

CALORIMETRY

Saddleback College Physics Department

Purpose

- (A) To predict the final temperature of a mixture of hot and cold water and compare it to the actual value.
- (B) To calculate the specific heat of a metal sample and compare it to the handbook value.
- (C) To calculate the latent heat of fusion of ice and compare it to the handbook value.

Theory

Speedometers measure speed, voltmeters measure voltage and accelerometers measure acceleration. In this experiment you will use a device called a calorimeter. The name suggests that it measures a quantity connected with heat since the calorie is a unit of thermal energy. Unfortunately, the amount of thermal energy present in an object cannot be measured as directly as some of the other quantities mentioned above.

The textbook discusses the transformation of energy from one form to another; e.g. from mechanical energy to thermal energy, from chemical energy to thermal energy to mechanical energy, and so on. In this experiment you will investigate some of the properties of thermal energy in transit or heat.

You will measure the heat absorbed or given off by a substance by observing the change in temperature of a standard substance. The calorimeter you will use is simple and consists of an insulated container in which measured quantities of materials can be mixed together without an appreciable amount of heat being gained from or lost to the environment.

You are to make several assumptions about the nature of heat, then, use those assumptions to predict what will happen when you mix two samples which are initially at different temperatures. If your prediction is correct, then you can feel somewhat confident in your assumptions, until they lead to a prediction that turns out to be wrong.

Assumption One: (Zeroth Law of Thermodynamics): Assume that if two objects at different temperatures are brought into thermal contact, heat will flow from the warmer to the cooler object until they reach thermal equilibrium (i.e. the same temperature).

Assumption Two: Assume that the heat transferred between your two samples inside the calorimeter is CONSERVED. In symbols:

$$(1) \quad Q_{gain} + Q_{lost} = 0 \quad \text{or} \quad \text{heat lost by hot object} = \text{heat gained by cold object}$$

The heat gained by a cold object or lost by a hot object is the product of the mass, the specific heat and the change in temperature for that object. In symbols:

$$(2) \quad Q = cm\Delta T$$

where Q is heat in calories or Joules, m = mass in grams, c is specific heat in $\frac{\text{cal}}{\text{g } C^\circ}$ or $\frac{\text{J}}{\text{kg } C^\circ} = \frac{\text{J}}{\text{kg } K}$.

The specific heat of water is $1 \frac{\text{cal}}{\text{g } C^\circ}$ (or $4186 \frac{\text{J}}{\text{kg } C^\circ}$ or $4186 \frac{\text{J}}{\text{kg } K}$), which means it takes 1 calorie of heat to raise 1 gram of liquid water by 1 Celsius degree.

The heat, Q that must be added or removed from a substance in order for it to change phase is given by:

$$(3) \quad Q = m L$$

where m = mass of the substance, L = latent heat of fusion or vaporization. The latent heat of fusion for water is $80 \frac{\text{cal}}{\text{g}} = 33.5 \times 10^4 \frac{\text{J}}{\text{kg}}$, which means that 80 calories (or 33.5×10^4 Joules) of heat must be removed from 1 gram (or 1 kg) of liquid water in order for it to change phase, i.e. go from a liquid to a solid.

Equipment

Styrofoam cups (1lid and 2 bottoms), thermometer, metal sample, ice, boiling water, cold water, balance scale

Procedure

PART I:

- 1) Use the assumptions above to predict the final temperature of a mixture formed when you mix two quantities of water, which are initially at different temperatures. Record your prediction in the DATA section of your lab.
- 2) Use the following steps to test your assumptions. Measure and record the mass of two empty Styrofoam cups. Put about $\frac{1}{3}$ cup of cold water in one cup and about the same amount of hot water in the other cup, then record the mass and temperature of each sample of water. Now mix the two together in one of the cups, stir gently with a thermometer and record the final temperature of the mixture. Does the equilibrium temperature of the mixture confirm that your prediction is correct?
- 3) Calculate the value of $Q = cm\Delta T$ for the hot water and the cold water. Are your assumptions confirmed, or is the difference between the two products greater than can be accounted for by uncertainties in measurement?
- 4) Try another mixture using unequal quantities of water, for example $\frac{1}{4}$ cup of hot water and $\frac{1}{2}$ cup of cold water. Before you mix the two, try to predict the final temperature. What did you predict the temperature of the mixture to be? What final temperature do you actually observe? Is the uncertainty in your thermometer readings and mass measurements enough to account for the difference between your predicted and observed values?

PART II:

- 1) Measure and **record** the mass of a small metal sample, making **note** of what type of metal you believe it to be. Put just enough cold water in a Styrofoam cup (bottom and lid) calorimeter to cover the sample. Tie a thread to the metal sample and suspend it in a beaker of boiling water.
- 2) **Record** the mass and temperature of the water in the calorimeter. When the sample has been immersed in the boiling water long enough to be heated uniformly (two or three minutes)

measure and **record** the temperature of the boiling water. Then lift the sample out and hold it just above the surface of the boiling water for a few seconds to let the water drip off, then QUICKLY transfer it to the cold water of known mass and temperature.

- 3) Stir the water/sample mixture gently with a thermometer and **record** the temperature when the metal sample and water have come to equilibrium. Watch the water temperature closely and **record** the **maximum** temperature reached. This change will be small!

PART III:

- 1) If we place an ice cube in some water in the calorimeter cup, heat will flow from the water into the ice and melt the ice, more heat will then be used to raise the melted ice (now water) to the final equilibrium temperature.
- 2) Use your calorimeter to find the latent heat of fusion of ice. Start with about $\frac{1}{2}$ cup of water that is a little above room temperature, and **record** its mass and temperature. Place an ice cube on a paper towel for a moment to dry off water on its surface, and then transfer it quickly into the calorimeter.
- 3) Stir gently with a thermometer until the ice is melted and the mixture reaches an equilibrium temperature. **Record** this temperature and the mass of the water plus melted ice. What was the mass of the ice that you added? **Record** this value!

Analysis

- 1) PART I:

Answer the following questions from the procedure. *What did you predict the temperature of the mixture to be? What final temperature do you actually observe? Is the uncertainty in your thermometer readings and mass measurements enough to account for the difference between your predicted and observed values?*

- 2) PART II:

Calculate the specific heat of the metal sample using equations (1) and (2) in the theory. Look up and the theoretical specific heat of your metal in either your textbook or the Handbook of Chemistry and Physics. At this point you are forced to identify the material of your metal sample. **Record** the theoretical value of specific heat and **compare** it to your experimental value. What are the significant sources of error?

- 3) PART III:

Calculate the latent heat of fusion for ice, using your data from part III and equations (1), (2) and (3). Compare your experimentally calculated value to the theoretical value in your textbook or in the Handbook of Chemistry and Physics. What are the significant sources of error?

In your conclusion, be sure to summarize all of your results, showing all of the experimental (calculated) values and all of the theoretical values as well as their respective % differences.