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Laboratory Safety

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Experiment 1

Heat Capacities of Materials

Arthur B. Ellis, Ann Cappellari,
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Notes for Instructors

Purpose

To determine the specific heats of some materials and to relate them to atomic composition.

Method

Solid samples of known mass are heated in boiling water at a known temperature and transferred to an equal mass of water at a known temperature. The relative temperature changes give the relative specific heat, which is converted to gram specific heat by using the known heat capacity of water; and to the molar specific heat by using the formula weight. Students discover that heat capacity is more strongly related to the number of atoms of matter present than to the mass of matter. For simple solids, the molar specific heat value will be $3R \times p$, where R is the gas constant, 8.3 J/°mol, and p is the number of atoms in the formula, as discussed in Chapter 2.

In this lab, students must decide what experiments should be performed, design the experiments, and pool the results at the end, in addition to gathering data.

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Materials

Solid samples such as aluminum, copper, lead, zinc, brass (a Cu–Zn alloy), pyrite (FeS_2), galena (PbS), quartz (SiO_2), fluorite (CaF_2), dolomite or chalk ($CaCO_3$), and polymers. Samples should be insoluble in water, nonporous, and denser than water. A maximum dimension of 1 inch will allow the samples to fit into a typical Styrofoam cup and still be covered by the mass of water that corresponds to the mass of the sample. The mass of the samples should ideally lie between ~40 and 100 g. One or two of the metals (preferably a relatively dense element such as copper) should either be cut into pieces of about 20 g each so that stacks of 40, 60, 80 and 100 g can be made; or else cut into pieces of about 20, 40, 60, 80, and 100 g.

String

Thermometers

Two Styrofoam cups in a 250-mL beaker *or* 250-, 400-, and 600-mL beakers that are nested.

Large beaker for boiling water

Heat Capacities of Materials

Purpose

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Introduction

Heat is a form of energy, often called thermal energy. Energy can be transformed from one form to another (electrical energy is converted to heat and light in an electric light bulb, for example) but it cannot be created or destroyed; rather, energy is conserved. The higher the temperature of a material, the more thermal energy it possesses. In addition, at a given temperature, the more of a given substance there is, the more total thermal energy the material possesses.

On an atomic level, absorbed heat causes the atoms of a solid to vibrate, much as if they were bonded to one another through springs. A rough sketch of this motion is shown in Figure 1 for a simple cubic structure.

The atoms of the solid can vibrate collectively to progressively greater extents as temperature is increased. This means that the ability of a metal to exchange heat will change as the temperature is changed. Above a certain temperature, however, all vibrations are fully activated and the maximum capacity of the solid to absorb and release heat per degree change in temperature is achieved. In the plot of heat capacity versus temperature shown in Figure 2, the flat region on the right side of the graph corresponds to this situation. For most elements that are metals, this maximum heat capacity is attained by the time room temperature is reached. (Room temperature is approximately 25 ^{o}C or 298 K).

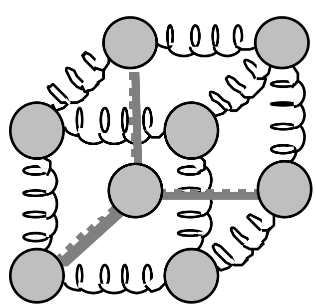

Figure 1. Atoms vibrate as though they were bonded by springs. The shaded circles represent atoms at the corners of a cubic unit cell.

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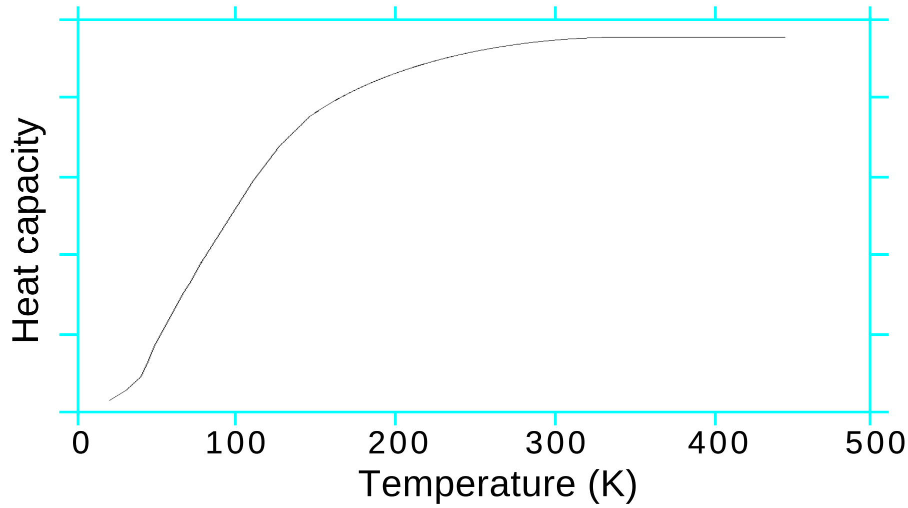

Figure 2. Plot of heat capacity versus temperature for a typical metal.

The heat capacity of a material is often quantified by its "specific heat," which is the amount of heat needed to change the heat content of exactly 1 g of the material by exactly 1 °C. In comparing two materials of the same mass, the material that requires the transfer of more heat to change its temperature by a given amount has the higher specific heat, because more heat will be transferred per gram.

Relative specific heat values can be determined in the following way: When two materials, each initially at a different temperature, are placed in contact with one another, heat always flows from the warmer material to the colder material until the two materials are at the same temperature. At this point, the two materials are in thermal equilibrium, meaning they are at the same temperature and there is no additional net flow of heat.

From the law conservation of energy, the heat lost by the initially warmer object must equal the heat gained by the initially cooler object:

$$(\text{heat lost})_{\text{hot object}} = (\text{heat gained})_{\text{cold object}} \tag{1}$$

The heat quantities in this equation would be calculated as (specific heat) x (mass of the object) x (the temperature change of the object), because specific heat has units of energy per (mass x degrees). If ΔT is the change in temperature, equation 1 becomes

$$(\text{sp. heat} \times \text{mass} \times \Delta T)_{\text{hot object}} = (\text{specific heat} \times \text{mass} \times \Delta T)_{\text{cold object}} \tag{2}$$

If the two materials that are exchanging heat have the same mass, a simpler expression results:

$$(\text{specific heat} \times \Delta T)_{\text{hot object}} = (\text{specific heat} \times \Delta T)_{\text{cold object}} \tag{3}$$

That is, the product (specific heat x ΔT) is the same for the two materials: The larger the specific heat of a material, the smaller its corresponding temperature change and vice versa; the two quantities are inversely related. Another way to write equation 3 is as a ratio:

$$\frac{(\text{specific heat}_{\text{hot object}})}{(\text{specific heat}_{\text{cold object}})} = \frac{(\Delta T_{\text{cold object}})}{(\Delta T_{\text{hot object}})} \tag{4}$$

In the experiment you will perform, you will heat various materials *of known mass* in a hot-water bath *of known temperature* by suspending them on a string; the materials will come into thermal equilibrium with the hot water. You will then transfer each material in turn to a beaker containing water *at a known, lower temperature having the same mass as your sample.* The highest temperature reached after you have transferred the material into the water can be taken as the temperature at which thermal equilibrium has been achieved. You are going to investigate how the temperature change depends on the choice of material being used to transfer heat; on the amount of that material; and on the other experimental conditions you employ. Your group will pool its data to draw conclusions about heat capacity and the atomic composition of the material.

Warm-up Exercises

1. Consider metal chunks (A, B, and C) of three different elements at 100 °C, all of the same mass. Each is then dunked into its own container of water at 20 °C, with the mass of the water being the same as that of the metal chunk. If A, B, and C produce final temperatures of 60, 40, and 30 °C , calculate their specific heats relative to the specific heat of water. That is, if the specific heat of water is considered to be "1", what are the specific heats of A, B, and C?

2. Now suppose that the same hot metals are dunked into a liquid other than water that has a smaller specific heat than water. How will the final temperatures at thermal equilibrium compare to those found in water? Again, the mass of the liquid is the same as that of each of the metals.

3. You can regard equation 3 as being analogous to balancing two people on the two sides of a see-saw. Sketch this situation for metals A and C and indicate what quantities would correspond to the masses of the people, to their distance from the center (the fulcrum) of the see-saw, and what the balancing represents.

General Procedure

Wear eye protection.

Obtain a thermometer.

Partially fill a large beaker with water and heat it to boiling. Note its temperature. Obtain several samples of materials, weigh them, and tie strings securely around each of them. Suspend the metals in the boiling water. (The samples should be suspended, because the bottom of the container being heated may be warmer than the boiling water.)

Either place two Styrofoam cups inside a 250-mL beaker or place a 250-mL beaker inside a 400-mL beaker and place them both inside a 600-mL beaker. Add a volume of room-temperature water of the same mass as your sample to the inner container and note its temperature; the density of water is 1.00 g/mL, so the number of grams of water needed to match the mass of the sample you will immerse in the water is simply the same number in milliliters.

Transfer a sample from the hot-water bath to the room-temperature water in the Styrofoam cup (or 250-mL beaker), stir, and note the maximum temperature reached, which is used to calculate ΔT for the water and for the sample.

Follow the same procedure using other samples of material.

Questions

Answer the following questions collectively, devising and conducting experiments to obtain data where appropriate. Everyone should obtain a copy of all the data collected by the group and use it as the basis for individual written answers. *You should decide as a group what experiments need to be done and who should do what in order to answer the questions.*

1. Divide your efforts so that all of the available materials can be evaluated; ideally, several teams should evaluate each material so results can be averaged. Determine the specific heats of all materials relative to water and describe any trends.
2. If you use different quantities of a particular material, how does the change in temperature caused by transferring the material from the hot to the room temperature water vary (a) if the amount of water stays constant and (b) if the weight of water continues to match the weight of the sample? Support your answers with data.
3. What is likely to happen to the temperature change you will measure if you do not transfer the sample fast enough from the hot water to the cool water? If you don't give the sample time to come to thermal equilibrium with the hot water? How can you test this? Support your answer with data and indicate what effect these errors have on the specific heat you calculate.

4. Speculate on why the cool water is placed in nested containers rather than just one container. Besides being transferred to the water, where else can the heat from your sample go? Does this make your relative specific heat values appear to be larger or smaller? How might you correct for this?

5. Your experiment has provided relative measurements of specific heats of different materials. If you know the actual specific heat of water, you can compare your data with values that have been reported in the scientific literature. Water has a specific heat of 4.18 J/°g, meaning that it takes 4.18 joules of energy to raise the temperature of 1 g of water by 1 °C. Multiply your relative values by the value for water to put units on your specific heat values. Convert your gram specific heats (joules per degree-gram) to molar specific heats (joules per degree-mole), if it makes sense to do so.

6. Does the heat capacity correlate better with the mass of the sample, the number of atoms in the sample, or the density of the sample? (You may use literature values for the density or determine them yourself.) Tabulate your results. What trends do you see? What causes those trends, if any?