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ELEMENTS OF PHYSICS

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To relate the BTU and the calorie, we use the experimental fact that for small temperature intervals the heat required to raise the temperature of a given quantity of water is proportional to the temperature interval. Since 1 C deg = $\frac{9}{5}$ F deg, we see that

> $\frac{9}{5}$ BTU will raise the temperature of 1 standard pound, or 453.6 g, of water from 14.5° C to 15.5° C; and thus that
>
> 1 BTU will raise the temperature of $\frac{5}{9} \times 453.6\ \text{g} = 252.0\ \text{g}$ of water from 14.5° C to 15.5° C.

But it requires 252.0 cal to do just this; hence

$$1\ \text{BTU} = 252.0\ \text{cal}. \tag{2}$$

The thermal units were first defined in the days when it was believed that heat was an invisible, weightless substance known as 'caloric fluid,' which was pictured as flowing from bodies at high temperature to bodies at low temperature somewhat as water flows from high elevations to low elevations. Indeed, it can be seen that equation (1) is entirely consistent with the conservation of a substance. Later, the work of Rumford and Joule, to be described in Sec. 3, demonstrated that heat actually represents transfer of *a form of energy* from hot bodies to cold bodies and that quantity of heat can therefore be measured in conventional energy units such as the J and the fp. The thermal units, calorie and BTU, are still used almost exclusively in specifying the quantities that determine the thermal properties of matter in spite of the fact that modern techniques of electrical calorimetry usually determine these quantities initially in joules.

2. Specific Heat, Latent Heat

> The **specific heat** of a substance is the quantity of heat required to increase the temperature of unit mass of the substance one degree.

Specific heat is ordinarily denoted by the symbol c; it is expressed in cal/g·C deg, kcal/kg·C deg, or BTU/lb·F deg. Since

$$\frac{1\ \text{BTU}}{1\ \text{lb}\times 1\ \text{F deg}} = \frac{252.0\ \text{cal}}{453.6\ \text{g}\times \frac{5}{9}\ \text{C deg}} = 1\ \frac{\text{cal}}{\text{g·C deg}} = 1\ \frac{\text{kcal}}{\text{kg·C deg}}, \tag{3}$$

the numerical value of the specific heat is the same in any of these three units. We notice that the cal and BTU were defined in such a way that the specific heat of water would be exactly one unit in either system; since the specific heat of any other substance must bear a constant ratio to that of water, the specific heat of any substance must have the same numerical value in either system, as demonstrated by (3).

While the specific heat of a substance varies slightly with the temperature at which the temperature change takes place, it will be adequate for our present discussion to assume that specific heat is a constant independent of temperature. Then we can determine the heat Q necessary to raise the temperature of a mass m of a substance by ΔT degrees by multiplying the specific heat c by m and by ΔT. Hence,

$$Q = c\,m\,\Delta T. \qquad \left\{\begin{array}{l} Q \text{ in cal} \\ c \text{ in cal/g·C deg} \\ m \text{ in g} \\ \Delta T \text{ in C deg} \end{array}\right\} \text{ or } \left\{\begin{array}{l} Q \text{ in BTU} \\ c \text{ in BTU/lb·F deg} \\ m \text{ in lb} \\ \Delta T \text{ in F deg} \end{array}\right\} \tag{4}$$

Values of the specific heat of various solids and liquids are given in Table I; the specific heats of gases will be considered in Chapter 18. It will be noted that water has the highest specific heat of any of the common substances listed in Table I.

TABLE I

Specific Heats of Solids and Liquids

(In cal/g·C deg or BTU/lb·F deg)

Metallic solids		Nonmetallic solids		Liquids	
Aluminum	0.212	Ice	0.48	Water	1.00
Brass	0.090	Clay	0.22	Ethyl alcohol	0.58
Copper	0.094	Coal	0.3	Gasoline	0.5
Gold	0.031	Concrete	0.16	Mercury	0.033
Iron and steel	0.11	Glass	0.12–0.20	Mineral oil	0.5
Lead	0.031	Limestone	0.22	Methyl alcohol	0.60
Platinum	0.032	Marble	0.21	Olive oil	0.47
Silver	0.056	Paraffin	0.69	Petroleum	0.51
Tin	0.055	Rubber	0.48	Sea water	0.93
Zinc	0.094	Wood	0.3–0.7	Turpentine	0.41

It is also convenient to define a quantity called the *heat capacity:*

> The **heat capacity** of a body is the quantity of heat required to raise the temperature of the body by one degree.

Heat capacity is measured in units such as cal/C deg or BTU/F deg.

Specific heats can be determined calorimetrically by the method of mixtures. The following example will illustrate the principle (1) on which all calorimetric computations are made. In this example, it is assumed that the calorimeter is perfectly insulated.

EXAMPLE. *A copper container (Fig. 1) of* 250-g *mass contains* 400 g *of water. Container and water are initially at room temperature of* 20.0° C *as measured by the thermometer. A block of copper of* 1-kg *mass is heated to* 100° C *by placing it in the steam from water boiling at normal atmospheric pressure. It is then removed from the steam and quickly placed in the water of the calorimeter. The copper block cools, the water and the container become warmer, and the final temperature, as read on the thermometer, is found to be* 34.5° C. *From these data determine the specific heat* c_{Cu} *of copper.*

The fundamental equation (1) becomes

$$\begin{Bmatrix}\text{heat lost by}\\ \text{copper block}\end{Bmatrix} = \begin{Bmatrix}\text{heat gained}\\ \text{by water}\end{Bmatrix} + \begin{Bmatrix}\text{heat gained by cop-}\\ \text{per container}\end{Bmatrix}.$$

The value of each of these terms is given by (4) as $mc\,\Delta T$. Substituting the given values of m and ΔT in the metric units listed in (4), and setting $c=1$ cal/g·C deg for water, we have

$$1000\text{ g}\cdot c_{\text{Cu}}\cdot 65.5\text{ C deg} = 400\cdot 1\cdot 14.5\text{ cal} + 250\text{ g}\cdot c_{\text{Cu}}\cdot 14.5\text{ C deg},$$

whence $$c_{\text{Cu}} = 0.094\text{ cal/g}\cdot\text{C deg}.$$

In the above substitution, 14.5 C deg is the temperature increase of water and container; 65.5 C deg is the temperature decrease of the copper block.

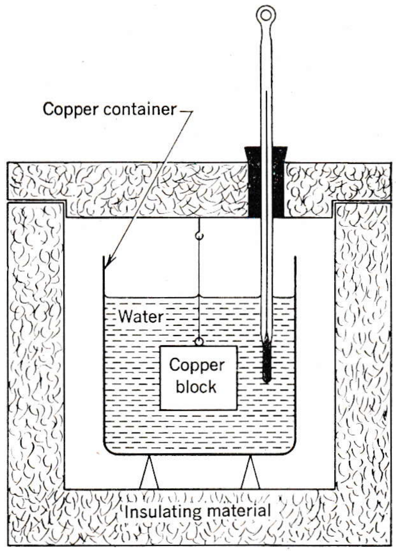

Fig. 1. A simple calorimeter for determination of the specific heat of copper by immersion of a heated block in water.

We now turn to consideration of *latent heats*. If we start with a piece of ice at −20° C, we must add heat to it (0.48 cal/g·C deg according to Table I) in order to raise its temperature to 0° C. At this point further addition of heat causes some of the ice to melt. *During the melting process no temperature rise takes place.* For each 80 cal of heat added, one gram of ice melts. Only after sufficient heat has been added to melt *all* the ice does any temperature rise of the water take place. The value, 80 cal/g, is known as the *latent heat of fusion* of ice.

> The **latent heat of fusion** of a substance is the heat that must be added to unit mass of the solid at its melting point to change it to liquid at the same temperature and pressure.

If we put a pan of water on the stove and turn on the gas, the tempera-