Lab #10 Thermal Physics

Purpose: To determine the specific heat of metal samples, and to determine the heat of fusion for ice.

Equipment:

Calorimeter	Triple beam balance
Thermometer	Several metal samples
Hot plates	Containers for water
Tongs	Ice at < 0° C

Discussion:

Part One, Specific Heat

As a form of energy, heat has the same units as work or energy; ft-lbs., n • m (or joules), calories (*), BTUs, KWHR, etc. How much heat (Q) it takes to warm an object on three things: how big it is (mass), how many degrees its temperature is to be raised (°C), and from what material it is made. How much heat it takes to warm an amount of some material 1.0 °C is called its heat capacity. A much more meaningful physical characteristic of a material is the amount of heat it takes PER UNIT MASS to warm the material by 1.0 °C. This is known as the specific heat capacity, or, more commonly, specific heat, usually identified by lower-case c. Its units are calories / gram • °C or joules/ kg • °C.

$$\text{Amount of heat} = Q = m \cdot c \cdot \Delta T$$

Water, for example, has a very high specific heat: 4,186 J/kg • °C or 1.0 cal/g • °C. So it would take:

1 cal to heat 1 gram of water 1 °C
10 cal to heat 1 gram of water 10 °C
10 cal to heat 10 grams of water 1 °C
100 cal to heat 10 grams of water 10 °C

Copper has a specific heat of 387 J/kg • °C, so to heat up 500 g of copper from 23 °C to 35 °C would require:

$$Q = m \cdot c \cdot \Delta T = (.5\text{kg}) \cdot (387\ \text{J/kg} \cdot {}^\circ\text{C}) \cdot (12.0\ {}^\circ\text{C}) = 2340\ \text{J of heat.}$$

The conversion factor between joules and calories is:

1 calorie = 4.186 joules

*Note that physics calories are 1/1000 of a nutritional Calorie, or Kcal.

In an insulated heat transfer process; the heat lost by the warmer objects is transferred to (gained by) the cooler ones. For today's laboratory exercise, this means:

Ht. lost by sample + Ht. gained by water + Ht. gained by cup = 0

OR: $(m \cdot c \cdot \Delta T)_{sample} = (m \cdot c \cdot \Delta T)_{water} + (m \cdot c \cdot \Delta T)_{cup} = 0$

When the equation is written this way, we must be careful with the signs of the $\Delta T = T_f - T_i$ = positive for an object that is warmed, but negative for an object that is cooled. (This goes along with the normally accepted convention that heat gained is considered positive and heat lost is negative.) To avoid this consideration, we may rewrite the equation above as follows:

Heat lost by hot things must = heat gained by cold things

OR: $(m \cdot c \cdot \Delta T)_{sample} = (m \cdot c \cdot \Delta T)_{water} + (m \cdot c \cdot \Delta T)_{cup}$

(Now we need only consider the abs. value of each temp. change.)
In this lab, you will know 8 of the 9 elements in this equation and can solve it for the specific heat of the metal sample.

Procedure:

1. Find the mass of the inner cup of the calorimeter.

2. Fill the cup about 1/3 full with water from the tap. Find the mass of the water and cup; then determine the mass of the water.

3. Assemble the inner cup into the outer cup, put the thermometer in the rubber stopper and into the plastic cover, making certain that its tip is immersed in the water. Record the room temperature of the water and inner cup combination (initial temperature). <u>TRY</u> TO READ THE THERMOMETER TO THE NEAREST <u>TENTH</u> OF A DEGREE!

4. Using tongs, very quickly (to minimize heat loss) transfer one of the metal samples to the inner cup and cover quickly. Gently agitate the calorimeter in a horizontal circular motion and record the temperature of the water and cup after all three (water, cup, sample) achieve thermal equilibrium (final temperature).

5. Determine the mass of the now-cool metal sample, and calculate its specific heat.

6. Repeat this process (steps 3-5) for a different type metal.

7. Compare your findings with the accepted values for the specific heat of the specimens listed below. Determine percent error.

Lead (Pb)	128 J/kg • °C	Tin (Sn)	210 J/kg • °C
Alum. (Al)	900 J/kg • °C	Zinc (Zn)	375 J/kg • °C
Copper (Cu)	387 J/kg • °C	Water	4186 J/kg • °C

DATA (Part 1)

Type of metal sample	________	________
Mass of cup (kg)	________	________
Mass of water + cup (kg)	________	________
Initial temp. of water & cup (°C)	________	________
Initial temp. of metal sample (°C)	________	________
Final temp. of water, cup, & sample (°C)	________	________
Mass of sample (kg)	________	________
Change in temp. of cup (°C)	________	________
Change in temp. of water (°C)	________	________
Change in temp. of sample (°C)	________	________
Specific heat of sample (J/kg • °C)	________	________

CALCULATIONS: (Clearly show all eight numbers used, incl. UNITS!)

Discussion: Part Two, Heat of Fusion

When any substance is changed from a solid to a liquid or from a liquid to a gas (or the reverse), then the substance undergoes a PHASE CHANGE, and a certain amount of heat per unit mass is transferred in order to make this change. For example, when a solid melts, this phase change takes place at the material's melting point and the amount of heat absorbed by the solid per unit mass to cause this change is called the heat of fusion (L_f). it is the same amount of heat per unit mass that is lost by a liquid when, after being cooled to its freezing point, it solidifies. Similarly, if a material changes phase from a liquid to a gas, the amount of heat absorbed per unit mass is called the heat of vaporization (L_v), and the temperature at which this takes place is the boiling point of the material. Again, this heat of vaporization is also equal to the amount of heat lost by a gas when, after being cooled to its boiling point, it condenses. Heats of fusion/vaporization have units of joules per kilogram, or calories per gram:

$$Q \text{ (phase change)} = m \cdot L_f \qquad \text{or} \qquad Q = m \cdot L_v$$

In today's laboratory exercise, heat goes from the room-temperature water and cup into (a) warming the ice to 0 °C, (b) changing the phase of the ice at 0 °C to water at 0 °C, and finally (c) warming the "new water" from 0 °C to the final equilibrium temperature:

Heat lost by hot things = Heat gained by cold things, OR:

Ht. lost (cup) + Ht. lost (water) = Ht. gained (warm ice to 0 °C) + Ht. gained (melt ice)

+ Ht. gained (warm "new water")

$$(m \cdot c \cdot \Delta T)_{cup} + (m \cdot c \cdot \Delta T)_{water} = (m \cdot c \cdot \Delta T)_{warm\ ice} + (m \cdot L_f)_{melt\ ice}$$

$$+ (m \cdot c \cdot \Delta T)_{warm\ \text{"new water"}}$$

Knowing the masses of the water, cup, and ice used in this procedure, their temperature changes, and the specific heats for water, ice, and the cup, it is a simple algebraic exercise to determine the heat of fusion for ice. (We know 13 of the 14 elements of the equation above).

Procedure:

1. Recall and record the mass of the inner cup.

2. Fill the cup at least _ full of tap water. Determine and record the mass of the water and cup to find the mass of the water.

3. Assemble the inner cup and the insulation ring into the outer cup. Place thermometer in the rubber stopper and into the plastic lid, making sure that the thermometer tip will be immersed under water when the lid is in place. Record the initial temperature. Again, reading the thermometer ACCURATELY is absolutely critical.

4. Using tongs, quickly transfer a specified amount of ice into the inner cup and cover. Gently agitate the calorimeter until all the ice is melted. When the temperature stops dropping, then eventually begins to rise slowly, record the minimum as final temperature.

5. Determine and record the total mass of the cup, original water, and melted ice. Calculate the mass of “new water” (which equals the mass of the ice you put into the cup).

6. Calculate the heat of fusion for ice, and compare it to its accepted value of 335, 000 J/kg, and compute % error.

DATA (Part 2)

Mass of cup	____________________	kg
Mass of water and cup	____________________	kg
Mass of water	____________________	kg
Mass of water, “new” water, & cup	____________________	kg
Mass of ice	____________________	kg
Initial temp. of water and cup	____________________	°C
Initial temp. of ice	____________________	°C
Final equilib. temp. of all water & ice	____________________	°C
Change in temp. of water and cup	____________________	°C
Change in temp. of ice	____________________	°C
Change in temp. of “new water”	____________________	°C

CALCULATIONS: (Clearly show all 13 numbers used, including UNITS!)