Experiment 9

Specific Heat & Heats of Transformation

In this experiment you combine materials at different temperatures and wait until the mixture reaches a final common temperature at **thermal equilibrium.** You model the heat exchange process within the mixture using **calorimetry**, which is an application of conservation of energy using the idea that heat is a form of energy. The calorimetry analysis allows you to determine some of the thermal properties of the materials. The **specific heat capacity** is a measure of how heat added to / lost by a material results in a rise / fall of temperature. The **latent heat of transformation** is a measure of how heat added to / lost by a material results in a change of state, or **phase change**.

Preliminaries.

In most cases, when heat energy Q is added to an object initially at temperature T_i, the object's temperature increases to a final temperature T_f. The temperature rise depends on the mass m of the object and the type of material of which the object is composed. The relationship incorporating these ideas is

$$Q = mc\,(T_f - T_i) \qquad \text{(eq. 1)}$$

where c is the specific heat capacity of the material.

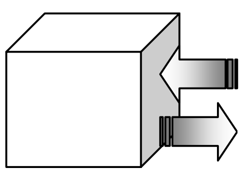

Figure 1: Heat Flow Sign Convention

Eq. 1 also describes the thermal behavior of an object when heat is removed. In this case, the final temperature is less than the initial temperature, giving a negative result for the heat Q. The sign of the heat energy Q is positive when heat is added and negative when heat is removed. This convention is shown in Figure 1.

In some cases, heat energy Q can be transferred to an object without having a temperature change occur. These cases occur when there is a phase change or change of state (examples are ice becoming water and dry ice becoming carbon dioxide vapor). The heat energy transferred in a phase change is directly proportional to the mass m changing its phase, the material and the type of phase change (boiling, melting, etc.). The relation is

$$Q = \pm\, mL \qquad \text{(eq. 2)}$$

where L is the latent heat of transformation appropriate for the type of phase change taking place. In using eq. 2, the sign must be chosen appropriately according to the sign convention discussed above. As an example, if the object is melting, heat is being added so that the plus sign is correct. The opposite is true if the object is freezing.

If objects at different temperatures are brought into contact in an insulated container they will transfer heat energy among themselves until thermal equilibrium is achieved. In thermal equilibrium, all temperatures equalize and remain constant thereafter. Hot objects cool off to the equilibrium temperature and the cold objects warm up. The hot objects lose heat energy and the cold ones gain heat energy. Objects may also change phase in their approach to thermal equilibrium, resulting in heat gains and losses. The heat energy gained and lost by the initially cold and hot objects can be expressed in terms of their masses, specific heats, latent heats, and temperature changes through eqs. 1 and 2. The net heat energy transfer for the system as a whole is zero *if the system is isolated*. This statement may be written symbolically by eq. 3, where Q_i represents one of the possible heat exchanges (due to different objects, changes in temperature, changes in phase) as the system progresses from its initial state to thermal equilibrium.

$$\sum Q_i = 0 \qquad \text{(eq. 3)}$$

This is a direct consequence of viewing heat as a form of energy and applying conservation of energy. The heat energy gains in parts of the system exactly balance the heat energy losses in other parts. Eq. 3 is a conservation of energy equation describing the processes in which the algebraic sum of all the relevant heat exchanges is zero.

The determination or use of thermal properties of the materials in a system consisting of objects exchanging heat is called calorimetry. Calorimetry experiments are generally difficult to analyze due to a source of error which is difficult to quantify – heat exchange between the system and its surroundings.

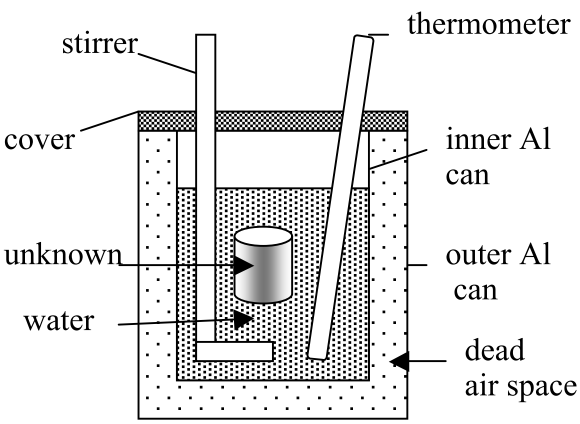

Figure 2: A Simple Calorimeter

Various steps can be taken to minimize this problem. First, the system can be as thermally insulated as possible. This is accomplished by placing the system in a **calorimeter**. The calorimeter used in these experiments is shown in Figure 2. The components of the system are assembled inside an aluminum cup set inside a bigger aluminum cup. The air space between the two provides the thermal insulation. The inner cup, however, is in thermal contact with its contents and *must be considered part of the system*. A lid on the container limits heat flow out the top. Second, the shorter the duration of the experiment, the less time for heat exchange. Once the components of the system are combined, the experiment needs to run as quickly as possible. Stirring of the system contents, using the stirrer shown in Figure 2, speeds the distribution of heat. This results in a more rapid approach to thermal equilibrium.

Procedure.

Caution: This lab uses boiling water and steam. Do not let the boiling water or steam touch your skin or clothes.

Part A. The Specific Heat Of Copper

• You can heat the copper by placing it some time in boiling water.

• Calculate the specific heat of copper and compare it to the accepted value

• Determine the mass of the inner aluminum cup m_{Al} of the calorimeter. How would this change your calculation?

Part B. The Latent Heat Of Fusion Of Ice

Devise a method to calculate how much heat is absorbed by ice in order to melt. Compare it to the accepted value.

Questions (Answer clearly and completely).

1. What value do you determine for the specific heat of copper? What is the percent difference from accepted?

2. One possible problem in Part A with the described procedure involves transferring the hot copper into the calorimeter. It may happen that drops of hot water are attached to the copper during the transfer. How would this affect the calculated value of the specific heat of copper? Would it be too big or too small? Why?

3. What value do you determine for the latent heat of fusion of ice? What is the percent difference from accepted?

4. Why does the procedure in Part B call for *warm* water? Is there an advantage to using above room temperature water? Explain carefully. [Hint: Why does the procedure of Part A call for cool water?]

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