

Title: Specific Heat and Latent Heat

Object: To determine the specific heat of various substances and the latent heat of melting for ice.

Introduction: *Specific heat,  $c$* , is an indication of how much the temperature of an object will change when internal energy is increased or decreased. Specific heat varies from substance to substance which means that two objects of the same mass but of different substances will change temperatures by different amounts when their internal energies are increased or decreased by the same amount. For example, have you ever held a hot piece of pizza by its crust only to have the moister parts burn your mouth when you take a bite? The meats and cheeses have high specific heat, while the crust has a low specific heat. The specific heat of a substance would be the slope of a graph of the change in energy vs. the mass of the object times the change in temperature. ( $E$  has been used for other kinds of energy, so we'll use  $U$  specifically for internal energy though many textbooks use  $Q$ .)

$$\Delta U = cm\Delta T \quad (1)$$

Therefore the specific heat of a substance can be calculated by measuring the change in temperature of an object which corresponds to a known change in the internal energy of the object (such as by doing mechanical work on the object).

Specific heats of different substances can be compared by measuring temperature changes as one object heats another. For example, if a sample of hot metal is dropped into some cold water then they both arrive at a common temperature somewhere between the extremes. By using a good calorimeter to minimize interaction with the air in the room we can assume that the decrease in internal energy of the metal is equal to the increase in internal energy of the water.

$$-\Delta U_m = \Delta U_w \quad (2)$$

$$-c_m m_m \Delta T_m = c_w m_w \Delta T_w \quad (3)$$

We'll first do some mechanical work on some lead and assume the work increases the internal energy of the lead, and then use that assumption to calculate the specific heat of lead. Then we'll use our new knowledge (and equation 3) to calculate the specific heat of water; and then use the specific heat of water (and equation 3) to calculate the specific heat of a different metal.

*Latent heat,  $L$* , is the amount of energy (per unit mass) needed to effect a change of state at a constant temperature. For example, the latent heat of melting (or fusion),  $L_f$ , for ice is the amount of energy (per unit mass) needed to change the ice at  $0^\circ\text{C}$  to water at  $0^\circ\text{C}$ ; the accepted value is 79.8 cal/g. Likewise, the latent heat of vaporization,  $L_v$ , for water is the energy (per unit mass) needed to change the liquid water at the boiling temperature to the gaseous state at that temperature; the accepted value is 539 cal/g. We can modify equation 3 to handle a change of state. Specifically, when ice cubes at  $0^\circ\text{C}$  are put in warm water:

$$-c_w m_w \Delta T_w = m_{\text{ice}}(L_f)_{\text{ice}} + c_{\text{melted ice}} m_{\text{ice}} \Delta T_{\text{melted ice}} \quad (4)$$

Procedure: Both lead and mercury are poisonous so don't eat the lead and don't break the thermometers. Wash your hands after lab.

Using equation 1:

1. Fill a plastic bottle about 1/3 full of lead shot and find the mass of the lead.
2. Measure the initial temperature (in °C) of the lead shot with a sensitive thermometer.
3. Drop the bottle  $n$  times (about 20 or 30) in quick succession from the lab bench to the floor.
4. Immediately measure the final temperature of your lead shot (the rise won't be much). Record the highest reading the thermometer reaches.
5. Calculate the mechanical work done from  $W = n F d$  where  $F$  is the weight (in newtons) of the shot and  $d$  is the distance (in meters) from the floor to the lab bench. The result will be in units of Nm; and one Nm is equal to one joule. Now convert from joules to calories.
6. Make a plot of change in internal energy (assumed to be equal to the work done) (in calories) vs. mass times the change in temperature (in g °C). Your lead shot sample will provide one data point; also acquire the data from the other groups in the class on this procedure for other data points to plot. It is important to realize that specific heat varies with temperature so this graph won't be a perfectly straight line. But over small temperature ranges  $c$  doesn't change much and so the graph will be approximately linear; therefore use a ruler to eyeball a best-fit straight line.
7. The slope of your straight line is your experimental value for the specific heat of lead (in cal/g °C) near room temperature; compare it to the accepted value.

Using equation 3:

1. Measure the mass of a solid piece of lead or a quantity of lead shot.
2. Suspend the solid lead by a string or the shot in a cup in the steam generator to raise its temperature to the temperature of boiling water (not necessarily 100 °C at our altitude).
3. Put a quantity (about 1/2 the mass of your lead sample) of cold water in the calorimeter and measure its mass and initial temperature. Include 12.2 grams to account for the effect of the aluminum calorimeter cup.
4. Record the initial hot temperature of the lead.
5. Quickly dry the lead (if you used the solid) and put it in the cold water in the calorimeter.
6. As they equilibrate (gently use your stirring rod occasionally), record the highest water temperature as the final temperature for both the lead and the water.
7. Use these data along with the specific heat for lead (use the accepted value) to compute the specific heat for water (in cal/g °C) from equation 3.

Now that you have found the specific heat for water, use the procedure above (for equation 3) to find the specific heat of a different solid metal sample, such as aluminum or steel or brass.

Using equation 4:

1. Fill your calorimeter cup one-half to two-thirds full of warm water and determine the mass of the water. Add 12.2 g for the cup.
2. Record the initial temperature of the water. Obtain a few ice cubes from a water-and-ice mixture (this ensures that they are very close to  $0^{\circ}\text{C}$ ) and carefully dry them with paper towels. Then put them in the calorimeter, being careful not to splash water out, and gently stir until completely melted. You would like to end as much below room temperature as you started above. Record the final temperature as the lowest reading after the ice is completely melted.
3. The mass of the ice is most easily determined by re-weighing the cup now and subtracting the mass of the cup (the true mass, not the 12.2 g) and the original warm water.
4. Use equation 4 to compute the latent heat of ice. (Use the accepted value for  $c_w$ .)

Apparatus: Sketch the apparatus.



Using equation 3:

Mass of lead (in grams):  $m_{Pb} =$  \_\_\_\_\_

Initial Temp.:  $T_{Pb_i} =$  \_\_\_\_\_ Final Temp.:  $T_{Pb_f} =$  \_\_\_\_\_  $\Delta T_{Pb} =$  \_\_\_\_\_

Mass of water (add 12.2 g to account for the calorimeter cup):  $m_w =$  \_\_\_\_\_

Initial Temp.:  $T_{w_i} =$  \_\_\_\_\_ Final Temp.:  $T_{w_f} =$  \_\_\_\_\_  $\Delta T_w =$  \_\_\_\_\_

Specific heat of water (solve equation 3 for  $c_w$ ):  $c_w =$  \_\_\_\_\_

Compare with accepted value:  $c_w =$  \_\_\_\_\_ % error = \_\_\_\_\_

Using equation 3 for a different metal (metal is \_\_\_\_\_) :

Mass of water (add 12.2 g to account for the calorimeter cup):  $m_w =$  \_\_\_\_\_

Initial Temp.:  $T_{w_i} =$  \_\_\_\_\_ Final Temp.:  $T_{w_f} =$  \_\_\_\_\_  $\Delta T_w =$  \_\_\_\_\_

Mass of metal (in grams):  $m_m =$  \_\_\_\_\_

Initial Temp.:  $T_{m_i} =$  \_\_\_\_\_ Final Temp.:  $T_{m_f} =$  \_\_\_\_\_  $\Delta T_m =$  \_\_\_\_\_

Specific heat of metal (solve equation 3 for  $c_m$ ):  $c_m =$  \_\_\_\_\_

Compare with accepted value:  $c_m =$  \_\_\_\_\_ % error = \_\_\_\_\_

Using equation 4 (for latent heat of ice):

Mass of warm water (add 12.2 g to account for the calorimeter cup):  $m_w =$  \_\_\_\_\_

Initial Temp.:  $T_{w_i} =$  \_\_\_\_\_ Final Temp.:  $T_{w_f} =$  \_\_\_\_\_  $\Delta T_w =$  \_\_\_\_\_

Initial Temp.:  $T_{ice_i} =$  0°C Final Temp.:  $T_{w_f} =$  \_\_\_\_\_  $\Delta T_{\text{melted ice}} =$  \_\_\_\_\_

Mass of ice (in grams):  $m_{ice} =$  \_\_\_\_\_

Latent heat of ice (solve equation 4 for  $(L_f)_{ice}$ ):  $(L_f)_{ice} =$  \_\_\_\_\_

Compare with accepted value:  $(L_f)_{ice} =$  \_\_\_\_\_ % error = \_\_\_\_\_

Questions:

1. Why might the results in the first procedure not be very accurate? (Does *all* of the work change the internal energy of the *lead*?)
2. How could one improve the results of this whole experiment?
3. Water is often used as a heat storage medium. Is there a physics reason for this, or is it just cheap and plentiful? What property of water determines the answer to this question? How unique is water with respect to this property when compared to nearly all other liquids and solids?
4. If you dropped a block of aluminum from the top of a building, would its temperature rise more or less than if you dropped a block of lead? Explain your reasoning.
5. How many calories would it take to raise 100.0 g of ice at  $-5.0^{\circ}\text{C}$  to  $+5.0^{\circ}\text{C}$ ? Show your work. There will be a term in your answer that wasn't in equation 4 because in the procedure you were careful to start with ice at  $0.0^{\circ}\text{C}$ ; for that term you will need to look up  $c_{\text{ice}}$ .

Conclusions: