Physics that Textbook Writers Usually Get Wrong

II. Heat and Energy

By Robert P. Bauman

Jacob Bronowski[1] has observed that the most important discovery of science is science itself. By *science*, we generally understand a process of discovery and communication based on unambiguous terminology and symbols. Present textbooks tend to follow historical definitions and terminology, even when these are ambiguous. The ambiguous terminology works against our best efforts, despite efforts of authors to clarify the meanings.

As we look back at the history of physics, we can see that Leibniz proposed *vis viva*, mv^2, as the important quantity for mechanics in the 17th century, about the same time that physics based on momentum was developed. Today energy is often considered an easier subject than momentum. It is a scalar, rather than a vector quantity, and is more a part of our everyday language. This familiarity of language, however, rather than being helpful, has proved to be a major stumbling block to understanding several aspects of energy.[2]

Despite great advances in mechanics and the seemingly clear demonstrations by Rumford at the end of the 18th century, it was not until the middle of the 19th century that energy was fully recognized as a "conserved" quantity. It may appear surprising to us today that it took so long to recognize such an important concept. The source of the delay seems clear, however, if we examine residual problems of nomenclature and definitions that perpetuate misunderstandings.
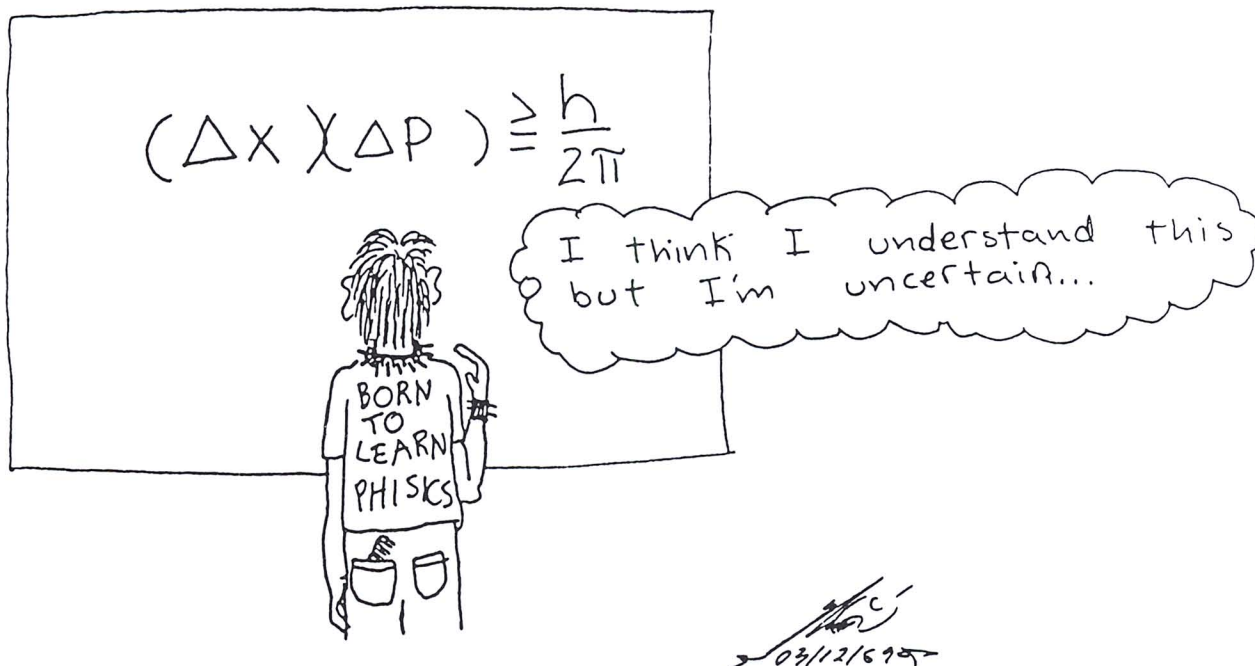
Thermal Energy and Temperature

Start with a simple quiz. What is the commonly accepted term for the following quantities?

    a. The property measured by a thermometer.

    b. Energy in a body that is responsible for determining the reading on a thermometer.

    c. Energy transferred to, or from, a body as a consequence of a difference in the property read by a thermometer.

    d. The process that will increase the reading of a thermometer in contact with a body.

    e. The process of adding energy to a body because of a difference in the property measured by a thermometer between the body and its surroundings, whether or not the process causes any change in thermometer reading of the body.

The commonly accepted language of science differs very little from the vernacular. Item **a** is, of course, called temperature. Nevertheless, it is surprisingly often called *heat*. When we speak of red heat or white heat, or contrast heat and humidity, we are speaking of temperature.

Cartoon by Ehren Stillman, a high-school student from Deroche, BC, Canada.

Item **b** is almost always called *heat* in physics textbooks (including some thermodynamics textbooks). It is combined with modifiers to give heat capacity and latent heat. In contrast, most thermodynamics textbooks profess to limit the term *heat* to mean item **c**, although item **c** is often called heat flow or heat transfer, more consistent with item **b** as the definition of heat.

Item **d** is the verb, *to heat*. We heat an object when we raise its temperature by means of a temperature difference with the surroundings, or by doing work on it, or by triggering an internal conversion of energy. But, as in item **e**, we also heat a block of ice to melt it, without raising its temperature.

Confusion extends to symbolic representation. So long as item **c** could be interpreted as a change in item **b** (a change in *heat* or *caloric*), item **c** could appropriately be written as ΔQ or dQ. As it slowly became recognized that **c** was *not* the same as a change in **b**, it was recognized that **c** was simply a quantity of energy transferred and therefore should be represented by Q. (That did not solve the question of how to represent a small quantity of energy transferred, which will be considered below.)

Note in particular that if we accept the equation

$$\Delta E = Q \qquad (1)$$

where this is the only mode of energy transfer, it follows that

$$\Delta Q = \Delta (\Delta E) \qquad (2)$$

and

$$dQ = d (\Delta E) \qquad (2a)$$

but these are not the intended meanings.

With the confusion in everyday language, it is not surprising that students enter our courses without a clear distinction between items **a**, **b**, and **c**. Given the overlapping of terminology in the classroom, it is not surprising that students leave our courses with a weak distinction between **a**, **b**, and **c**. Student perceptions are often best exhibited by the nature of the mistakes they make. Experience has shown that students beginning a thermodynamics course tend to confuse isothermal ($\Delta T = 0$), isoergic ($\Delta E = 0$), and adiabatic ($Q = 0$).

To distinguish the several quantities collectively and individually known as *heat*, we adopt the following terminology.

a. The property read by a thermometer will be called *temperature*.

b. Energy within a body, responsible for the temperature, will be called *thermal energy*.

c. Transfer of energy, as a consequence of a temperature difference, will be called *thermal energy transfer* and represented by the symbol Q. A small (infinitesimal) quantity of thermal energy transferred will be represented by q. Note that q is not a change in any identifiable quantity, and therefore is *not* a differential (exact or inexact).

Heat will be retained as a generic term, useful when it is not necessary to distinguish between meanings or processes. For example, the verb *heat* will loosely designate either the process of adding energy (as Q, as W, or as an arbitrary mixture) or the process of increasing the temperature (by $Q > 0$ or $W > 0$ or by internal energy conversion).

Precise definition of thermal energy is difficult.[3] Some authors have claimed that it cannot be defined, but these authors then use the term *heat* to represent the quantity, without definition.

"Physics that Textbook Writers Usually Get Wrong"

Conservation Laws

If a ball is dropped from a height of 1.00 m and rebounds to a height of 0.80 m, is energy conserved? In most textbooks the answer depends on the chapter in which the question appears. Students are told that there are *nonconservative* forces, for which energy is not conserved, and they are told that energy is always conserved. What are they to believe?

In addition to the several meanings of conservation and conservative in everyday language and in environmental studies, conservation has two primary meanings in science textbooks.

> 1. Conservation indicates that a property of the system does not change; it is a "constant of the motion." Energy is not a constant of the motion for the bouncing ball, or in problems involving friction.
>
> 2. Conservation indicates no creation or loss, but does not limit changes for the system. Then conservation of energy, for example, would be expressed in the form

$$\Delta E_{system + surroundings} = 0 \tag{3}$$

When we speak of "conservation laws," it is the second meaning that is intended. Yet most occurrences of the term in physics textbooks are in the first sense. The problem would be mitigated somewhat if students were told that the term was being used in conflicting ways, but that is not done.

It may appear that the difference is minor, but the ambiguity appears to block a clear understanding of what the conservation laws mean. For example, most physics students and instructors have difficulty identifying which of the following thermodynamic functions are conserved quantities (in the second sense): T (temperature), P (pressure), V (volume), E (energy), H (enthalpy: $H = E + PV$), and S (entropy).

It should be clear that temperature cannot be limited by a conservation law. A beaker of water may be split into many parts, each at the same T, so we get "more" T. Or we can mix one (large) beaker of water at 90°C with a second (small) beaker at 20°C and get a T that is neither the sum nor the average. A reaction carried out in an insulated bomb may change the temperature of the system with no effect on the surroundings.

More generally, an intensive property cannot be subject to a conservation law. It follows that P, also, cannot be a conserved quantity, as illustrated by the reaction in an insulated bomb, where P inside changes with no effect on the surroundings.

Energy is a conserved quantity, as already mentioned. We will return to it later. Enthalpy, however, is not conserved. For example, even an ideal gas changes PV when T changes, so $E + PV$ changes in the confined, insulated chemical reaction.

Entropy is the extreme counter example to a conserved quantity. For *every* process that actually occurs (in finite time), the entropy of system + surroundings increases. That is the second law of thermodynamics. We can approach arbitrarily close to equilibrium, or reversibility, but overall there are only increases of entropy (of system + surroundings), never decreases.

We skipped over volume. Obviously we can change the volume of a system: inflate a balloon or warm a block of copper, for example. The question, then, is what happens to the volume of the surroundings as the volume of the system increases. In every case, the change in volume of surroundings is equal and opposite to the change in volume of the system. The total volume is constant. Volume is the best known and understood quantity subject to a conservation law.

Conservation laws are sometimes expressed as a special case. For example, conservation of energy is often stated in the form:

The energy of an isolated system is constant.

There are several important difficulties with this form. First, we cannot perform experiments on isolated systems. Second, if we do define our system in such a way that it is isolated (e.g., temporarily), then $Q = 0$ and $W = 0$ and we are cut off from much of the familiar apparatus of thermodynamics.

More important, however, is the fact that if a system is isolated, then its momentum and angular momentum and mass are constant. It would be a weak form of conservation that applies only when *all* conserved quantities are constant for a system!

A less common, but important, variation is the statement that $\oint dE = 0$. The sum of the changes in energy around any

complete cycle or closed path is zero. This statement is valid, and illustrates that energy is a property of the system or a "state function," but has little if anything to do with conservation of energy. It is also true, for example, that $\oint dT = 0$ and $\oint dS = 0$, yet neither T nor S is conserved.

Finally, we must raise the question of whether energy is still a conserved quantity under special relativity, or whether it is only "mass-energy" (often interpreted to mean a sum of terms) that is constant. What special relativity tells us is that "rest mass," m_o, is not conserved.[4] However, there is no exception known to conservation of energy. Therefore if we follow Einstein and define mass from the equation $m = E/c^2$, then mass is also conserved. Of course, any linear combination, such as $m + E/c^2$, is then conserved as well, because m and E are each conserved.

The First Law of Thermodynamics

Most thermodynamicists agree that the first law of thermodynamics is the law of conservation of energy. If the surroundings are defined as all of the universe, except the system influenced by the process under consideration, we can avoid cosmological arguments without limiting ourselves to special cases.

The first law is often said to be expressed by the first-law equation,[5]

$$\Delta E = Q + W \qquad (4)$$

where Q is the thermal energy transfer to the system and W is energy added to the system as work. How the energy is stored in the system, if at all, is not hinted at by the first-law equation.

To connect the first-law equation to the first law, we must accept Eq. (3) as our conservation law, rather than a limitation to isolated systems (for otherwise $Q = W = 0$), but three steps are required in addition.

(a.) Show that

$$Q_{system} = -Q_{surroundings} \qquad (5)$$

and

$$W_{system} = -W_{surroundings} \qquad (6)$$

[or at least that $(Q + W)_{system} = -(Q + W)_{surroundings}$].

(b.) Show that Q and W are always defined, regardless of the nature of the process.

(c.) Show that there is no other way of transferring energy.

Usually (a.) is accomplished only implicitly, by defining Q and W in such a way as to make the equivalence necessary. [Even so, some authors explicitly deny the validity of Eqs. (5) and (6).] By contrast, the last two requirements cannot be met. Neither Q nor W is operationally defined for many processes, such as those involving friction.[6] And the primary energy loss mechanism for a pan of boiling water is by evaporation, which is neither Q nor W. Similarly, most of the

energy transferred to New York City is transferred by railcar, tankcar, and pipeline; these are neither Q nor W.

Additional complications are introduced to confuse the student who tries to synthesize the bits and pieces. The ΔE in Eq. (4) is often labeled as *internal energy*, but the typical misstatement of the work-energy theorem tells the student that $W = \Delta(K.E.) \neq \Delta E_{internal}$. If Eq. (4) is to be tied to the first law of thermodynamics, it must represent the total change in energy of the system, not solely the change in internal energy.

Furthermore, many authors fail to distinguish between *internal energy* and *thermal energy*. The former includes such components as chemical energy and strain energy (as in a deformed spring); the latter is the portion of internal energy that changes reversibly with temperature change.[7]

One must wonder how students are expected to make the leap from the first-law equation to the conservation principle.

Summary

Terminology carried over from caloric theory of two centuries ago encourages misconceptions and leads to confusing notations. Although *heat* has replaced *caloric*, the model of an identifiable fluid (often misrepresented by Q) that moves around remains. Careless terminology also leads to confusion between conservation laws, constants of the motion, and state functions.

The "first-law equation," $\Delta E = Q + W$, is often confused with the first law, so that students are never shown the tenuous connection. Also, this E is confused with internal energy, which in turn is confused with thermal energy. It is then not surprising that students find the subject difficult to grasp.

References
1. Jacob Bronowski, quoted in Robert Karplus, *Introductory Physics; a Model Approach* (W.A. Benjamin, Inc., New York, 1969), p. 461.
2. Robert P. Bauman, "Physics that textbook writers usually get wrong; I. Work," *Phys. Teach.* **30**, 264 (1992).
3. For many years a quantity known as *heat content* was included among the thermodynamic functions. It was equal (apart from a constant for each substance) to the integral, $\int_0^T q = \int_0^T C_p dT$, plus added Q values for phase transitions. As it became widely recognized that this integral did not represent the quantity b, and that the name was more misleading than helpful, the name was dropped. The quantity is now called *enthalpy*, defined as $E + PV$.
4. Rest mass does have the property of *invariance*, meaning that the rest mass of a particle is the same in every inertial reference frame. [See, for example, Edwin F. Taylor and John Archibald Wheeler, *Spacetime Physics* (Freeman, San Francisco, 1963), p. 134.] This is not the same as conservation of rest mass.
5. If W is defined as work done by the system, $\Delta E = Q - W$. This changes the sign wherever the symbol W appears, but makes no other change in the equations or predictions.
6. See Ref. 1 and Bruce Arne Sherwood and W.H. Bernard, "Work and heat transfer in the presence of sliding friction," *Am. J. Phys.* **52**, 1001 (1984).
7. Robert P. Bauman, *Modern Thermodynamics with Statistical Mechanics* (Macmillan, New York, 1992), pp. 14, 473.