THE ENERGY OF PHASE CHANGES
The Energy of Phase Changes

Introduction

Consider heating a solid: as the solid is warmed, energy from the source of heat is "put into" the solid, and the solid gains energy. If the heating is continued, the warming solid will eventually reach its **melting point** and convert to a liquid. To complete this transformation, more energy (e.g., in the form of heat) must be added to the substance. However, as the substance changes phase, the added energy goes into accomplishing this transformation and the temperature stays the same. Thus, a single melting point temperature can be determined at which the solid converts to the liquid. This is an example of a **phase transition**. The amount of heat required to accomplish the solid-to-liquid phase change at the melting temperature is an **enthalpy** and is called the **heat of fusion**.

Of course, if the resulting liquid is also heated its temperature will increase as more energy is put into it until the liquid eventually reaches its **boiling point**, at which another phase change occurs, this time from a liquid to a gas. Again, at this temperature, the continued heating provides energy for the liquid-to-gas phase transition while the temperature remains the same. The heat, or enthalpy, required for the liquid-to-gas phase change is known as the **heat of vaporization**. It is worth noting that the above assumes the substance has both solid-to-liquid and liquid-to-gas transitions. Some substances convert directly from a solid to a gas when heated sufficiently, a process called **sublimation**. The heat needed for this solid-to-gas transition is called the **heat of sublimation**.

In the processes described above, note that there are two steps. In the first step, the substance is heated, generally by placing it in contact with another substance that is warmer (i.e., at a higher temperature, $T$). Energy is then transferred as heat from the warmer substance to the cooler substance. This continues until they reach the same temperature. If the two substances are in a well-insulated container such that minimal heat escapes to the surroundings, the energy change of the initially cooler substance can be calculated from its mass $m$, temperature change $\Delta T$, and heat capacity or specific heat $s$, which is $4.184 \text{ J/(g} \cdot \text{oC)}$ for water, according to the equation

$$q = ms\Delta T.$$  \hspace{1cm} (1)

If the system is indeed well-insulated, then the energy change of the second, initially warmer substance is assumed to be equal, but opposite in sign, to the energy change of first substance. In the second step, at the temperature of the phase transition, the energy transferred as heat goes into accomplishing the change of phase while the temperature remains constant. If it is the cooler substance undergoing the phase transition, the heat energy required for the phase change comes from the warmer substance, which changes temperature as it loses heat. Thus, for a substance undergoing a phase change at the transition temperature, the heat energy required can be obtained from the change in temperature of its surroundings. The energy involved with the phase change can be stated in units of energy/g or energy/mol (typically expressed as J/g or kJ/mol); the latter is the **molar** heat for the phase transition (see the Glossary). **Note the difference in**
In this way, the energy required to change the phase of a substance can be measured using a calorimeter by isolating the substance with, e.g., a warmer substance that provides the required heat. Measuring the drop in temperature of the warmer substance gives the heat required for the phase transition.

The energies involved with phase changes are:

<table>
<thead>
<tr>
<th>Heat of ...</th>
<th>Symbol</th>
<th>Phase Change</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fusion</td>
<td>$\Delta H_{\text{fus}}$</td>
<td>solid $\rightarrow$ liquid; melting</td>
</tr>
<tr>
<td>Vaporization</td>
<td>$\Delta H_{\text{vap}}$</td>
<td>liquid $\rightarrow$ gas; boiling</td>
</tr>
<tr>
<td>Sublimation</td>
<td>$\Delta H_{\text{sub}}$</td>
<td>solid $\rightarrow$ gas; sublimation</td>
</tr>
</tbody>
</table>

The magnitudes of these changes are characteristic for each substance under the conditions of the experiment.

The measurement of heat produced or consumed in a process is called calorimetry (i.e., the measurement of Calories, another unit of heat energy). It has many applications. For example, the energy content of foods can be measured and the energy density of fuels for heating or rocket propulsion can be determined.

[You might find it interesting that for many elements, the product of the specific heat times the atomic weight has approximately the same value! This was used to determine the approximate atomic weight for some elements. (Petit and Dulong, Annales de Chimie et de Physique 10 pg 395, 1819 and W. F. Magie, A Source Book in Physics, McGraw-Hill: NY, 1935, p. 178)]

Pre-lab

**Safety:** Goggles must be worn at all times. At atmospheric pressure liquid N$_2$ boils at -196°C and dry ice (solid CO$_2$) sublimes at -78°C. These substances are very cold and can quickly cause frostbite to exposed skin, so care must be taken in handling these substances.

Take care to keep the thermometer and Vernier temperature probe from direct contact with the liquid nitrogen and the dry ice.

**Pre-lab Assignment:** Please write out the following in your lab notebook. This assignment must be completed before the beginning of lab. You will not be allowed to start the experiment until this assignment has been completed and accepted by your TA.

1) Briefly describe the objectives of this experiment.

2) Write out the experimental procedure in your lab notebook according to the “Guidelines for Keeping a Laboratory Notebook” handout.
In addition to these pre-lab requirements, *a short quiz will be given at the beginning of lab* based on the material in this lab write-up.

**Procedure**

**Overview**

In this three-part experiment, a measured amount of warm water will be placed into a pair of nested Styrofoam cups, and the temperature will be measured with a glass thermometer. A weighed amount of a cold substance undergoing a phase transition will be added to the water and the temperature of the rapidly cooling water will be monitored and recorded using a Vernier temperature probe. Using the temperature change of the water, its mass, and its heat capacity or specific heat of 4.184 J/(g°C), the heat lost by the water - and thus the heat gained by the cold substance - can be determined. This energy change, together with the mass of the second, cold substance, can then be used to determine the heat associated with the phase change in J/g and kJ/mol.

The temperature range used today will be approximately 20–70 ºC. Temperature will be measured initially with a glass thermometer and then monitored with the Vernier temperature probe.

**Part 1 - The Heat of Fusion of Water (Ice)**

In this part of the experiment you will use a simple calorimeter to determine the heat of fusion of ice.

1. Confirm that the Vernier LabPro interface box is connected to the computer *via a USB cable*. A power supply line should also run from the LabPro box to an electrical outlet. The stainless steel temperature probe should be connected to the CH1 port in the LabPro box.

2. Launch the LoggerPro software by clicking the LoggerPro icon. Configure the software for data collection by opening file number 18 from the “Chemistry with Vernier” folder. If the real-time temperature reading is not displayed, consult the [online instructions](#) or ask your TA for help.

3. Obtain 2 pairs of nested styrofoam cups. Label each nested pair before measuring and recording their masses. In this way, you will be able to determine the mass of water and other substances added to them. Prepare a data table in your lab notebook like the one shown on the following page.
<table>
<thead>
<tr>
<th>QUANTITY</th>
<th>TRIAL 1</th>
<th>TRIAL 2</th>
<th>TRIAL 3</th>
</tr>
</thead>
<tbody>
<tr>
<td>m[H₂O(l)]</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>m[H₂O(s)]</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Tᵢ[H₂O(l)]</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Tᵢ[H₂O(l)]</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ΔT[H₂O(l)]</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

4. Heat approximately 250 mL of water to 55 ºC - 65 ºC.

5. When the water has been heated to the desired temperature, pour approximately 60 mL into one of the nested pairs and obtain and record its mass. Preheating or pre-cooling the nested cups with a small amount of the water is recommended to avoid heat loss from the water to the container.

6. With your second cup pair on the balance, add approximately 20 grams of ice and record the mass. (Take your glass thermometer to the balance with you so that you can measure the temperature of your warm water just before mixing it with ice in Step 7.)

7. After first confirming the actual temperature of your water, pour the ice into the nested cup pair holding the water.

8. Return to your work station and immediately place the Vernier temperature probe into the calorimeter. Click the “Collect” button at the top of the LoggerPro software window to initiate data collection. Stir the contents gently with the probe. When all the ice has melted and the temperature has reached a minimum and begins to rise, you may stop the data collection. By clicking on the STAT button, you can obtain the minimum temperature. *Each group must collect at least three data sets for the heat of fusion.* (Remember in your calculations that the water formed from the melted ice is being warmed!)

9. Find an average value for $\Delta H_{fus}$ of ice and express it as a heat of fusion, *i.e.*, in units of J/g, and also as a molar heat of fusion, *i.e.*, in units of kJ/mol. Add your results to the class data on the board. Record all the values from all the groups for use in your lab report.

**Part 2 - The Heat of Vaporization of Nitrogen**

In this part of the experiment you will use a simple calorimeter to determine the heat of vaporization of nitrogen.

1. Use the same pairs of cups as in Part 1, and be certain that they are thoroughly dry. Prepare a data table in your lab notebook analogous to the one in Part 1.
2. Heat approximately 250 mL of water to 55 °C - 65 °C as before.

3. When the water has been heated to the desired temperature, pour approximately 60 mL into one of the nested pairs. Obtain and record its mass as you did in Part 1. Again, preheating or pre-cooling the nested cups with a small amount of the water or liquid nitrogen is recommended to avoid heat loss from the water to the container as well as to reduce loss of liquid nitrogen to the atmosphere.

4. With your second cup pair on the balance, add approximately 40 grams of liquid nitrogen and record the mass. (Take your glass thermometer to the balance with you so that you can measure the temperature of your warm water just before mixing it with liquid nitrogen in Step 5.)

5. After first confirming the actual temperature of your water, pour the nitrogen into the nested cup pair holding the water. Liquid nitrogen changes phase quickly. In order to prevent significant loss of mass, make the transfer to the water as soon as you remove the cups containing the liquid nitrogen from the balance - do not wait until you get back to your bench. A cloud of extremely cold vapor will form above the container. Fan this away to avoid cooling the contents of the cup.

6. Return to your computer and immediately initiate data collection using the Vernier temperature probe as you did in Part 1. Stir the contents gently with the probe. When all the liquid nitrogen has evaporated and the temperature has reached a minimum and begins to rise, you may stop the data collection. By clicking on the STAT button, you can obtain the minimum temperature. Each team must collect at least three data sets for liquid nitrogen.

7. Find average values for $\Delta H_{\text{vap}}$ of liquid nitrogen and express as a heat of vaporization in units of J/g and as a molar heat of vaporization in kJ/mol. Add your results to the class data on the board. Record all the values for use in your lab report.

**Part 3 - The Heat of Sublimation of CO$_2$ (Dry Ice)**

In this part of the experiment you will use a simple calorimeter to determine the heat of sublimation of CO$_2$, otherwise known as dry ice.

1. Use the same pairs of cups as in Part 1 and be certain that they are thoroughly dry. Prepare a data table in your lab notebook analogous to the one in Parts 1 and 2.

2. Heat approximately 250 mL of water to 55 °C - 65 °C as before.

3. When the water has been heated to the desired temperature, pour approximately 60 mL into one of the nested pairs. Obtain and record its mass as you did in Part 1. Again, preheating or pre-cooling the nested cups with a small amount of the water, or liquid nitrogen is recommended to avoid heat loss from the water to the container.

4. With your second cup pair on the balance, add approximately 15 grams of crushed dry ice and record the mass. (Take your glass thermometer to the balance with you so that you can measure the temperature of your warm water just before mixing it with dry ice in Step 5.)
5. After first confirming the actual temperature of your water, pour the dry ice into the nested cup pair holding the water. Be sure that all the dry ice is transferred! In order to prevent significant loss of mass, make the transfer to the water as soon as you remove the cups containing the dry ice from the balance - do not wait until you get back to your bench. A cloud of extremely cold vapor will form above the container. Fan this away to avoid cooling the contents of the cup.

6. Return to your computer and immediately initiate data collection using the Vernier temperature probe as you did in Part 1. Stir the contents gently with the probe. When all the dry ice has sublimed and the temperature has reached a minimum and begins to rise, you may stop the data collection. By clicking on the STAT button, you can obtain the minimum temperature. Each team must collect at least three data sets for dry ice.

7. Find average values for $\Delta H_{sub}$ of CO$_2$ and express as a heat of sublimation in units of J/g and as a molar heat of sublimation in kJ/mol. Add your results to the class data on the board. Record all the values for use in your lab report.

**Discussion Points**

Of the three processes you investigated today, which involved the greatest amount of heat per mole for the phase changes involved?

Are the phase changes investigated exothermic or endothermic?

Which of the materials studied floated in water?

Does the amount of heat involved depend upon the amount of material being melted, vaporized, or sublimed?

Do you think the amount of heat involved depends upon the identity of the substance being melted, vaporized, or sublimed?

What intermolecular changes are occurring?

What average values for $\Delta H_{fus}$, $\Delta H_{sub}$, and $\Delta H_{vap}$ were obtained by your group? How do these compare to the class results?

Discuss possible sources of error in this experiment. How could you eliminate or minimize these sources of error?

**Report**

Your lab report should be a formal, individual report prepared according to the “Guidelines for Laboratory Reports” you have been given. In addition to the categories discussed in these guidelines you should provide answers to all the questions posed in this laboratory experiment write-up.
Glossary

**boiling point**
the temperature at which a liquid spontaneously transforms to a gas

**calorimeter**
an experimental apparatus used for measuring the heat evolved in a chemical process

**enthalpy**
a measure of the total thermodynamic energy of a system, represented as $H$; the enthalpy is the sum of the internal energy and the work involved in changing the system volume; changes in enthalpy, $\Delta H$, are the same as the heat transferred in a constant pressure process (a fact used in this laboratory experiment)

**heat of fusion, $\Delta H_{\text{fus}}$**
the change in enthalpy required (e.g., by the application of heat at constant pressure) to change a solid into a liquid; usually given as the enthalpy per mass of material, typically in units of J/g; related to the molar heat of fusion which is the enthalpy per mole of material, typically in kJ/mol

**heat of sublimation, $\Delta H_{\text{sub}}$**
the change in enthalpy required (e.g., by the application of heat at constant pressure) to change a solid into a gas; usually given as the enthalpy per mass of material, typically in units of J/g; related to the molar heat of sublimation which is the enthalpy per mole of material, typically in kJ/mol

**heat of vaporization, $\Delta H_{\text{vap}}$**
the change in enthalpy required (e.g., by the application of heat at constant pressure) to change a liquid into a gas; usually given as the enthalpy per mass of material, typically in units of J/g; related to the molar heat of vaporization which is the enthalpy per mole of material, typically in kJ/mol

**melting point**
the temperature at which a solid spontaneously transforms to a liquid

**phase transition**
a change from one phase of matter (e.g., solid, liquid, gas) to another

**specific heat**
the energy required to raise the temperature of a 1 g substance by 1 °C; depends on the substance; has units of J/(g°C); also called the specific heat capacity

**sublimation point**
the temperature at which a solid spontaneously transforms to a gas