

35 Temperature, Internal Energy, Heat, and Specific Heat Capacity

As you know, temperature is a measure of how hot something is. Rub two sticks together and you will notice that the temperature of each increases. You did work on the sticks and their temperature increased. Doing work is transferring energy. So you transferred energy to the sticks and their temperature increased. This means that an increase in the temperature of a system is an indication of an increase in the internal energy (a.k.a. thermal energy) of the system. (In this context the word system is thermodynamics jargon for the generalization of the word object. Indeed an object, say an iron ball, could be a system. A system is just the subject of our investigations or considerations. A system can be as simple as a sample of one kind of gas or a chunk of one kind of metal, or it can be more complicated as in the case of a can plus some water in the can plus a thermometer in the water plus a lid on the can. For the case at hand, the system is the two sticks.) The internal energy of a system is energy associated with the motion of molecules, atoms, and the particles making up atoms relative to the center of mass of the system, and the potential energy corresponding to the positions and velocities of the aforementioned submicroscopic constituents of the system relative to each other. As usual with energy accounting, the absolute zero of energy in the case of internal energy doesn't matter—only changes in internal energy have any relevance. As such, you or the publisher of a table of internal energy values (for a given substance, publishers actually list the internal energy per mass or the internal energy per mole of the substance under specified conditions rather than the internal energy of a sample of such a substance), are free to choose the zero of internal energy for a given system. In making any predictions regarding a physical process involving that system, as long as you stick with the same zero of internal energy throughout your analysis, the measurable results of your prediction or explanation will not depend on your choice of the zero of internal energy.

Another way of increasing the temperature of a pair of sticks is to bring them into contact with something hotter than the sticks are. When you do that, the temperature of the sticks automatically increases—you don't have to do any work on them. Again, the increase in the temperature of either stick indicates an increase in the internal energy of that stick. Where did that energy come from? It must have come from the hotter object. You may also notice that the hotter object's temperature decreased when you brought it into contact with the sticks. The decrease in temperature of the hotter object is an indication that the amount of internal energy in the hotter object decreased. You brought the hotter object in contact with the sticks and energy was automatically transferred from the hotter object to the sticks. The energy transfer in this case is referred to as the flow of heat. Heat is energy that is automatically transferred from a hotter object to a cooler object when you bring the two objects in contact with each other. Heat is not something that a system has but rather energy that is transferred or is being transferred. Once it gets to the system to which it is transferred we call it internal energy. The idea is to distinguish between what is being done to a system, "Work is done on the system and/or heat is caused to flow into it", with how the system changes as a result of what was done to it, "The internal energy of the system increases."

The fact that an increase in the temperature of an object is an indication of energy transferred to that object might suggest that anytime you transfer energy to an object its temperature increases. But this is not the case. Try putting a hot spoon in a glass of ice water. (Here we consider a case

for which there is enough ice so that not all of the ice melts.) The spoon gets as cold as the ice water and some of the ice melts, but the temperature of the ice water remains the same ($0\text{ }^{\circ}\text{C}$). The cooling of the spoon indicates that energy was transferred from it, and since the spoon was in contact with the ice water the energy must have been transferred to the ice water. Indeed the ice does undergo an observable change; some of it melts. The presence of more liquid water and less ice is an indication that there is more energy in the ice water. Again there has been a transfer of energy from the spoon to the ice water. This transfer is an automatic flow of heat that takes place when the two systems are brought into contact with one another. Evidently, heat flow does not always result in a temperature increase.

Experiment shows that when a higher temperature object is in contact with a lower temperature object, heat is flowing from the higher temperature object to the lower temperature object. The flow of heat persists until the two objects are at one and the same temperature. We define the average translational kinetic energy of a molecule of a system as the sum of the translational kinetic energies of all the molecules making up the system divided by the total number of molecules. When two simple ideal gas systems, each involving a multitude of single atom molecules interacting via elastic collisions, are brought together, we find that heat flows from the system in which the average translational kinetic energy per molecule is greater to the system in which the average translational kinetic energy per molecule is lesser. This means the former system is at a higher temperature. That is to say that the higher the translational kinetic energy, on the average, of the particles making up the system, the higher the temperature. This is true for many systems.

Solids consist of atoms that are bound to neighboring atoms such that molecules tend to be held in their position, relative to the bulk of the solid, by electrostatic forces. A pair of molecules that are bound to each other has a lower amount of internal potential energy relative to the same pair of molecules when they are not bound together because we have to add energy to the bound pair at rest to yield the free pair at rest. In the case of ice water, the transfer of energy into the ice water results in the breaking of bonds between water molecules, which we see as the melting of the ice. As such, the transfer of energy into the ice water results in an increase in the internal potential energy of the system.

The two different kinds of internal energy that we have discussed are internal potential energy and internal kinetic energy. When there is a net transfer of energy into a system, and the macroscopic mechanical energy of the system doesn't change (e.g. for the case of an object near the surface of the earth, the speed of object as a whole does not increase, and the elevation of the object does not increase), the internal energy (the internal kinetic energy, the internal potential energy, or both) of the system increases. In some, but not all, cases, the increase in the internal energy is accompanied by an increase in the temperature of the system. If the temperature doesn't increase, then we are probably dealing with a case in which it is the internal *potential* energy of the system that increases.

Heat Capacity and Specific Heat Capacity

Let's focus our attention on cases in which heat flow into a sample of matter is accompanied by an increase in the temperature of the sample. For many substances, over certain temperature ranges, the temperature change is (at least approximately) proportional to the amount of heat that flows into the substance.

$$\Delta T \propto Q$$

Traditionally, the constant of proportionality is written as $\frac{1}{C}$ so that

$$\Delta T = \frac{1}{C} Q$$

where the upper case C is the heat capacity. This equation is more commonly written as

$$Q = C \Delta T \quad (35-1)$$

which states that the amount of heat that must flow into a system to change the temperature of that system by ΔT is the heat capacity C times the desired temperature change ΔT . Thus the heat capacity C is the "heat-per-temperature-change." Its reciprocal is a measure of a system's temperature sensitivity to heat flow.

Let's focus our attention on the simplest kind of system, a sample of one kind of matter, such as a certain amount of water. The amount of heat that is required to change the temperature of the sample by a certain amount is directly proportional to the mass of the single substance; e.g., if you double the mass of the sample it will take twice as much heat to raise its temperature by, for instance, 1 $^{\circ}\text{C}$. Mathematically, we can write this fact as

$$C \propto m$$

It is traditional to use a lower case c for the constant of proportionality. Then

$$C = cm$$

where the constant of proportionality c is the heat-capacity-per-mass of the substance in question. The heat-capacity-per-mass c is referred to as the mass specific heat capacity or simply the *specific heat capacity* of the substance in question. (In this context, the adjective specific means "per amount." Because the amount can be specified in more than one way we have the expression "mass specific" meaning "per amount of mass" and the expression "molar specific" meaning "per number of moles." Here, since we are only dealing with mass specific heat, we can omit the word "mass" without generating confusion.) The specific heat capacity c has a different value for each different kind of substance in the universe. (Okay, there might be some coincidental duplication but you get the idea.) In terms of the mass specific heat capacity,

equation 35-1 ($Q = C \Delta T$), for the case of a system consisting only of a sample of a single substance, can be written as

$$Q = mc\Delta T \quad (35-2)$$

The specific heat capacity c is a property of the kind of matter of which a substance consists. As such, the values of specific heat for various substances can be tabulated.

Substance	Specific Heat Capacity*
	$\left[\frac{\text{J}}{\text{kg} \cdot \text{C}^\circ} \right]$
Ice (solid water)	2090
Liquid Water	4186
Water Vapor (gas)	2000
Solid Copper	387
Solid Aluminum	902
Solid Iron	448

*The specific heat capacity of a substance varies with temperature and pressure. The values given correspond to atmospheric pressure. Use of these representative constant values for cases involving atmospheric pressure and temperature ranges between -100°C and $+600^\circ\text{C}$, as applicable for the phase of the material, can be expected to yield reasonable results but if precision is required, or information on how reasonable your results are is needed, you should consult a thermodynamics textbook and thermodynamics tables and carry out a more sophisticated analysis.

Note how many more Joules of energy are needed to raise the temperature of 1 kg of liquid water 1°C than are required to raise the temperature of 1 kg of a metal 1°C .

Temperature

Despite the fact that you are quite familiar with it, some more discussion of temperature is in order. Whenever you measure something, you are really just comparing that something with an arbitrarily-established standard. For instance, when you measure the length of a table with a meter stick, you are comparing the length of the table with the modern day equivalent of what was historically established as one ten-thousandth of the distance from the earth's north pole to the equator. In the case of temperature, a standard, now called the "degree Celsius" was established as follows: At 1 atmosphere of pressure, the temperature at which water freezes was defined to be 0°C and the temperature at which water boils was defined to be 100°C . Then a substance with a temperature-dependent measurable characteristic, such as the length of a column of liquid mercury, was used to interpolate and extrapolate the temperature range. (Mark the position of the end of the column of mercury on the tube containing that mercury when it is at the temperature of freezing water and again when it is at the temperature of boiling water. Divide the interval between the two marks into a hundred parts. Use the same length of each of those parts to extend the scale in both directions and call it a temperature scale.)

Note the arbitrary manner in which the zero of the Celsius scale has been established. The choice of zero is irrelevant for our purposes since equations 35-1 ($Q = C \Delta T$) and 35-2

($Q = mc\Delta T$) relate temperature *change*, rather than temperature itself, to the amount of heat flow. An absolute temperature scale has been established for the SI system of units. The zero of temperature on this scale is set at the greatest possible temperature such that it is theoretically impossible for the temperature of any system in equilibrium to be as low as the zero of the Kelvin scale. The unit of temperature on the Kelvin scale is the kelvin, abbreviated K. Note the absence of the degree symbol in the unit. The Kelvin scale is similar to the Celsius scale in that a change in temperature of, say, 1 K, is equivalent to a change in temperature of 1 C°. (Note regarding units notation: The units °C are used for a *temperature* on the Celsius scale, but the units C° are used for a *temperature change* on the Celsius scale.)

On the Kelvin scale, at a pressure of one atmosphere, water freezes at 273.15 K. So, a temperature in kelvin is related to a temperature in °C by

$$\text{Temperature in K} = (\text{Temperature in } ^\circ\text{C}) \cdot \left(\frac{1\text{K}}{^\circ\text{C}} \right) + 273.15 \text{ K}$$