Heat & Temperature

Knowledge/Understanding Goals:
- the difference between heat and temperature
- thermal equilibrium

Notes:
- **heat**: energy that can be transferred by moving atoms or molecules via transfer of momentum.
- **temperature**: a measure of the average kinetic energy of the particles (atoms or molecules) of a system.
- **thermometer**: a device that measures temperature, most often via thermal expansion and contraction of a liquid or solid.

Note that heat is the energy itself, whereas temperature is a measure of the quality of the heat—the average of the kinetic energies of the individual molecules.

Suppose you have a hot substance. If you increase the amount of the substance (without changing the temperature), you have increased the total amount of heat (energy) in the system. However, even though you have more of the substance and therefore more total energy, the temperature did not change.

Any natural (spontaneous) transfer of energy always goes from a higher-energy state to a lower-energy state. (For example, with respect to gravity, objects spontaneously roll downhill, not uphill.)

Heat always transfers from objects with a higher temperature (more kinetic energy) to objects with a lower temperature (less kinetic energy).

- **system**: the region being considered in a problem.
- **surroundings**: everything that is outside of the system.

*E.g.*, if a metal block is heated, we would most likely define the **system** to be the block, and the **surroundings** to be everything outside of the block.

- **thermal equilibrium**: when all of the particles in a system have the same average kinetic energy (temperature). When a system is at thermal equilibrium, no net heat is transferred. (*I.e.*, collisions between particles may still transfer energy, but the average temperature of the particles in the system—what we measure with a thermometer—is not changing.)

Heat Flow

We generally use the variable $Q$ to represent heat in physics equations.

Heat flow is always represented in relation to the **system**.
Heat Flow

<table>
<thead>
<tr>
<th>Heat Flow</th>
<th>Sign of $Q$</th>
<th>System</th>
<th>Surroundings</th>
</tr>
</thead>
<tbody>
<tr>
<td>from surroundings to system</td>
<td>+ (positive)</td>
<td>gains heat (gets warmer)</td>
<td>lose heat (get colder)</td>
</tr>
<tr>
<td>from system to surroundings</td>
<td>- (negative)</td>
<td>loses heat (gets colder)</td>
<td>gain heat (get hotter)</td>
</tr>
</tbody>
</table>

A positive value of $Q$ means heat is flowing into the system. Because the heat is transferred from the molecules outside the system to the molecules in the system, the temperature of the system increases, and the temperature of the surroundings decreases.

A negative value of $Q$ means heat is flowing out of the system. Because the heat is transferred from the molecules in the system to the molecules outside the system, the temperature of the system decreases, and the temperature of the surroundings increases.

This can be confusing. Suppose you set a glass of ice water on a table. When you pick up the glass, your hand gets colder because heat is flowing from your hand (which is part of the surroundings) into the system (the glass of ice water). This means the system (the glass of ice water) is gaining heat, and the surroundings (your hand, the table, etc.) are losing heat. The value of $Q$ would be positive in this example.

In simple terms, you need to remember that your hand is part of the surroundings, not part of the system.

The Laws of Thermodynamics

thermodynamics: the study of heat-related (thermal) energy changes (dynamics)

enthalpy: the “usable” heat content of an object or system

entropy: the “unusable” thermal energy in a system. Energy in the form of entropy is unavailable because it has “escaped” or “spread out”. (Entropy is sometimes called “disorder” or “randomness”, but this is not correct in the thermodynamic sense. The entropy of your room is a thermodynamic property, not a commentary on how much dirty laundry is on the floor!) A correct view of entropy requires an analysis of possible energy states and their probabilities, which is beyond the scope of this course.

The laws of thermodynamics state that:

0. If a system is at thermal equilibrium, every component of the system has the same temperature.

1. Heat always flows from a region of higher enthalpy to a region of lower enthalpy. This means you can’t get more heat out of something than you put in.

2. In almost every change, some energy is irretrievably lost to the surroundings. Entropy is a measure of this “lost” energy. The entropy of the universe is always increasing, which means on any practical scale, you always get out less energy than you put in.
3. In any closed system, the total energy (enthalpy + entropy + work) remains constant. If energy was “lost,” it turned into an increase in entropy.