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## Heat and the First Law of Thermodynamics

--Manuscript Draft--

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**Overview:**

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 you put ice cubes in a cup of room-temperature water, heat from the water will flow to the ice cubes and they will begin to melt. Oftentimes the term "heat" is used inconsistently – usually to refer simply 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 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 wasn't valid then one might expect adding ice cubes to an isolated room-temperature cup of water would cause the water to boil, which would imply a 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 the molecules in the substance. Its temperature is a measure of the *average* kinetic energy of all the 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 object's temperature rises. However the amount of rise in temperature depends on what kind of material the heat is flowing into. The amount of heat, $Q$, required to change the temperature in any given material is proportional to the mass $m$ of the material present, and to the temperature change $\Delta T$. This simple relationship is expressed as

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

where $c$ is a characteristic quantity of the material called the its *specific heat (*or sometimes called *specific heat capacity)*. Rearranging **Equation 1**, we see that

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

Hence the units of specific heat is J. The specific heat can be described as the amount of heat required to raise one gram of a substance by one degree Celsius. At standard atmospheric pressure, the specific heat of water is known to be 4.18 J/(g C°). In other words, if you had one gram of water and you supplied 4.18 Joules of energy to the water, its temperature would increase by one degree Celsius. However this is assuming the sample of water is sufficiently isolated from its surroundings. If it wasn't, 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

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

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 use a "coffee cup calorimeter", which is essentially just a styrofoam cup. Styrofoam sufficiently insulates the interior substance from the cup's surroundings so 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.

**1.1.** Obtain a scale, a sample of lead, 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, and a graduated cylinder, and scissors.

**1.2.** Cut a small portion off the top of one of the styrofoam cups, so it can act as a lid for the other cup. Also make a small hole in the bottom, large enough for the thermometer to fit through but not larger than the girth of the thermometer.

**1.3.** Measure 100 ml of water with the graduated cylinder and pour it into the unmodified styrofoam cup.

**1.4.** Place the modified styrofoam cup on top of the cup of water so it acts as a lid, and make sure it fits snugly. If not, make the appropriate modifications.

**1.5.** Measure the temperature of the water and record it in **Table 1.** The water should be at room temperature.

**1.6.** Fill the beaker with enough water so that the~~at your~~ lead sample can be completely submerged without touching the bottom. Place the sample in the beaker with the water and verify that there is enough water~~.m~~..

**1.7.** Remove the sample from the beaker and place the beaker ~~Place the beaker of water~~ on the heating element and turn it on. Heat the water until it reaches a boil.

**1.8.** Attach the string to the lead sample so it can be suspended into the boiling water.

**1.9.** Attach the thermometer and string to the clamp/rod so the bottom of the thermometer and the lead sample are suspended at the same height.

**1.10.** Submerge the thermometer/lead sample into the boiling water until the sample is completely covered with water, but not touching the bottom. There may be a temperature gradient in the water beaker since the heating element is transferring~~inputing~~ a large amount of energy from the bottom of the beaker.

**1.11.** Wait at least 5 minutes for the sample to come to thermal equilibrium with the boiling water. Once they are in thermal equilibrium the thermometer reading will be constant at 100 degrees C. At this point, ~~you think they are in thermal equilibri~~um, record the temperature on the thermometer in **Table 1.**

**1.12.** Remove the thermometer from the boiling water and set aside. Then remove the lead sample from the water and quickly (carefully) place the sample in the coffee cup calorimeter. Place the lid on the calorimeter and insert the thermometer through the top.

**1.13.** 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**.

**1.14.** Using the change of temperatures of both the water and lead sample, and knowing the specific heat of water, calculate the specific head of lead using **Equation 1.**

**Representative 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 in an isolated system, the amount of energy is ~~neither created or~~neither created nor destroyed, but energy can transfer between different objects within the system. When the hot piece of lead is put in the coffee cup calorimeter, heat

**Commented [AK1]:** Is the lead sample dropped into the beaker prior to boiling to judge the correct volume?

**Commented [AK2]:** Practically, what does this mean here? That the temperature reading on the thermometer has stabilized?

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, $Q_{in}$

$$Q_{out} = Q_{in}. \quad \textbf{(Equation 4)}$$

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

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

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

$$c_{lead} = (m_{water}\, c_{water}\ \Delta T_{\text{water}}) / (m_{lead}\ \Delta t_{lead}) \quad \textbf{(Equation 6)}$$
$$= (100\text{ g} \cdot 4.18\text{ J/(g C}^{o}) \cdot 3\text{ C}^{o}) / (75\text{ C}^{o} \cdot 44.1\text{ g})$$
$$= 0.127\text{ J/(g C}^{o})$$

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**

| | $T_i$ (C°) | $T_f$ (C°) | $m$ (g) |
|---|---|---|---|
| Water | 22 | 25 | 100 |
| Lead | 100 | 25 | 131.7 |

**Summary**

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 observing 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 from that system – 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. Since the styrofoam cups acts as a good insulator, the system was considered to be 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.

**Commented [AK3]:** The idea that the work done by the system on its surroundings is subtracted should be revisited here. Again, why was this a factor or nonfactor in this experiment.

**Applications:**

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