

1.3: Energy-Interaction Model

Energy Conservation

The scientific meaning of energy is rather tricky to convey in a sentence or two. There is a good reason for this: energy is an *abstract concept* that took scientists a long time to figure out. Although the concept of energy is truly universal in the sense that energy *changes* are associated with nearly all phenomena and processes, energy is not related to a single property of matter. For example, we all have an intuitive sense of “hotness” and we associate the concept of temperature with this property of matter. We associate the concept of force with the intuitive notion of push and pull. Energy, on the other hand, is associated with many properties or conditions of matter including temperature, force, motion, atomic level, mass, charge, and on and on. It is the fact that energy is so universal that makes it so difficult to define it precisely.

Another reason energy is difficult to pin down, is that the value of energy itself is seldom of importance; rather, it is the *changes* in the values of energy that seem to matter. In fact, we will see that *change in energy* is directly related to “how much” interaction occurred. This is the root of the *Energy-Interaction Model*, it helps explain how energy changes when an interaction occurs. One of the most fundamental conservation law in physics, is the *conservation of energy*, which is the main focus of this course. Mathematically, the statement of conservation of energy is written this way:

$$\Delta E_{total} = Q + W \quad (1.3.1)$$

The left side of the equation represent the sum of all the types of energy that changed in the physical system during a particular interaction, $\Delta E_{tot} = \sum \Delta E$. The right side of the equation represents all the ways that energy can be either added or removed from the physical system. Energy transfer as *heat*, Q , occurs when the system is interacting with an environment which is at a different temperature. Energy transfer as *work*, W , represents all other types of energy transfer that have to do with net force exerted on or by the system. When energy is transferred *to* the system, heat or work is *positive* resulting in an *increase* in the total energy of the system. Likewise, when energy is transferred *from* the system, heat or work is *negative* resulting in an *decrease* in the total energy of the system.

The following are useful statements about energy. Taken together they constitute a working definition of the concept of energy. This concept will be developed further throughout the course.

Statements that help define energy:

- Energy is an abstract concept that characterizes the interactions of matter.
- The *change* in energy of a physical system is a *quantifiable measure* of the degree of its interactions with other physical systems.

Here are some important things to think about when considering energy:

- Historically, many “forms” or “kinds” of energy have been identified. Sometimes, the form of energy labels the interaction that resulted in a change of energy, e.g., chemical, mechanical, or nuclear. It is often misleading to think of there being different forms, kinds, or types of energy, even though we have these “names of energy” in our vocabulary. Energy is energy, regardless of how it manifests itself. Thus, although kinetic energy and chemical energy might be describing a different physical properties of a system, the idea of energy is equivalent.
- Energy is not a *real physical thing* that resides in systems, even though we often talk about it as if it were a “real thing”. Be careful here. Remember there is nothing at all physical about this “thing” we call energy, even when our language sometimes suggests that there is.
- Interactions are modeled by treating energy as something that resides in (can be identified with) particular type of energy. When interactions occur, energies change. This changes in energy might result in an increase or decrease of a particular energy type or a change from one type of energy to another.
- When an interaction occurs, there is a change in one or more energy types. For each energy form that changes, there is an observable (in the sense of being detectable) and quantifiable change in some part of the physical system. The magnitude of the change of each type of energy is determined by the change in its associated observable parameter. We will refer to this observable parameter as the *indicator* associated with a particular type of energy. The questions of how to “divide up” the energy of some physical system and which type of energy to include in the model get to the heart of the modeling process.

Alert

The importance of always being conscious of the word “*change*” as in the phrase “change in energy” and of the way this idea is represented mathematically cannot be overstressed. We signify “change in energy” mathematically by writing “ ΔE ”. The uppercase Greek symbol “ Δ ” often is used in science to indicate a change in some quantity. Does “ ΔE ” look like “E” to you? Of course not, and they have very different physical meanings. A very common misconception is to treat these quantities equally:

CORRECT: $\Delta E_1 + \Delta E_2 + \Delta E_3 + \dots =$ (net energy transferred into or out of the physical system)

and

INCORRECT: $E_1 + E_2 + E_3 + \dots =$ (net energy transferred into or out of the physical system)

Only the first statement makes any physical sense.

But if we wrote the second one as:

CORRECT: $E_1 + E_2 + E_3 = E_{total}$

it would make perfect sense.

Yet the statement:

INCORRECT: $\Delta E_1 + \Delta E_2 + \Delta E_3 = E_{total}$

would make absolutely no sense.

The point is, whether we are talking about a change in energy or the value of the energy, it *matters a lot*. You must be always conscious of which you are thinking about and why it is the one you want to think about!

Physical System

We will be analyzing various phenomena as asking the questions, “Is energy conserved?” and “What can we learn from the phenomena using the concept of conservation of energy?”. The actual question, “Is energy conserved?”, does not have a meaning unless we first define the physical system whose energy we are trying to analyze.

For example, you throw a piece of ice into a drink, place in an insulated container which is initially at room temperature and ask whether the energy of ice is conserved. The answer will be “no”, since you observe that the ice will eventually melt. Does this mean that you violated the conservation of energy? Of course, not! Conservation of energy tells us if the energy of your system changed (ice in this example), that means that the energy had to be transferred from the environment (drink in this example). In our ice example, when ice melted its energy increased. Energy conservation then tells us that the energy was transferred to the ice. On the other hand, if we defined our physical system as the ice and the drink combined, then indeed the energy of our system would be conserved. How do we know? If our system is not interacting with other physical systems (the drink and ice are in an *insulated* container), energy is not entering or leaving the system, $Q = W = 0$. Thus, we can conclude that energy is conserved, $\Delta E_{tot} = 0$. Thus, the amount of energy that increased in the ice must equal to the amount of energy that decreased in the drink.

In reality no system is perfectly *insulated*, but it is a valid approximation to assume energy conservation (or lack of interaction with the environment) within a given time range. If you wait long enough, the drink will eventually come to thermal equilibrium with the environment outside the container. Thus, not only the definition of the physical system is necessary in order to analyze energy conservation, but also the time interval over which the interactions are being analyzed.

Interval

The *time interval* that needs to be considered depends on the question being asked. In situations involving thermal systems, the typical question is, “What is the final state of the physical system once thermal equilibrium is reached?”. *Thermal equilibrium* is a state when heat is no longer flowing since all the substances involved in the interaction have reached the same temperature. Thus, thermal equilibrium is often the logical final state for this type of interaction. Or you might want to ask how much ice to add to a drink in order to have a desired final temperature. In physics we often make approximations since it is easier to apply our models to simplified physical situations. You can learn a lot about nature but first starting with the most simplified situations and then slowly building upon them. For example, the drink in an insulated container will reach room temperature if you wait long enough. However, the drink and the ice will reach equilibrium much faster with each other, than the mixture will with the room outside. Thus, on the short time scale of the drink coming to equilibrium with the ice, we can assume that the interactions with the room are negligible.

One of the frequent ways beginning students go astray is to not be perfectly clear in their own mind how the beginning and end of a process or interaction is being determined. Often there is a point in time that can be identified with the beginning and the end of the interaction. This is not usually an actual clock time, such as ten minutes after one, but rather, something happens at a particular time. The "something" that happens can be easily pictured in our mind and remembered. It is when the "ice was placed in the drink", or "ball was thrown", or "just before the ball hit the ground". Or perhaps it was when "the coffee pot was turned on" or when "the coffee pot was turned off". Something physically happens that we make a conscious decision to identify with the beginning of the time interval and something else happens that we use to identify with the end of the time interval that corresponds to the interaction or process we are interested in. It is crucial to always clearly identify these starting and ending "events," which precisely determine the interval over which the process or interaction occurs. This will then help us determine what type and how much energy changed during this interval.

Closed and Open Physical Systems

There are two useful ways to express the principle of conservation of energy, one corresponding to closed physical systems and the other to open physical systems.

- **Closed physical system:** in a *closed physical system*, there is no transfer of energy into (or out of) that physical system from some other physical system. (This condition would generally preclude mass transfers as well, since mass transfers would also result in energy transfers.) Another way to state the condition of being closed is that all interactions occur *within* the identified physical system, and there are no interactions with other physical systems. For a closed system the statement of energy conservation becomes:

$$\Delta E_{total} = 0 \quad (1.3.2)$$

- **Open physical system:** in an *open physical system*, there can be a *net* transfer of energy into (or out of) that physical system from some other physical system. Another way to state the condition of being open is that interactions can occur between matter in the identified physical system and matter in other physical systems (which might simply be the *environment*). For an open system the change in total energy is given by:

$$\Delta E_{total} = Q + W \quad (1.3.3)$$

In ice/drink example, the ice together with the drink placed in an insulated container would constitute a *closed physical system*. If you choose to define the physical system as just the ice, this will then be an *open physical system* which is interacting with other systems, in this case the drink.

Energy Transfers: Heat Q and Work W

When energy is transferred into a *physical system* from *another physical system*, it is customary to name the energy transferred as either *heat* or *work*. The name heat, Q , is given to energy transfers that occur as result of a difference in temperatures. Heat "flows" from the physical system at the higher temperature to the physical system at the lower temperature.

Alert

Historically, the word "heat" was also used to mean what is now more commonly called thermal energy. If you are reading other textbooks, especially older ones, be sure you understand how the authors are using the word "heat". This is important, because the concept of energy transfer as heat is very different from the concept of thermal energy. We (and most modern authors) now restrict the use of the word "heat" to its meaning as a transfer of energy between physical systems as a result of temperature differences. Thus, stating that a physical system contains some amount of heat does not make sense. As a result, the notation ΔQ will never be used, since this implies a change in a quantity, which Q is not. However, the notation dQ will be used to mean an infinitesimally small amount of energy transferred as heat.

The term work, W , is used to describe the energy transferred between physical systems (objects), which exert forces on one another and move relative to one another. We will examine this type of energy transfer in Chapter 2.

One of the great advantages of an energy model is that we do not have to be concerned with the *details* of how energy transfers occur. We do not need to have a microscopic explanation, for example, of how friction causes increases in thermal energy and decreases in other non-thermal energies. On a microscopic scale, all kinds of energy transfers are taking place between individual atoms and molecules. The energy transfers we are talking about that occur between different physical systems are always the *net* transfers that occur as a result of an interaction between those physical systems.

Energy Changes

State of a Physical System

Related to the discussion regarding the beginning and ending of a process or interaction is the idea also mentioned previously that we are not concerned about the details of the interaction in an energy conservation approach. In fact, all we care about is how the *state of the system* changed from the beginning to the end of the process or interaction, but not what happened in between. By state of the physical system we mean the values of certain parameters that changed. For the *Energy-Interaction Model* we care only about the *indicators* that tell us how much the energy changed in each form of energy. This notion of state of the physical system will become more obvious as we work through more and varied phenomena using the *Energy-Interaction Model*.

Energy Types Related to Thermal and Chemical Processes

When dealing with thermal and chemical processes from a *macroscopic* perspective, it is convenient and useful to define energy types that correspond to the empirically determined enthalpies (the amount of heat added or removed from the substance), ΔH s, that correspond to the process or interaction. These include enthalpies associated with physical phase changes and with the formation of various molecular species. The indicator for the energy associated with these processes would be the *amount of substance* that changed phase or for a chemical reaction, the *amount of substance* that was formed or that “disappeared”. We refer to this energy type as *bond energy*, E_{bond} . The general form of the expression for the amount of energy change in these processes is:

$$\Delta E_{bond} = \pm |\Delta m \Delta H| \quad (1.3.4)$$

The indicator for the energy change is the *amount of substance* that changed, Δm . Energy has units of *joules*, J . Thus, when Δm will have units of kg , ΔH has units of J/kg . If the amount of substance is a molar quantity, then Δm will have units of moles, mol , and is more commonly written as Δn , and ΔH has units of J/mol . We will discuss the \pm sign in the equation shortly.

When the process is either melting or freezing, the enthalpy of melting, or *heat of melting*, is defined as ΔH_{melt} . When the process is either vaporization or condensation, the enthalpy of vaporization, or *heat of vaporization*, is defined as ΔH_{vap} .

When there are no phase changes or chemical reactions occurring and heat is added to a substance, its temperature changes. The amount of temperature change depends on the substance's *heat capacity*. Heat capacity is a quantity that determine how much energy is transferred to a system for a given temperature change:

$$C = \frac{Q}{\Delta T} \quad (1.3.5)$$

We call the type of energy that changes when temperature is changing *thermal energy*, $E_{thermal}$. When only temperature is changing during a process, $\Delta E_{total} = \Delta E_{thermal}$, and if no work being transferred, $W = 0$, Equation 1.3.1 simplifies to $\Delta E_{thermal} = Q$. Combining this with Equation 1.3.5 we can write thermal energy as:

$$\Delta E_{thermal} = C \Delta T \quad (1.3.6)$$

Temperature is the indicator for this energy. The SI units of temperature is *kelvin*, K , thus, heat capacity has units of J/K . Sometimes, the units of *celcius*, $^{\circ}C$, are used, where the conversion between celcius and kelvin is $^{\circ}C = K - 273$. Thus, 273K corresponds to $0^{\circ}C$.

When temperature of the system increases, so does its thermal energy, and vice versa. Heat capacity depends on the amount of substance you have, it requires more energy to increase the temperature by a given amount of a larger quantity of a substance.

The amount of energy required to increase the temperature of a substance also depends on the type of substance. It is useful to define a quantity which only depends on the type of substance, but not on the amount. This is know as the *specific heat*. For example, it takes more energy to change the temperature of water compared to other common substances. This explains why coastal regions do not have large fluctuations in temperature between day and night. The water's high specific heat does not allow a big drop in temperature when the sun sets. On the other hand, sand has a low specific heat, explaining the large variations in temperature between day and night in a desert.

When the specific heat is given per unit mass, the heat capacity can be written as:

$$C = mc_m \quad (1.3.7)$$

where m is mass in units of kg , and c_m is the specific heat in units of $\frac{J}{K \cdot kg}$. When the specific heat is given per amount of moles, sometimes known as *molar heat capacity*, the heat capacity can be written as:

$$C = nc_{mol} \quad (1.3.8)$$

where n is the number of moles and c_{mol} is molar heat capacity in units of $\frac{J}{K \cdot mol}$.

Separating internal energies into bond and thermal energies is very useful for three reasons:

1. they correspond to changes in indicators we can directly observe.
2. they depend on parameters, ΔH and C , that are tabulated for most substances.
3. During a physical or chemical reaction, only one of them will be changing at a given time.

Both phase changes and chemical reactions involve changes in bond energy that correspond to making and breaking of atomic and/or molecular bonds. Temperature changes corresponding to random atomic motions, results in thermal energy changes. The two types of energy we have defined above are closely related to bond energy and thermal energy we will develop from an atomic perspective in Chapter 3, but are not exactly the same. However, rather than use a different name for the energies involving the empirically determined parameters, ΔH and C , and macroscopic indicators, Δm and ΔT , we will simply refer to them as the bond energy and thermal energy, remembering that we will refine our understanding of these processes in Chapter 3 and see how the macroscopic energies we define here relate to the microscopically defined bond and thermal energies using particle models.

The algebraic signs of thermal and bond energies

It is very important to make sense of the algebraic sign of the change in energy based on what physically is happening. This is actually very simple to do, once you understand the role of an indicator. However, it is easy to make a simple algebraic slip-up when actually calculating numerical values. You should always check to see if the final algebraic sign makes sense.

Any thermal energy for which the temperature increases during the process will always have a positive change in energy. Likewise, any thermal energy for which the temperature decreases will always have a negative change in energy. This is consistent with the simple notion that thermal energy increases with increases in temperature, because at higher temperatures there is “more vigorous” motion of the particles. When doing numerical calculations the sign in the change of thermal energy will correspond to the sign in the change in temperature, as indicated by Equation 1.3.6.

Any bond energy for which bonds are broken during the process or interaction will always have a positive change in energy. Likewise, any system for which bonds are formed during the process or interaction will always have a negative change in energy. This is consistent with the common experience of having to *add energy* (and thus increase the bond energy) to vaporize liquid water that is at 100°C . The bonds that had existed in the liquid phase disappear (are broken) in the process of the liquid changing to a vapor. You will probably need to struggle mentally with this last point: broken bonds have more energy than intact bonds. Work on this until it “seems obvious” to you. Therefore, when the substance is either melting or vaporizing the change in bond energy is positive (a plus sign is chosen on the right-hand-side of Equation 1.3.4). When the substance is either freezing or condensing the change in bond energy is negative (a minus sign is chosen on the right-hand-side of Equation 1.3.4).

Alert

A very common misconception is to think that bond energy should increase when bonds are formed and decrease when bonds break. In other words, bond energy is misinterpreted as the the energy stored in bonds. It is important to analyze each energy term in terms of the *Energy-Interaction Model*, which states that if energy is transferred to a system, the system's energy must increase, and vice versa. The *Three-Phase Model* tells us that energy needs to be added to a substance when it is vaporizing from a liquid to a gas, and removed when it is condensing from a gas to a liquid. Thus, the system's energy (bond energy in this case) increases as bonds break and decreases as bonds form.

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