SPECIFIC HEAT AND HEAT OF FUSION

PURPOSE
This laboratory consists of two separate experiments. The purpose of the first experiment is to measure the specific heat of two solids, copper and rock, with a technique known as the method of mixtures. The purpose of the second experiment is to measure the heat of fusion of ice.

SPECIFIC HEAT OF SOLIDS
The specific heat of a substance is defined as the energy per unit mass that is required to change the temperature of the substance by one degree (see your text). The term “heat” is used to describe energy in motion or being transferred, just as we use the term “wind” to describe air in motion or being transferred. The specific heat is usually given the symbol c. The unit of 'c' used in the laboratory is Cal g\(^{-1}\) C\(^{-1}\). In general, the specific heat of a substance depends on its temperature. The graph Figure 1 shows the temperature dependence of the specific heat (here given in J kg\(^{-1}\) K\(^{-1}\)) of zinc. Most solids, including the ones we will be using, have a similar temperature dependence.

Near room temperature (say 293 K) the specific heat varies only slightly with temperature, and so in this neighborhood (within ±25 K) the specific heat is nearly constant. In this experiment we shall neglect the small variations in the specific heat with temperature and treat all specific heats as constants independent of temperature.

The apparatus used to measure specific heats is a calorimeter as shown in Figure 2. It consists of two aluminum cups and a lid with a hole provided for a thermometer. The smaller of the two cups can be suspended within the other cup by means of an "O-ring" made of an insulating material.

The experiment is carried out by the method of mixtures. A known mass of water at room temperature is added to the inner cup. The amount of water must be such that it completely immerses the sample, whose specific heat is to be measured. The sample is placed within a boiling water bath so as to reach the boiling temperature of water. It is then quickly transferred to the inner cup and the lid is placed on the calorimeter. The mixture is gently stirred until the final equilibrium temperature is reached.
The temperature is monitored with a thermometer.

In order to determine the specific heat of the sample experimentally, we must make some physical assumptions. The calorimeter is specifically designed to limit heat losses to the environment. It is then natural to assume that no heat energy is transferred to the surroundings. This assumption allows us to consider energy transfers that occur only between the sample, the water, and the calorimeter. Is this assumption reasonable? We know that the calorimeter must leak some energy to the surroundings, for if we were to wait long enough, the entire system would reach room temperature. The amount of energy transfer to the surroundings depends on the temperature difference between the system and the surroundings.

There are three ways heat energy can escape from the system. These are convection, conduction, and thermal radiation. An isolated system, by definition, is immune to all three types of energy transfer. We need to minimize all these to obtain accurate answers.

Heat transfer by convection relies on convection currents that physically carry material (in this case heated air) from one region to another. Since the calorimeter is covered with a lid, heat transfer due to convection can be minimized.

The layer of air between the inner and outer cups provides thermal insulation for minimizing conduction.

The polished aluminum cups are intended to provide mirror-like boundaries to reflect thermal radiation back into the system, reducing the radiation losses. The radiation effect can also be minimized by starting the original temperature of the mixture as far below room temperature as the final equilibrium temperature is above room temperature. (The idea here is that radiation would add heat at the start, but at the end of the experiment the system will lose heat by radiation. Therefore, the heat added at the start would be close to the heat lost at the end. Therefore, the radiation losses would be further minimized.) For the substances involved in this experiment and initially heated to the temperature of boiling water, a good estimate for the initial temperature of calorimeter and water is about 10°C (this may vary) below the room temperature since, the final equilibrium temperature is expected to be about 10°C above the room temperature. The choice of such an initial temperature will reduce the radiation loss.
For our calculations, we can write specific variables as follows:

\[ T_i = \text{initial temperature of the water and calorimeter (inner cup + stirrer) - (approximately 10}^\circ\text{C below room temperature)} \]

\[ T_s = \text{initial temperature of the sample (approximately 100}^\circ\text{C)} \]

\[ T_e = \text{equilibrium temperature of the sample, water, and calorimeter} \]

\[ m_w = \text{mass of water in g} \]

\[ m_c = \text{mass of calorimeter (inner cup + stirrer) in g} \]

\[ m_s = \text{mass of sample in g} \]

\[ c_w = \text{specific heat of water} = 1 \text{ Cal g}^{-1} \circ\text{C}^{-1} \]

\[ c_{AI} = \text{specific heat of aluminum} = 0.221 \text{ Cal g}^{-1} \circ\text{C}^{-1} \]

\[ c = \text{unknown specific heat of sample} \]

We give you names for the variables so you can be specific in the following calculation. We can equate the heat energy gained by the water and the calorimeter with the heat energy lost by the sample. We do so by using the equation

\[ \Delta Q = m \cdot c \cdot \Delta T \]

where \( \Delta Q \) is the heat transferred, \( m \) is the mass, \( c \) is the specific heat, and \( \Delta T \) is the temperature change. The challenge as you work this out is to correctly identify the mass and \( \Delta T \) for a given part of the experiment.

The energy transfer to the thermometer is ignored in the above analysis, even though it is obvious that the thermometer can’t work without energy transfer to it. The thermometer presents a problem because part of it is interacting directly with the system while the rest is completely outside of the system. There are (slightly complicated) ways to include the effect of the thermometer but we shall not pursue them here. It is useful, though, to ask in what direction this error would push the result. Would the specific heat seem to be too big, or too small?
HEAT OF FUSION OF ICE

The heat of fusion of a substance is defined as the energy per unit mass required to change the phase of the substance from solid to liquid. Knowledge of the heat of fusion of a substance can give valuable information about its microscopic properties. For instance, a large heat of fusion implies that the bonds between the molecules which make up the substance are strong, while small values indicate weak bonding. The heat of fusion is an example of a Latent Heat, so named because the heat is not changing temperature, so its effect is in some way “Latent” or hidden.

In this experiment we will measure the heat of fusion of water using the calorimeter. A known mass of water is added to the inner cup of the calorimeter such that its temperature is about 10°C above room temperature. A few ice cubes are added to the cup. The ice cubes should initially be at the melting temperature. How can we ensure this? The object is to add enough ice to produce a significant temperature change in the system yet not so much ice that it does not all melt. You can experiment with different masses of ice to find an optimum amount or you can find out approximately how much ice to add by using the textbook value for the heat of fusion and working the problem backwards. The ice should be blotted with a paper towel to remove all moisture on it before adding it to the cup. Also, care must be taken not to splash water from the cup, while adding ice pieces. (Can you see why?)

The system is covered with the lid and stirred gently as before until the temperature reaches its lowest (equilibrium) value. The analysis of this problem resembles that of the specific heat experiment.

The variables are labelled as follows:

\[ T_i = \text{initial temperature of the water and calorimeter (inner cup + stirrer)} \]
\[ \text{(approx.} 10^\circ\text{C above room temperature)} \]
\[ T_m = \text{melting temperature of the ice (approx.} 0^\circ\text{C)} \]
\[ T_e = \text{equilibrium temperature of the system} \]
\[ m_w = \text{mass of water in g} \]
\[ m_c = \text{mass of calorimeter (inner cup + stirrer) in g} \]
\[ m_l = \text{mass of ice in g} \]
\[ c_w = \text{specific heat of water} = 1 \text{ Cal g}^{-1} \text{ } ^\circ\text{C}^{-1} \]
\[ c_{Al} = \text{specific heat of aluminum} = 0.22 \text{ Cal g}^{-1} \text{ } ^\circ\text{C}^{-1} \]
\[ L = \text{unknown heat of fusion of ice} \]
Under the condition that all the ice melts we can equate the heat energy lost by water and calorimeter to the heat energy gained by the ice plus the heat energy gained by the water from the melted ice. We can write the basic equation for melting as

\[ \Delta Q = L \Delta m \]

where \( \Delta Q \) is the energy transferred (i.e. the heat) that melts a mass \( \Delta m \), with a heat of fusion of \( L \). As above, you will equate heat transferred to deduce \( L \). The problem is now trickier:

The water lost heat \( \Delta Q = mc\Delta T \). The ice gained heat in the amount \( \Delta Q = L \Delta m \), and then water created from the melting ice, plus the now goes up by a heat \( \Delta Q = mc\Delta T \). The water melted from ice, and the water we put the ice into, both end up at the same \( T \).

We are leaving it to you write all the math out correctly. Have you lab instructor check it. Learning how to do such a calculation is one of the educational goals of the lab.
EXPERIMENTAL PROCEDURE

A. DETERMINATION OF SPECIFIC HEATS

1. Weigh the pieces of copper and rock separately on the side arm balance to the nearest gram and enter their masses on your data sheet.

2. Suspend the copper piece in boiling water.

3. Weigh the calorimeter (inner cup with the stirrer) and note its mass in the Table 1.

4. Fill the calorimeter about two-thirds with cold water at a temperature 10°C below room temperature and weigh it again with the stirrer. Find the mass of water added to the calorimeter.

5. When the copper piece is in boiling water for five minutes or more, note the temperature of the sample \( T_s \), that of boiling water, and enter this in Table 1.

6. You may have noted a slow rise in the temperature of the cold water and the calorimeter while waiting for the copper piece to reach the temperature of boiling water. This is due to absorption of radiated heat from the surroundings. Note the initial temperature of calorimeter and water, just before transferring the copper piece and enter this value \( T_i \) in the Table 1.

7. Transfer the copper piece from boiling water into the calorimeter as quickly as you can. Close the lid and stir the mixture carefully without spilling any water outside the calorimeter. Note the highest temperature recorded by the mixture. Allow sufficient time to make sure that the mixture has reached the highest temperature. The temperature will decrease after reaching this highest temperature due to radiation losses.

8. Compute the heat gained by the water and the calorimeter. (What is the \( \Delta T \) you are using?)
   Equate the heat gained above with the heat lost by the copper sample.
   What is the \( \Delta T \) for the sample?
   Using the heat capacity equation, find the heat capacity.

8. Repeat steps 2-7 for the rock sample. Be sure to replace the water in the calorimeter each time with cold water as mentioned in step 4.
B. HEAT OF FUSION OF ICE:

10. Weigh the calorimeter (including the stirrer) making sure that they are dry. Enter its mass in Table 2.

11. Fill the calorimeter two-thirds with warm water at a temperature about 10°C above room temperature. Weigh the calorimeter and its contents and enter the mass in Table 2.

12. Note the initial temperature ($T_i$) of calorimeter and water and also the melting temperature ($T_m$) of ice by inserting the thermometer in the ice bath and enter these values in Table 2.

13. Add small pieces of ice (about five or six pieces should be sufficient) to water in the calorimeter. Make certain that no moisture is on the ice by drying it with a paper towel.

14. Stir the contents carefully by keeping the ice under water. Record the final temperature ($T_f$), after all the pieces of ice have melted.

15. Weigh the calorimeter and its contents (stirrer, water and melted ice) and determine the mass of ice melted.
# TABLE I

## SPECIFIC HEAT OF COPPER AND ROCK

<table>
<thead>
<tr>
<th></th>
<th>COPPER</th>
<th>ROCK</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass ($m_a$)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of Calorimeter ($m_c$)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of Calorimeter + water</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of water ($m_w$)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Initial temperature ($T_s$) of the sample</td>
<td>°C</td>
<td>°C</td>
</tr>
<tr>
<td>Initial temperature ($T_i$) of Calorimeter+water</td>
<td>°C</td>
<td>°C</td>
</tr>
<tr>
<td>Final temperature ($T_e$) of the mixture</td>
<td>°C</td>
<td>°C</td>
</tr>
<tr>
<td>Specific heat ($c_{Al}$) of calorimeter</td>
<td>Cal g$^{-1}$°C$^{-1}$</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th></th>
<th>COPPER</th>
<th>ROCK</th>
</tr>
</thead>
<tbody>
<tr>
<td>Specific heat in Cal g$^{-1}$°C$^{-1}$</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
**TABLE 2**

**HEAT OF FUSION OF ICE**

<table>
<thead>
<tr>
<th>Description</th>
<th>Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of calorimeter ((m_c))</td>
<td></td>
</tr>
<tr>
<td>Mass of calorimeter + water</td>
<td></td>
</tr>
<tr>
<td>Mass of water ((m_w))</td>
<td></td>
</tr>
<tr>
<td>Initial temperature ((T_i)) of calorimeter(with stirrer) + water</td>
<td></td>
</tr>
<tr>
<td>Melting temperature ((T_m)) of ice</td>
<td></td>
</tr>
<tr>
<td>Final Temperature ((T_e)) of mixture</td>
<td></td>
</tr>
<tr>
<td>Mass of calorimeter + water + melted ice</td>
<td></td>
</tr>
<tr>
<td>Mass of ice ((m_I))</td>
<td></td>
</tr>
<tr>
<td>Specific heat ((c_{Al})) of calorimeter</td>
<td></td>
</tr>
</tbody>
</table>

Compute the heat lost by the water and the calorimeter.  (What is the \(\Delta T\)?)

Some of this heat went into warming the water that was formed when the ice melted. How much heat is required for warming the melted ice?  (What is the \(m\), and what is the \(\Delta T\) used here?)

Therefore, how much heat is left to go into melting the ice?

From that number, calculate the Latent heat (heat of fusion) of ice melting.

**Heat of fusion of ice**  \(L = \) ___________ Cal g\(^{-1}\)