Experiment 7 - Finding the Specific Heat of a Metal

Heat is needed to warm up anything, but the amount of heat is not the same for all things. Therefore it is desirable to know the **specific heat** of a substance. The amount of heat needed to raise the temperature of one gram of a substance by one Celsius degree is call the "specific heat" or "specific heat capacity" of that substance. For pure water, that amount is one **calorie**, which is equivalent to 4.184 joules. Almost all other substances have lower specific heats than water does.

Specific heat is defined in terms of one gram of substance and a one degree temperature rise, but of course it is applied to other-sized samples and other temperature changes. So, in general terms:

Heat that must be put into a sample = (temperature change) × (grams sample) × (specific heat)

or $Q = s \bullet m \bullet \Delta T$, where **Q** is the amount of heat, **s** is the specific heat, **m** is the mass of the sample, and **ΔT** is the temperature change. This equation can be used to calculate the amount of heat that must be involved when the other three values are known or measured. If a substance is cooling off, rather than warming up, the term Q represents the heat that is given off by the sample (instead of the heat that must be put into the sample). This same equation can also be used to calculate a specific heat, when the other three terms are known or measured. That is what you will be doing in the present experiment.

The heat that is taken from a hot sample into cooler surrounding material, such as water, can be measured in an insulated container called a **calorimeter**. A weighed amount of the substance being studied is heated to some known (measured) temperature and is then poured quickly into the calorimeter that already contains a measured amount of water at a known, measured temperature. Heat flows from the hot substance into the cooler water, until the sample temperature and the water temperature become equal. (This final temperature will be somewhere in between the initial temperatures of the two substances.)

When two objects at different temperatures are placed in contact with each other, heat always flows from the hotter to the cooler object. Heat will flow until the two reach **thermal equilibrium**, when they are at the same temperature. It will always be true that the number of calories (or joules) of heat lost by the originally hot object will equal the number of calories (or joules) gained by the originally cold object. In other words, the amount of heat lost is equal to the amount of heat gained ($Q_{lost} = Q_{gained}$). In this experiment, the amount of heat that is lost by a sample of metal as it cools is equal to the amount of heat gained by the water in the calorimeter. This assumes that no heat is lost from the calorimeter to its surroundings (the room), and that the amount of heat that is absorbed by the calorimeter itself is so small we can ignore it. Thus:

Heat lost by the metal = heat gained by the water, or

$$Q_{metal} = (\Delta T_m)(m_m)(s_m) = Q_{water} = (\Delta T_w)(m_w)(s_w)$$

where the subscripts m and w identify the metal and the water. In this equation, you will know both ΔT values because you will measure initial and final temperatures. You will know all of the water values, and all but one of the metal values. You will then be able to solve for the "unknown" value, the specific heat of the metal.

(Note: in the above equation, we are assuming that we will be using the absolute value of ΔT. If you are not using the absolute values of ΔT, then $Q_{metal} = - Q_{water}$.)

Safety Precautions:

- Wear your safety goggles.
- Wash your hands after handling the metals.

Waste Disposal:

- There is no waste for this experiment. Return the metal samples so that they can be re-used.

Procedure

1. Obtain a calorimeter (2 nested Styrofoam cups), a thermometer, and a sample of metal in a large test tube.
2. Weigh the sample of metal in the test tube. Then, carefully transfer the metal sample to a clean dry beaker while you weigh the empty test tube, so you can determine the actual weight of the sample alone. NOTE: If it is awkward to weigh the tube and sample on the balance pan, you can support it in a small beaker. Since the beaker will be part of the weight when the empty tube is weighed also, its own weight will cancel out.
3. Replace the metal in the test tube and put the test tube in a beaker of water. The beaker should contain enough water so that the top of the metal is below the surface of the water. Heat the water to boiling and allow it to boil for a few minutes to allow the metal to attain the temperature of the boiling water (100°C).

4. While the water is boiling, weigh the calorimeter. Add about 60 mL of water to the calorimeter and weigh it again. The weight of the water in the calorimeter can be calculated by subtraction.
5. After allowing the water some time to equilibrate its temperature with that of the calorimeter, measure the temperature of the water as accurately and possible. The temperature should be measured to the nearest 0.5°C for greatest accuracy.
6. Take the test tube out of the beaker of boiling water and quickly pour the metal into the water in the calorimeter**. Be sure that no water adhering to the outside of the test tube runs into the calorimeter when you are pouring the metal.**
7. Swirl the water and metal to reach equilibrium, and record as accurately as possible the <u>maximum temperature</u> reached by the water. This is considered the <u>final temperature</u> of both the metal and the water.
8. At the end of the experiment, dry the metal by heating it again in the test tube in boiling water, and then pouring the hot metal onto a paper towel. Return the metal sample in its tube to the place from which you borrowed it.
9. Complete the report sheet, including answers to the two questions below.

Calculations

1. Determine the mass of water in the calorimeter, the mass of metal used, the temperature change of the water, and the temperature change of the metal. Note that the water and the metal each have the same final temperature, but since they started at different temperatures, they have different temperature <u>changes</u>.
2. Using the specific heat of water as 1.00 cal/g•°C or 4.184 J/g•°C, calculate the heat absorbed by the <u>water</u> in calories and in joules. (Use $Q = sm\Delta T$.)
3. Determine the amount of heat given off by the metal in calories and in joules.
4. Use $Q = sm\Delta T$ to determine the heat capacity of the metal. (Make sure to use the heat given off by the <u>metal</u>, the mass of the <u>metal</u>, and the temperature change of the <u>metal</u> in this calculation.)

Questions

1. A large error can be caused in this experiment by allowing some hot water to enter the calorimeter along with the hot metal. If this happens, would it make the specific heat value obtained for the metal higher than it really is or lower? Explain.

2. Another error is caused by the fact that the calorimeter itself absorbs some of the heat from the metal. Does this error make the specific heat value obtained for the metal higher than it really is or lower? Explain.