Prelab #2
Coffee Cup Calorimetry

Heat is a form of energy, sometimes called thermal energy that will pass *spontaneously* from an object at a high temperature to an object at a lower temperature. If the two objects are in contact they will, given sufficient time, both reach the same temperature. Heat flow is ordinarily measured in a device called a calorimeter. A calorimeter is simply a container with insulating walls, made so that essentially no heat is exchanged between the contents of the calorimeter and the surroundings. A coffee cup or thermos is an example of containers designed to prevent heat from escaping its contents. Within the calorimeter chemical reactions may occur or heat may pass from one part of the contents to another, but no heat flows into or out of the calorimeter from or to the surroundings.

A. Specific Heat.

When heat energy flows into a substance, the temperature of that substance will increase. The quantity of heat energy ($Q$) required to cause a temperature change in any substance is equal to the specific heat ($S$) of that particular substance times the mass ($m$) of the substance times temperature change ($\Delta T$), as shown in Equation 1.

$$Q = S \times m \times (\Delta T)$$

Equation (1)

The specific heat can be considered to be the amount of heat required to raise the temperature of one gram of the substance by one degree Celsius. Amounts of heat energy are measured in either joules or calories. To raise the temperature of 1 gram of water by 1 degree Celsius, 4.184 joules of heat must be added to the water. The specific heat of water is therefore 4.184 joules/g°C. Since 4.184 joules equals one calorie, we can also say that the specific heat of water is 1 calorie/g°C. That is not a coincidence, this is how the calorie was originally defined- the amount of energy needed to raise 1g of water by 1°C!

B. Calorimetry

The specific heat of a metal can readily be measured in a calorimeter. A calorimeter is a container with a known heat capacity. When a heating or cooling process in a calorimeter occurs, the amount of energy (heat) gained or lost in the process is absorbed by the calorimeter. This energy can be calculated from equation (1) if the calorimeter’s change in temperature is measured, since its heat capacity ($S$) and mass ($m$) are known. Very often, a calorimeter is composed of pure water in an insulating container. If one assumes that a negligible amount of heat escapes the insulated container and there is a negligible change in temperature of the container during the process, then all the energy released during the process of interest goes in to the water. Since we know the heat capacity and the mass of the water used in calorimeter and can measure the temperature change of the water, then we can use equation (1) to calculate the energy gained or lost by the water. The key point is that the energy gained (or lost) by the water is equal to the energy lost (or gained) in the process (such as cooling of the metal) that occurs in the calorimeter.
A pre-weighed amount of metal is heated to some known temperature and is then quickly transferred into a calorimeter that contains a measured amount of water at a known temperature. Energy (heat) flows from the metal to the water, and the two eventually equilibrate (come to the same temperature) at some temperature between the initial temperatures of the water and metal. Assuming that no heat is lost from the calorimeter to the surroundings, and that a negligible amount of energy is absorbed by the calorimeter walls, the amount of energy that flows from the metal as it cools is equal to the amount of energy absorbed by the water. In other words, the energy that the metal loses is equal to the energy that the water gains. Since the metal loses energy (T\text{final} is less than T\text{initial}; therefore \Delta T is negative) Q\text{metal} is negative. The water in the calorimeter gains energy (T\text{final} is greater than T\text{initial}; therefore \Delta T is positive) and Q\text{water} is positive. Since the total energy is always conserved (cannot disappear), we can write equation (2):

\[ Q_{\text{metal}} + Q_{\text{water}} = 0 \quad \text{equation (2)} \]

Rearranging equation (2) gives (note the negative sign):

\[ Q_{\text{metal}} = -Q_{\text{water}} \quad \text{equation (3)} \]

In this experiment we measure the mass of water in the calorimeter and mass of the unknown metal and their initial and final temperatures. Using equation (1), the heat energy gained by the water and lost by the metal can be written:

\[ Q_{\text{water}} = S_{\text{water}} \times m_{\text{water}} \times \Delta T_{\text{water}} \quad \text{equation (4)} \]

\[ Q_{\text{metal}} = S_{\text{metal}} \times m_{\text{metal}} \times \Delta T_{\text{metal}} \quad \text{equation (5)} \]

(Note that \(\Delta T_{\text{metal}}\) < 0 and \(\Delta T_{\text{water}}\) > 0, since \(\Delta T = T_{\text{final}} - T_{\text{initial}}\))

We can calculate \(Q_{\text{water}}\) from the experimental data; we measured \(m_{\text{water}}, \Delta T_{\text{water}},\) and we know \(S_{\text{water}}= 4.184 \text{ J/g}^\circ\text{C}\). We know that the heat energy lost by the metal is equal (but opposite sign) as the heat energy gained by the water. To determine the specific heat of your unknown metal (\(S_{\text{metal}}\)), substitute \(Q_{\text{metal}}\) (in equation (5)) with \(-Q_{\text{water}}\):

\[ -Q_{\text{water}} = S_{\text{metal}} \times m_{\text{metal}} \times \Delta T_{\text{metal}} \quad \text{equation (6)} \]

In this equation, we know 3 of the 4 variables: we calculated \(Q_{\text{water}}\) from experimental data; we measured \(m_{\text{metal}}\) and \(\Delta T_{\text{metal}}\). We can now solve equation (6) for \(S_{\text{metal}}\). Equation (6) can be rearranged to give \(S_{\text{metal}}\) as a function of the known variables (\(Q_{\text{water}}, m_{\text{metal}},\) and \(\Delta T_{\text{metal}}\)):

\[ S_{\text{metal}} = \frac{-Q_{\text{water}}}{m_{\text{metal}} \times \Delta T_{\text{metal}}} \quad \text{equation (7)} \]

You will use this procedure to obtain the specific heat of an unknown metal. You will then compare the specific heat of your unknown metal to a table containing values of specific heats for several metals in order to determine the identity of your metal.
Prelab Questions:

1) Why can we substitute \( Q_{\text{metal}} \) (in equation (5)) with \( -Q_{\text{water}} \)?

2) Why is \( \Delta T_{\text{metal}} < 0 \)?

3) Why is \( \Delta T_{\text{water}} > 0 \)?

4) Should \( Q_{\text{metal}} \) be positive or negative? Why?

5) Should \( Q_{\text{water}} \) be positive or negative? Why?

6) A metal sample weighing 45.2 g and at a temperature of 100.0°C was placed in 38.6 g of water in a calorimeter at 25.2°C. At equilibrium the temperature of the water and metal was 32.4°C.

   a. What was \( \Delta T \) for the water? \( \Delta T = T_{\text{final}} - T_{\text{initial}} \)

   _____ °C

   b. What was \( \Delta T \) for the metal?

   _____ °C

   c. Using the specific heat of water (4.184 J/g°C), calculate how much heat flowed into the water?

   _____ joules

   d. Calculate the specific heat of the metal, using equation (7).

   _____________ joules/g°C