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Heat and the First Law of Thermodynamics

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Overview

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When a pot of water is placed on a hot stove, heat is said to "flow" from the stove to the water. When two or more objects are placed into thermal contact with each other, heat spontaneously flows from the hotter objects to the colder ones, or in the direction that tends to equalize the temperature between the objects. For example, when ice cubes are put in a cup of room-temperature water, heat from the water flows to the ice cubes and they begin to melt. Often, the term "heat" is used inconsistently, usually to simply refer to the temperature of something. In the context of thermodynamics, heat, like work, is defined as a transfer of energy. Heat is energy transferred from one object to another because of a difference in temperature.

Furthermore, the total energy of any isolated thermodynamic system is constant—that is, energy can be transferred to and from different objects within the system and can be transformed to different types of energy, but energy cannot be created or destroyed. This is the first law of thermodynamics. It is very similar to the conservation of energy law discussed in another video, but in the context of heat and thermodynamic processes. In the case of ice cubes in water, if the first law of thermodynamics was invalid, then one might expect that adding ice cubes to an isolated room-temperature cup of water would cause the water to boil, which would imply the creation of energy.

Principles

There is a clear distinction between internal energy, temperature, and heat. The internal energy of a substance refers to the total energy of all molecules in the substance. Its temperature is a measure of the *average* kinetic energy of all individual molecules. Consider two pieces of hot metal in thermal equilibrium resting next to each other, one half the size of the other. They both have the same temperature, but the smaller piece of metal has half the thermal energy than the other. Finally, heat, as discussed above, is the *transfer* of energy from different objects.

If heat flows into an object, the temperature of the object rises. However, the amount of the rise in temperature depends upon the kind of material that the heat flows into. The amount of heat, Q , required to change the temperature of any given material is proportional to the mass m of the material present and to the temperature change ΔT . This simple relationship is expressed as:

$$Q = mc \Delta T, \text{ (Equation 1)}$$

where c is a characteristic quality of the material called its *specific heat* (or sometimes called *specific heat capacity*). Rearranging **Equation 1** gives:

$$c = Q / (m \Delta T). \text{ (Equation 2)}$$

Hence, the units of specific heat is J. The specific heat can be described as the amount of heat required to raise 1 g of a substance by 1 °C. At standard atmospheric pressure, the specific heat of water is known to be 4.18 J/(g°C). In other words, if 4.18 J of energy is supplied to 1 g of water, its temperature would increase by 1 °C. However, this is assuming that the sample of water is sufficiently isolated from its surroundings. If it is not, some of the energy being transferred to the water could be lost to the environment surrounding the water—the surrounding air, for instance. This kind of energy loss, or transfer, is referred to as the system "doing work." The first law of thermodynamics can then be written as:

$$\Delta U = Q - W, \text{ (Equation 3)}$$

where U is the total internal energy of a system, Q is the heat added to the system, and W is the work done *by* the system.

This lab will feature a "coffee cup calorimeter," which is essentially a Styrofoam cup. Styrofoam sufficiently insulates the interior substance from the surroundings of the cup so that the system will do no work and $W = 0$.

Procedure

1. Measure the specific heat capacity of lead and demonstrate the first law of thermodynamics.

1. Obtain a scale, a lead sample, two Styrofoam cups, a 300-mL (or larger) beaker, a heating element, a thermometer, a piece of string, water at room temperature, a rod attached to stand with clamps, a graduated cylinder, and scissors.
2. Cut a small portion off the top of one of the Styrofoam cups so that it can act as a lid for the other cup. Make a small hole in the bottom, large enough for the thermometer to fit through, but not larger than the girth of the thermometer.
3. Measure 220 mL of water using the graduated cylinder and pour it into the unmodified Styrofoam cup. Alternatively, 220 g of water can be weighed.
4. Place the modified Styrofoam cup on top of the cup of water so that it acts as a lid; make sure that it fits snugly. If not, make the appropriate modifications.

5. Measure the temperature of the water and record it in **Table 1**. The water should be at room temperature.
6. Fill the beaker with enough water so that the lead sample can be completely submerged. Place the sample in the beaker with the water and verify that there is enough water. Heat the water to boiling using the heating element.
7. Attach the string to the lead sample so that it can be suspended into the boiling water. Place the sample in the water, with the string accessible for moving the sample later.
8. Wait at least 5 min for the sample to come to thermal equilibrium with the boiling water. When the lead sample is removed from the boiling water, it will decrease in temperature very rapidly. Measure the temperature of the sample outside of the boiling water. Proceed to place the sample into the coffee cup calorimeter immediately after taking its temperature. It may be well below 100 °C. Record this temperature in **Table 1**.
9. Swirl the coffee cup/lead system around to ensure a uniform mixture. Watch the temperature on the thermometer as it changes. Once it stops changing, record that temperature in **Table 1**.
10. Using the temperature changes of both the water and the lead sample, and given the specific heat of water, calculate the specific heat of lead using **Equation 1**.

Results

Using the values recorded in **Table 1**, the specific heat of lead can be calculated. From the first law of thermodynamics, it is known that energy is neither created nor destroyed in an isolated system, but energy can transfer between different objects within the system. When the hot piece of lead is put in the coffee cup calorimeter, heat will be supplied from the lead to the water, and that heat transfer is *conserved*; that is, the heat output of the lead, Q_{out} , equals the heat input of the water, Q_{in} .

$$Q_{out} = Q_{in} \text{ (Equation 4)}$$

As in **Equation 3**, the total energy U is constant. Using **Equation 1**, **Equation 4** can be equivalently written as:

$$m_{lead} c_{lead} \Delta T_{lead} = m_{water} c_{water} \Delta T_{water} \text{ (Equation 5)}$$

With the specific heat of water known to be 4.18 J/(g°C) and the information from **Table 1**, c_{lead} can be solved for:

$$c_{lead} = (m_{water} c_{water} \Delta T_{water}) / (m_{lead} \Delta T_{lead}) \text{ (Equation 6)}$$

$$= (220 \text{ g} \cdot 4.18 \text{ J/(g}^\circ\text{C)} \cdot 1.2^\circ\text{C}) / (43.4^\circ\text{C} \cdot 201 \text{ g})$$

$$= 0.127 \text{ J/(g}^\circ\text{C)}.$$

The accepted value for the specific heat of lead is 0.128, so the results here are in excellent agreement, with only a 1.5% difference.

Table 1. Experimental results.

	T_i (°C)	T_f (°C)	m (g)
Water	18.5	19.7	220
Lead	63.1	19.7	201

Applications and Summary

The first law of thermodynamics applies to the entire universe-no energy can be created or destroyed throughout the universe, but all kinds of energy transfers and transformations do take place. Plants convert energy from sunlight into the chemical energy stored in organic molecules, many of which we subsequently eat. Nuclear power plants that produce much of our electricity use heat transfer from hot radioactive rods to produce steam, which powers turbines that generate electricity. Refrigerators work by using electricity to pull heat out of the system. An evaporator filled with coolant and a condenser perform work on the refrigerator to effect a negative heat transfer.

Heat transfer was observed in a closed system between a piece of hot lead and room-temperature water. The specific heat capacity was measured by measuring temperature changes in known quantities of water and lead. If the Styrofoam cup system was not sufficiently insulated from its surroundings, heat from the system would have been lost-in other words, the hot water/lead would have done work on the surroundings, as in **Equation 3**. If this was the case, the calculations performed in this lab would have been much more difficult to make, since the surrounding air readily dissipates heat to its surroundings. Because Styrofoam cups acts as a good insulator, the system was considered independent of the surrounding air. The first law of thermodynamics was observed, as no energy was created or destroyed during the experiment; the energy of the closed system was conserved.