Applications and Analogies

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Heat Capacity, Body Temperature, and Hypothermia

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Heat capacity, in conjunction with thermochemistry, is discussed in chemistry courses at the high school level, in college general chemistry, and again in physical chemistry. Changing the temperature of a substance depends on three factors—mass, specific heat capacity, and energy—and thus these temperature changes perplex many introductory students. Articles in this *Journal* have offered clever analogies and approaches in the teaching of heat capacity (*1*, *2*). In addition, the dramatic effect of high heat capacity of liquid water is exemplified in many general chemistry textbooks as the moderating effect of lakes and oceans on weather (*3–6*). However, weather effects, while important, are macroscopic in nature and are often difficult for introductory students to incorporate into their experiences, particularly if they do not reside near any large bodies of water. Offered here is an application of the significance of water's high heat capacity, which has interesting life science implications.

In chemistry, we normally define "room temperature" as 25 °C, which is 77 °F—in actuality, a rather warm room. However, if we immerse our finger in a cup of room-temperature water, the water will feel cool. This is, of course, due to the fact that as warm-blooded animals, we radiate heat. We are easily able to heat the air around our bodies because air has a fairly low density (0.0012 g/mL at 25 °C for dry air at 1.00 atmosphere pressure) (*7*), and a relatively small specific heat capacity (1.01 J/g °C) (*6*). Air that is 25 °C feels warm because little of our body heat is lost in warming it, but four times more heat is required to increase the temperature of water, as its specific heat capacity is 4.18 J/g °C. Furthermore, the density of water is far greater, 0.997 g/mL at 25 °C (*7*). As a result, our finger is in contact with a much greater mass of water to which it is losing significantly more heat than it was losing to the equivalent volume of air it contacted before immersion.

Complete immersion of my index finger requires approximately 12 mL of water.[1] The heat lost by the finger in raising the temperature of the water just 1 °C is

$$12\ \mathrm{mL} \times 0.997\frac{\mathrm{g}}{\mathrm{mL}} \times 4.18\frac{\mathbf{J}}{\mathrm{g}\ {}^\circ\mathrm{C}} \times 1.0\ {}^\circ\mathrm{C} = 50\ \mathbf{J}$$

as compared with the same volume of air:

$$12\ \mathrm{mL} \times 0.0012\frac{\mathrm{g}}{\mathrm{mL}} \times 1.01\frac{\mathbf{J}}{\mathrm{g}\ {}^\circ\mathrm{C}} \times 1.0\ {}^\circ\mathrm{C} = 0.015\ \mathbf{J}$$

This oversimplified treatment does not take into account the transfer of heat from the air or water to the surroundings. Rather, it assumes that the system (finger plus air or water) is completely insulated. If one takes the surroundings into account, the calculated heat lost by the finger would be greater in both cases, although more for the dry finger than for the one immersed in water. Nonetheless, this activity serves to illustrate the dramatic difference between the energy required to change the temperature of water and air.

This difference has serious implications for the maintenance of body temperature and the prevention of hypothermia, or unintentional lowering of the body temperature below the normal range. Because cells have a high water content, their heat capacity is correspondingly large. Thus humans (and other warm-blooded animals) are resistant to internal temperature changes when exposed to ambient temperature fluctuations. This is particularly true when coupled with other temperature-protective features such as perspiration, shivering, vasoconstriction and dilation, and subcutaneous fat.

It also follows that, once the body temperature *has* been significantly lowered, raising it back to normal ranges requires a significant metabolic output of energy. Once its temperature has dropped below a certain point, the body cannot recover without an external heat source; it simply cannot produce enough energy by metabolism to increase its temperature, particularly since metabolism is slowed by the lowered temperature (*8*). Consequently, many deaths occur each year as a result of hypothermia. Not all of these fatalities are a result of winter's subfreezing temperatures in cold-weather states, however. Water's high heat capacity can play a role here as well. Exposure to water that is below normal body temperature can rapidly accelerate the hypothermia process (*8*), which is illustrated semiquantitatively in the calculations above by the greater amount of energy needed to heat water than air. Many accidental drownings are linked to unconsciousness resulting from hypothermia, and one does not need to be totally immersed in water to be at risk. For example, four soldiers died of hypothermia in Florida in 1995 even though the air temperature was relatively warm (*9*). Their deaths were attributed to the fact that they were briefly immersed in chest-deep water, and because they were thoroughly wet, they were unable to maintain a safe body temperature.

The energetic ramifications of water's high heat capacity in the maintenance and disruption of body temperature in warm-blooded animals offers an interesting and relevant example that links the concepts of heat capacity and temperature change to energy and metabolism. A discussion of the causes and prevention of hypothermia can provide an interesting context for an introduction to calorimetry.

Note

1. Significantly less water is required to merely wet the author's finger. The 12 mL in a long test tube will provide a layer of water approximately 2 mm thick around the index finger.

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