat of the reaction or process is calculated. In this experiment you will use a coffee cup calorimeter (one Styrofoam™ cup inserted into another) with a simple cardboard cover to measure the specific heat of an unknown metal and, if time permits (extra credit anyone?) the heat of solution of an unknown solid.

I. Specific Heat of a Metal

The amount of heat which one object transfers to another depends upon the mass of the object, the temperature change it undergoes, and its specific heat. The specific heat, s, of an object is an intensive property describing the number of Joules it takes to raise the temperature of 1.00 g of the object 1.00°C. Several common substances are listed in the table below with their corresponding specific heats. Note that water has the highest specific heat. One very practical consequence of this is water's ability to moderate climates. In hot weather water can store large amounts of heat with a small temperature change, while in cold weather this heat is liberated with a small temperature change; the heat of the summer can be stored for the winter.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific Heat, s (J/g°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Al(s)</td>
<td>0.900</td>
</tr>
<tr>
<td>Au(s)</td>
<td>0.129</td>
</tr>
<tr>
<td>C(graphite)</td>
<td>0.720</td>
</tr>
<tr>
<td>Hg(l)</td>
<td>0.139</td>
</tr>
<tr>
<td><strong>H₂O(l)</strong></td>
<td><strong>4.184</strong></td>
</tr>
<tr>
<td>C₂H₅OH(l)</td>
<td>2.46</td>
</tr>
<tr>
<td>CaCO₃(s)</td>
<td>0.85</td>
</tr>
<tr>
<td>CCl₄(l)</td>
<td>0.86</td>
</tr>
<tr>
<td>Cu(s)</td>
<td>0.385</td>
</tr>
</tbody>
</table>
EXPERIMENT #12  SPECIFIC HEAT AND HEAT CAPACITY

The quantity of heat gained or lost by a substance or object, $q$, is equal to the product of its specific heat, $s$, mass, $m$, and change in temperature $\Delta T$.

$$q = (s) \times (m) \times (\Delta T) \quad (1)$$

where

$$\Delta T = T_{\text{final}} - T_{\text{initial}} \quad (2)$$

The portion of equation (1) that is the product of the specific heat and mass is known as the heat capacity $C$.

$$C = (s) \times (m) \quad (3)$$

Heat capacity is an *extensive property*.

If hot metal at a known temperature is placed into a calorimeter containing a known mass of water at a known temperature, the metal cools (loses heat energy), the water warms (gains heat energy), and the calorimeter warms (gains heat energy). Since a calorimeter isolates the system – water, metal, and calorimeter – from the surroundings, we have from the conservation of energy

$$q_{\text{calorimeter}} = 0 \quad (4)$$

But,

$$q_{\text{calorimeter}} = 0 = q_{\text{metal}} + q_{\text{water}} + q_{\text{coffee cup}} \quad (5)$$

We will simplify the calculations by assuming that the heat capacity of the coffee cup is negligible. So

$$q_{\text{metal}} = -q_{\text{water}} \quad (6)$$

The $q_{\text{metal}}$ and $q_{\text{water}}$ are obtained by proper application of equation (1)

$$q_{\text{metal}} = s_{\text{metal}} \times m_{\text{metal}} \times \Delta T_{\text{metal}} \quad (7)$$

$$q_{\text{water}} = s_{\text{water}} \times m_{\text{water}} \times \Delta T_{\text{water}} \quad (8)$$

Since the metal cools, $\Delta T_{\text{metal}}$ will be negative; since the water warms, $\Delta T_{\text{water}}$ will be positive. The negative sign in equation (6) will be eliminated, and a positive value of $s_{\text{metal}}$ will be obtained.

$$s_{\text{metal}} = \frac{-s_{\text{water}} \times m_{\text{water}} \times \Delta T_{\text{water}}}{m_{\text{metal}} \times \Delta T_{\text{metal}}} \quad (9)$$
**Example**

(a) How much heat is required to raise the temperature of 250. g of water from 22°C to 98°C?

\[
q = (4.18 \ \text{J/g} \cdot \text{°C}) \times (250. \ \text{g}) \times (98°C - 22°C) = 7.9 \times 10^4 \ \text{J}
\]

(b) What is the molar heat capacity of water?

\[
\text{molar heat capacity} = (4.184 \ \text{J/g} \cdot \text{°C}) \times (18.02 \ \text{g/mol}) = 75.40 \ \text{J/mol} \cdot \text{°C}
\]

**II. Heat of a Solution (Optional)**

When a chemical reaction or solution process is carried out in a calorimeter, the analysis is very similar. Reactants at room temperature are added to the calorimeter; any water in the calorimeter is also at room temperature. If the process is exothermic, the temperature of the liquid and calorimeter will rise, \( \Delta T \) will be positive; if the process is endothermic, the temperature of the liquid and calorimeter will go down, \( \Delta T \) will be negative. Since a calorimeter isolates the system -- water, process, and calorimeter -- from the surroundings, we have from the conservation of energy

\[
q_{\text{calorimeter}} = 0 \quad (4)
\]

But

\[
q_{\text{calorimeter}} = 0 = q_{\text{process}} + q_{\text{water}} + q_{\text{coffee cup}} \quad (11)
\]

We will simplify the calculations by assuming that the heat capacity of the coffee cup is negligible. So

\[
q_{\text{process}} = -q_{\text{water}} \quad (12)
\]

The \( q_{\text{water}} \) is obtained by proper application of equation (1)

\[
q_{\text{water}} = s_{\text{water}} \times m_{\text{water}} \times \Delta T_{\text{water}} \quad (13)
\]

Since the process is performed at constant pressure, \( q_{\text{process}} \) is equal to \( \Delta H_{\text{process}} \) and

\[
\Delta H_{\text{process}} = -s_{\text{water}} \times m_{\text{water}} \times \Delta T_{\text{water}} \quad (14)
\]
Example
When a student mixes 50.0 mL of 1.00 M HCl and 50.0 mL of 1.00 M NaOH in a coffee cup calorimeter, the temperature of the resultant solution increases from 21.0°C to 27.5°C. Calculate the enthalpy change for the reaction (The enthalpy change is the heat change at constant pressure.), and the enthalpy change per mole of HCl (or NaOH), assuming the specific heat of the calorimeter is zero, the total volume of the solution is 100.0 mL, the density of the solution is 1.00 g/mL, and the specific heat of the solution is 4.18 J/g°C -- the same as that for water.

From equation (12) we have,

\[ q_{\text{reaction}} = -q_{\text{water}} \]
\[ = -(4.18 \text{ J/g°C}) \times [(100.0 \text{ mL}) \times (1.00 \text{ g/mL})] \times (27.5°C - 21.0°C) \]
\[ = -2120 \text{ J} \]

[enthalpy change per mole HCl] \[= \frac{-2120 \text{ J}}{0.0500 \text{ L} \times 1.00 \text{ M}} = -4.24 \times 10^4 \text{ J/mol} \]

**EXPERIMENTAL PROCEDURE:** (Work in pairs.)

I. Specific Heat of a Metal
1. Obtain a pair of Styrofoam™ cups, cardboard cover, and a thermometer. The thermometer will also serve as a the stirring rod. Place about 700 mL of deionized water in a 1000 mL beaker, set on a hot plate, and begin heating the water to boiling.

2. Obtain a sample of unknown metal. Weigh 60-80.xxx g of metal into a 25 x 200 mm test tube. Loosely place a cork stopper into the mouth of the test tube.

3. Place the test tube, metal, and stopper into the beaker of boiling water. The water level should be above the level of metal in the test tube. Heat the metal in the boiling water bath for at least 10 minutes. Maintain the level of the water. The temperature of the boiling water is the temperature of the metal when placed in the calorimeter. Use the calorimeter thermometer to measure the temperature of the boiling water to the nearest 0.5°C. Wipe the thermometer dry before replacing in the calorimeter.

4. While the metal is being heated, weigh the calorimeter on the top-loading balance, add 40 mL of deionized water, and weigh the calorimeter again. The mass difference is the mass of water in the calorimeter.
5. Support calorimeter in a 400 mL beaker. Place the cover on the calorimeter, insert the thermometer (The bulb of the thermometer should be under the water.), and measure the temperature of the water to the nearest 0.5°C.

6. **Perform the following sequence quickly:** Take the test tube with the metal out of the boiling water bath, remove the stopper, and pour all of the metal into the calorimeter. Do not let water adhering to the outside of the test tube enter the calorimeter. Replace the cardboard cover and stir the water as best you can with the thermometer. Do not try to stir the metal; the thermometer bulb may break. Record to 0.5°C the maximum temperature reached by the water.

7. Repeat the experiment using about 50 mL of water in the calorimeter. Dry the metal before reusing it. Perform this operation in the hood. Remove most of the water by drying the metal with a paper towel. Place the semi-dry metal in a beaker, rinse the metal with acetone several times by decantation, and properly dispose of the acetone in the labeled waste container. **THE DRYING OPERATION WITH THE ACETONE MUST BE PERFORMED IN THE HOOD.**

8. After the second trial, dry the metal as above, replace it in the test tube, stopper the tube, and return to the unknown rack.

II. **Heat of a Solution (Optional)**

   Place about 50 mL of deionized water in the calorimeter and weigh as in **Part I.** Measure the temperature of the water to 0.5°C. The temperature should be within a degree or two of room temperature. On a piece of weighing paper, weigh out 5 g of your unknown solid to the nearest tenth of a milligram (analytical balance) and add to the calorimeter. Place the cover on the calorimeter and stir the contents with the thermometer. Determine the maximum or minimum temperature reached to the nearest 0.5°C. All of the solid must dissolved and at least a 5 degree temperature change must occur for the trial to be valid. If necessary, repeat the experiment, making adjustments in the mass of solid or volume of water. Properly dispose of the unknown solution.
DATA AND CALCULATIONS: Specific Heat and Heat Capacity

I. Specific Heat of a Metal

Unknown: __________

<table>
<thead>
<tr>
<th></th>
<th>Trial 1</th>
<th>Trial 2</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of stoppered test tube plus metal</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of test tube and stopper</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of calorimeter</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of calorimeter and water</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of water, $m_{\text{water}}$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of metal, $m_{\text{metal}}$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Initial temperature of water in calorimeter</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Initial temperature of metal</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Maximum temperature of metal and water in calorimeter</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

$\Delta T_{\text{water}}$
$\Delta T_{\text{metal}}$
$q_{\text{water}}$

Specific heat of the metal, $s$

Calculations
## II. Heat of Solution (Optional)

Unknown: __________

<table>
<thead>
<tr>
<th></th>
<th>Trial 1</th>
<th>Trial 2</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of solid, $m_{\text{solid}}$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of water, $m_{\text{water}}$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Initial temperature</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Maximum/minimum temperature of solution</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$q_{\text{water}}$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\Delta H_{\text{solution}}$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\Delta H_{\text{solution, Joules/gram}}$</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Calculations**
ADDITIONAL ASSIGNMENT I: Specific Heat and Heat Capacity

1. A piece of silver of mass 362 g has a heat capacity of 85.7 J/°C. What is the specific heat of silver?

2. A 6.22-kg piece of copper metal is heated from 20.5°C to 324.3°C. Calculate the heat absorbed (in kJ) by the metal.

3. Calculate the amount of heat liberated (in kJ) from 366 g of mercury when it cools from 77.0°C to 12.0°C.
4. A coffee cup calorimeter contains 150. g of water at 24.6°C. A 110. g block of molybdenum metal is heated to 100.°C and then placed in the water in the calorimeter. The temperature of the water rises to 28.0°C, at which point it stops rising. What is the specific heat of molybdenum metal? Ignore the heat capacity of the calorimeter.

5. When 2.50 g of NaOH were dissolved in 49.0 g water in a calorimeter at 24.0°C, the temperature of the solution went up to 37.1°C. Find ΔH for the reaction as it occurred in the calorimeter and find ΔH per gram of NaOH.
ADDITIONAL ASSIGNMENT II: Specific Heat and Heat Capacity

1. A 0.1375-g sample of solid magnesium is burned in a constant-volume bomb calorimeter that has a heat capacity of 3024 J/°C. The temperature increases by 1.126°C. Calculate the heat given off by the burning Mg, in kJ/g and in kJ/mol.

2. A 44.0-g sample of an unknown metal at 99.0°C was placed in a constant-pressure calorimeter containing 80.0 g of water at 24.0°C. The final temperature of the system was found to be 28.4°C. Calculate the specific heat of the metal. (The heat capacity of the calorimeter is 12.4 J/°C.)
3. When 50.0 mL of 0.100 M AgNO₃ and 50.0 mL of 0.100 M HCl are mixed in a constant pressure calorimeter, the temperature of the mixture increases from 22.30°C to 23.11°C. The temperature increase is caused by the following reaction:

\[
\text{AgNO}_3(aq) + \text{HCl}(aq) \rightarrow \text{AgCl(s)} + \text{HNO}_3(aq)
\]

Calculate ΔH for this reaction, assuming that the combined solution has a mass of 100.0 g and a specific heat of 4.184 J/(g•°C).