I. Introduction

The heat of fusion ($H_f$) is the amount of energy required to melt or freeze a substance and it is often expressed on a “per gram” or “per mole” basis. In this experiment, you will be determining the heat of fusion for ice and comparing it to the known value (6.02 kJ/mol).

$H_f$ is determined by adding a known amount of $0^\circ$C ice to a known amount of hot water in an insulated calorimeter that you will build before coming to lab. Obviously, the hot water loses heat energy as it cools down to a final temperature, $T_f$. However, the ice gains heat in two steps.

1. Melting ice: Because the initial temperature of the ice is $0^\circ$C, it will take heat from the hot water and melt.

2. Warming water: After the ice has melted, it leaves behind water at $0^\circ$C that must be warmed by the hot water that surrounds it to the final temperature of the mixture $T_f$.

Calculations and details are outlined below.

You will construct the calorimeter used in these experiments before coming to class. The calorimeter and lab report are both worth 10 points and will be graded separately. The calorimeter grade will be determined as follows:

- Overall design idea: 2 points
- Insulation choice: 2 points
- Built to specifications: 2 points
- Durability: 2 points
- Originality: 2 points

At the end of class, your calorimeter will compete against others in the calorimeter contest. In “heats” of four, each constant will receive ~75 mL of hot, near-boiling water. LoggerPro will be used to monitor the temperature of each calorimeter and the calorimeter that keeps its water hot the longest will advance to compete with other finalists to determine the overall winner. Some calorimeters will be selected for display in the display cases in the front of the Science Building.

What is a good insulator?

One of the best insulators is actually nothing at all… a vacuum! A cup of hot chocolate in a porcelain mug cools down because circulating, convective, air around the mug quickly carries away heat energy. In an insulating Thermos bottle (figure at right), the inner beverage container is surrounded by a vacuum jacket. In this case there is no air in contact with the container to circulate and carry away heat and the beverage stays hot (…or cold) much longer. Of course the opening at the top of the Thermos bottle is a potential heat leak but a well designed plug or stopper can minimize the losses.

Air by itself is actually a pretty good insulator as long as it isn’t allowed to circulate. This is the principle behind insulating foams. By trapping air in small bubbles (left), convective circulation can be kept to a minimum while still taking advantage of air’s insulating properties. Generally speaking, most common insulation materials take advantage of the effect by trapping the air in small bubbles/pockets and thus prevents it from circulating.

Good insulators also minimize heat lost by conductive and radiative processes. Conductive heat loss occurs as heat energy is transferred through a material to the outer surroundings. For example, if you heat the end of a copper wire in a flame, heat energy will be conducted through the copper from one end to the other. Eventually, the whole copper wire will become too hot to hold with your fingers. Thermos bottles minimize conductive heat loss by using glass (a poor heat conductor) in their construction and by making the glass walls very thin.
Radiative heat exchange occurs when objects emit or receive invisible infrared light. Sitting around a campfire you feel warm because of the infrared light that shines on you. A similar effect is used in restaurants where I.R. lamps are used to keep food warm until they are served to customers. Thermos bottles are often coated with a mirror-like substance to keep infrared radiation from the outer surroundings from being absorbed.

**What insulator should I choose for my calorimeter?**

Good question. Here are some of the materials that have performed well in the past: fiberglass, rock wool, spray foam* and Styrofoam. Other materials that didn’t do very well (Why?) include jello, sand and rice. Also, try to design your calorimeter so that the inner enclosure doesn’t absorb very much heat energy. For example, don’t use a metal inner container as it will absorb a lot of heat energy and lower the temperature of the hot water prematurely.

*Spray foam must be exposed to air to properly cure.

**Heat Transfer Calculations**

Some of the heat energy that melts the ice comes from the hot water. The amount of heat energy lost by the hot water is given by the following formula:

\[
Q_{\text{hot water}} = m_{\text{hot water}} \times c_{\text{H}_2\text{O}} \times (T_F - T_i)
\]

Equation #1

where \( Q \) is the energy lost by the hot water, \( m \) is the mass of the hot water, \( c \) is the specific heat of water \((4.184 \text{ J/g } \circ C)\), and \( T_F \) and \( T_i \) are the final and initial temperatures of the hot water respectively.

*As the ice melts* its temperature remains constant at \( 0 \circ C \). The energy required to melt the ice is given by:

\[
Q_{\text{melt}} = m_{\text{ice}} \times H_f
\]

Equation #2

where \( Q \) is the energy gained by the ice, \( m_{\text{ice}} \) is the mass of the ice, and \( H_f \) is the unknown heat of fusion in this experiment.

Note also that energy is also gained by the water formed when the ice cube melts. This water is initially at \( 0 \circ C \) and must be warmed up to the final temperature of the mixture. The energy required for this step is also furnished by the hot water & calorimeter and is given by the following formula:

\[
Q_{\text{cold water}} = m_{\text{ice}} \times c_{\text{H}_2\text{O}} \times (T_f - 0.00)
\]

Equation #3

Where \( m_{\text{ice}} \) is used since the ice has now melted and exists as additional liquid water. Note that the initial temperature of the water formed via this process is \( 0 \circ C \) (i.e. the ice that melts forms liquid water at \( 0 \circ C \))

In today's experiment, you will measure all masses and temperatures. The only unknown will be the heat of fusion, \( H_f \). You can then calculate actual values for the amount of heat lost by the warm water, the calorimeter, and the amount of heat gained by the ice water via the equations above.

The amount of heat required to melt the ice is then calculated via the following formula describing conservation of energy for this process:

\[
\frac{Q_{\text{lost}}}{Q_{\text{hot water}}} = - \frac{Q_{\text{gained}}}{(Q_{\text{melt}} + Q_{\text{cold water}})}
\]

Equation #4

Values for \( Q_{\text{hot water}}, Q_{\text{cold water}} \) are calculated and used to determine the value for \( Q_{\text{melt}} \) and \( H_f \).
II. Procedure

Calorimeter Construction

The mass of the calorimeter plus water & ice must be below the weight limit of the laboratory balances (400 grams).

The stainless steel Vernier temperature probe shown at right (note dimensions) will be used to monitor the temperature of the solution and to stir the ice/water mixture. Make sure that it can easily reach the ice/water mixture at the bottom of your calorimeter.

If your calorimeter uses a lid, make sure to provide a hole for the temperature probe. The temperature probe must be able to reach the ice/water mixture at the bottom of the calorimeter.

Calorimetry

Place approximately 30 grams of ice in a beaker to warm. The ice must be melting slightly to guarantee that its temperature is 0°C.

Pour approximately 50 mL of hot tap water (~ 45°C) into your dried, pre-weighed calorimeter. Reweigh the water filled calorimeter and record the mass.

*These amounts are only approximate and it may be necessary to use more or less water/ice depending on the volume of your calorimeter and how deep it is.

Insert the temperature probe into your calorimeter and click on the “Collect” button in Logger Pro.

Do NOT stop collecting data until the entire experiment is finished (approx. 6 minutes total time).

Stir well at all times. Inadequate stirring leaves the temperature probe in hot or cold pockets and produces erratic temperature behavior.

After collecting data for ~ 2 minutes, pour off any water that has formed in the ice-containing beaker, dry the ice with a paper towel and add approximately 10 grams to the calorimeter. Continue collecting temperature data and try to keep the temperature probe immersed in the hot water for best results.

Collect data for another 3-4 minutes stirring at all times. The ice should be completely melted at the end of the data collection run. If not, repeat the run with a smaller amount of ice.

After stopping the data acquisition, touch the end of the temperature probe to an inside surface of the calorimeter to transfer any water clinging to the probe. Reweigh the calorimeter/water and record the mass in your notebook.

Repeat the procedure for the calorimeter as many times as necessary to get two good trials. Repeat the procedure for your partner’s calorimeter.

Use the Logger Pro graph to determine the initial and final temperatures for each trial and record these values in the data table below.
III. Data Analysis

1. Calculate the masses of the hot water and ice from your experimental measurements.

2. Use equations 1 & 3 to determine the heat lost by the hot water ($Q_{\text{hot water}}$) and the heat gained by the ice water ($Q_{\text{cold water}}$).

3. Calculate the heat required to melt the ice ($Q_{\text{melt}}$) via equation 4.

4. Determine the value of $H_f$ using equation 2 in both Joules per gram and Joules per mole units.

5. Use your book or the internet to find an accurate value of $H_f$ in units of Joules per gram of ice.

6. Determine how closely your value of $H_f$ compares to the known value by performing a $\Delta\%$ calculation for both trials.

IV. Data Table

<table>
<thead>
<tr>
<th></th>
<th>Calorimeter Trial #1</th>
<th>Calorimeter Trial #2</th>
<th>Calorimeter Trial #3 (optional)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$m_{\text{calorimeter}}$</td>
<td>(g)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$m_{\text{calorimeter}} + \text{hot water}$</td>
<td>(g)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$m_{\text{calorimeter}} + \text{hot water} + \text{ice}$</td>
<td>(g)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$m_{\text{ice}}$</td>
<td>(g)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$m_{\text{hotwater}}$</td>
<td>(g)</td>
<td></td>
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</tr>
<tr>
<td>$T_i$</td>
<td>($^\circ$C)</td>
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</tr>
<tr>
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<td>($^\circ$C)</td>
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<td>(J)</td>
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<td></td>
</tr>
<tr>
<td>$Q_{\text{cold water}}$</td>
<td>(J)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$Q_{\text{melt}}$</td>
<td>(J)</td>
<td></td>
<td></td>
</tr>
<tr>
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<td>($J/g_{\text{ice}}$)</td>
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<tr>
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<tr>
<td>$H_f \Delta%$</td>
<td></td>
<td>6.02 kJ/mol</td>
<td>6.02 kJ/mol</td>
</tr>
</tbody>
</table>
V. Individual Report

*All must be word processed, clearly worded and on one page.*

- **Upper right corner:** Your name, Your lab section number, Date of experiment

- **Data table:** Obtain a copy of the data table from the lab-handout website and fill in your values using MS Word. Results will be graded according to how closely they compare to my calculations. Don’t round until you have a value for \( H_f \)

- **Answers to the following questions:**

  1. Explain why it was necessary to let the ice sit out at room temperature for several minutes before using it in the experiment?

  2. A student measures incorrectly and mistakenly records a value for the ice’s mass that is too high. How will this mistake affect the calculated value for \( H_f \)? Explain.

  3. Ice diving involves wearing SCUBA gear, digging a hole through the ice on a lake and swimming around under the ice. How is it possible for temperatures below the ice to actually be warmer than surface temperatures?

  4. Now that you’ve tested your calorimeter, describe one change you would make to improve its performance.