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So You Thought a Glass Thermometer Measured Temperature

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Of cource it measures temperature; it just isn't the temperature you may have thought. The reason is that a thermometer is an active measuring device, and its presence influences the value of the measurement. When a room-temperature glass thermometer is placed in contact with a warmer sample for the purpose of determining the sample's initial temperature, there is heat lost by the sample to warm the thermometer; hence, the thermometer indicates the thermometer-sample equilibrium temperature. A cursory inspection of elementary physics texts[1–5] indicates this important feature of the measurement process is not discussed. The questions of interest here are: (1) what physical quantities influence the difference between the sample's initial temperature and the thermometer's equilibrium temperature, and (2) how large a sample is needed to minimize that difference (so that the simple assumption that the thermometer temperature is the initial sample temperature is valid).

Analysis

An ordinary inexpensive glass-alcohol thermometer commonly used in classrooms is 30 cm in length and consists of 25 g of glass and 2 g of alcohol. The products of heat capacity and mass show the thermometer requires 6 cal to raise its temperature one degree. If the sample temperature is not too high above room temperature and the sample is large, the sample heat loss to the thermometer is small. However, if the sample's temperature is near that of boiling water and the sample size is not too large, then the heat loss is significant. When the sample temperature is 80°C, then the heat needed to raise a room-temperature thermometer to 80°C is 342 cal. When the sample is 342 g or less of water, then the sample's initial temperature will drop by one degree or more.

To deal with the questions on a quantitative basis, let T_s, M_s, and C_s be the sample's initial temperature, mass, and heat capacity, respectively. We take the thermometer's initial temperature to be T_0, its equilibrium temperature to be T, and the thermometer's mass and heat capacity as M_t and C_t. Since the thermal conductivity of glass is small, and the time required for equilibrium short (a minute or two), the thermometer's heat-absorbing mass in contact with the sample is given by $M_t\,(L/L_0)$. L_0 is the thermometer's length, and L is the length in contact with the sample. Assuming a closed-system model of only the sample and thermometer, the sample's heat loss equals the thermometer's heat gain, and we obtain

$$M_sC_s(T_s - T) = M_t(L/L_0)C_t(T - T_0). \tag{1}$$

Solving for T_s yields,

$$T_s = T + \gamma\Delta T, \tag{2}$$

where $\Delta T = T - T_0$, and

$$\gamma = M_tLC_t/M_sC_sL_0. \tag{3}$$

Gamma is a ratio of two energies, the energy absorbed by the thermometer per degree to the energy liberated by the sample per degree.

Let's get some feel for the numbers by taking the case of 100 g of water with an initial temperature $T_s = 80$°C. Solving Eq. (2) for T, and assuming full immersion of the thermometer with $L/L_0 = 1$, the thermometer's equilibrium temperature is $T = 76.6°C$. The temperature loss of the water to warm the thermometer is 3.4°C. That seems quite large, but is there any support for it from experimental evidence? Yes.

Experimental Results

Initially, an ordinary classroom calorimeter made from Styrofoam, an inexpensive glass thermometer, and a semiconductor digital thermometer (Extech Corp. Model #421305) good to the nearest 0.1°C were used to measure the initial and equilibrium temperatures. It quickly became apparent the heat loss by the Styrofoam calorimeter presented a problem; when about half full of 80°C water, the temperature loss was 2.7°C/min. To reduce the loss, a tall stainless-steel travel coffee cup with a vacuum jacket was used as a calorimeter. When filled about half full of hot water, the stainless-steel calorimeter had a temperature loss somewhere between 0.2°C and 0.3°C/min.

One hundred grams of hot water were placed in the calorimeter, and the temperature was monitored with the digital thermometer by placing a small part of its tip in the water. After three to four minutes, the temperature stabilized, and it was taken to be the initial temperature T_s. About one-third of a room-temperature glass thermometer was immersed in the hot water, and after about a minute or so, the thermometer's temperature T stabilized to the equilibrium value. The result for five experimental trials indicated a temperature difference close to 1°C, which is within about 15% of the value expected from Eq. (2).

In the previous discussion, the digital-thermometer temperature is taken to be the sample's initial temperature. In fact, the digital-thermometer temperature is the equilibrium temperature of the sample digital-thermometer system. Since the corresponding value of $M_t C_t (L/L_0)$ for the probe is less than one calorie per degree when about one-tenth of the probe is immersed, the corresponding γ value is less than 0.01 for the sample of 100 g of water, and hence, the difference between the temperatures is negligible.

Comments

Clearly the fact that the glass thermometer is an active measuring device is demonstrated where the temperature difference is dependent upon the ratio of the heat absorbed by the thermometer per degree to the heat liberated by the sample per degree. Hence, it follows that textbook calorimetry problems and laboratory calorimetry exercises can be made more realistic by including the heat absorption or loss by the thermometer.

Our second question is how large a sample is necessary to insure the equilibrium temperature T is a good approximation to the initial temperature T_s. For purposes of discussion, let's say T is to be within 0.5°C of T_s, where T_s is 80°C, the thermometer is at room temperature, and the whole thermometer is in contact with the sample, so $L/L_0 = 1$. We solve Eq. (2) and obtain $M_s C_s = 684$ cal/°C. If the sample is water, we need 648 g, and when it is copper metal ($C_s =$ 0.092 cal/g°C), we need in excess of 5 kg! In the case of ice at initially −10°C, we need 792 g to obtain an equilibrium temperature of −9.5°C. The fact that the heat capacity of ice is one-half that of water is responsible for the large amount of ice.

Realistically, we rarely have the whole thermometer in contact with the sample. So for the samples mentioned above using Eq. (2), when only one-third of the thermometer is in contact with the samples, then only one-third the sample mass is needed: 228 g of water, 1.6 kg of copper, and 264 g of ice. These cases illustrate the importance of the thermometer contact with the sample, and they point to an important experimental technique rarely observed when students perform calorimetry experiments: When using a thermometer to determine a sample's temperature, the smaller the portion of the thermometer in contact with the sample, the smaller the γ value, and the closer the equilibrium temperature is to the initial sample temperature.

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