

Specific Heat

Specific Heat

The specific heat is the amount of [heat](#) per unit mass required to raise the [temperature](#) by one degree Celsius. The relationship between heat and temperature change is usually expressed in the form shown below where c is the specific heat. The relationship does not apply if a [phase change](#) is encountered, because the heat added or removed during a phase change does not change the temperature.

$$Q = cm\Delta T$$

heat added specific heat mass change in temperature

The specific heat of water is 1 calorie/gram °C = 4.186 joule/gram °C which is higher than any other common substance. As a result, water plays a very important role in temperature regulation. The specific heat per gram for water is much higher than that for a metal, as described in the [water-metal example](#). For most purposes, it is more meaningful to compare the molar specific heats of substances.

The molar specific heats of most solids at room temperature and above are nearly constant, in agreement with the [Law of Dulong and Petit](#). At lower temperatures the specific heats drop as quantum processes become significant. The low temperature behavior is described by the [Einstein-Debye model](#) of specific heat.

Law of Dulong and Petit

The [specific heat](#) of copper is 0.093 cal/gm K (.389 J/gm K) and that of lead is only 0.031 cal/gm K (.13 J/gm K). Why are they so different? The difference is mainly because it is expressed as energy per unit mass; if you express it as energy per mole, they are very similar. It is in fact that similarity of the molar specific heats of metals which is the subject of the Law of Dulong and Petit. The similarity can be accounted for by applying [equipartition of energy](#) to the atoms of the solids. From just the translational degrees of freedom you get $3kT/2$ of energy per atom. Energy added to solids takes the form of atomic vibrations and that contributes three additional degrees of freedom and a total energy per atom of $3kT$. The specific heat at constant volume should be just the rate of change with temperature (temperature derivative) of that energy.

$$\text{Energy per mole} = 3kTN_A$$

k= Boltzmann's constant
T= Temperature in Kelvins
 N_A = Avogadro's number

The Law of Dulong & Petit

$$C_V = \frac{\partial}{\partial T}(3kTN_A) = 3kN_A / \text{mole} = 24.94 \text{ J / mole K}$$

When looked at on a molar basis, the specific heats of copper and lead are quite similar:

Copper $0.386 \text{ J/gm K} \times 63.6 \text{ gm/mole} = 24.6 \text{ J/mol K}$

Lead $0.128 \text{ J/gm K} \times 207 \text{ gm/mole} = 26.5 \text{ J/mol K}$

Equipartition of Energy

The theorem of equipartition of energy states that molecules in thermal equilibrium have the same average energy associated with each independent degree of freedom of their motion and that the energy is

$$\frac{1}{2}kT \text{ per molecule}$$
$$\frac{1}{2}RT \text{ per mole}$$

k = Boltzmann's constant
R = gas constant

$$\frac{3}{2}kT$$
$$\frac{3}{2}RT$$

For three translational degrees of freedom, such as in an ideal monoatomic gas.

The equipartition result

$$KE_{avg} = \frac{3}{2}kT$$

serves well in the definition of [kinetic temperature](#) since that involves just the translational degrees of freedom, but it fails to predict the [specific heats](#) of polyatomic gases because the increase in internal energy associated with heating such gases adds energy to rotational and perhaps vibrational degrees of freedom. Each vibrational mode will get $kT/2$ for kinetic energy and $kT/2$ for potential energy - equality of kinetic and potential energy is addressed in the [virial theorem](#). Equipartition of energy also has implication for electromagnetic radiation when it is in equilibrium with matter, each mode of radiation having kT of energy in the [Rayleigh-Jeans law](#).

For the translational degrees of freedom only, equipartition can be shown to follow from the [Boltzmann distribution](#).

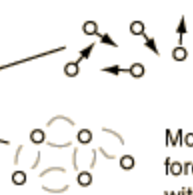
$$KE_{avg} = \left[\overline{\frac{1}{2}mv^2} \right] = \frac{3}{2}kT$$

Internal Energy

Internal energy is defined as the [energy](#) associated with the random, disordered motion of molecules. It is separated in scale from the macroscopic ordered energy associated with moving objects; it refers to the invisible [microscopic energy](#) on the atomic and molecular scale. For example, a room temperature glass of water sitting on a table has no apparent energy, either [potential](#) or [kinetic](#). But on the microscopic scale it is a seething mass of high speed molecules traveling at hundreds of meters per second. If the water were tossed across the room, this microscopic energy would not necessarily be changed when we superimpose an ordered large scale motion on the water as a whole.

Does a glass of water sitting on a table have any energy?

No apparent energy of the glass of water on a macroscopic scale.



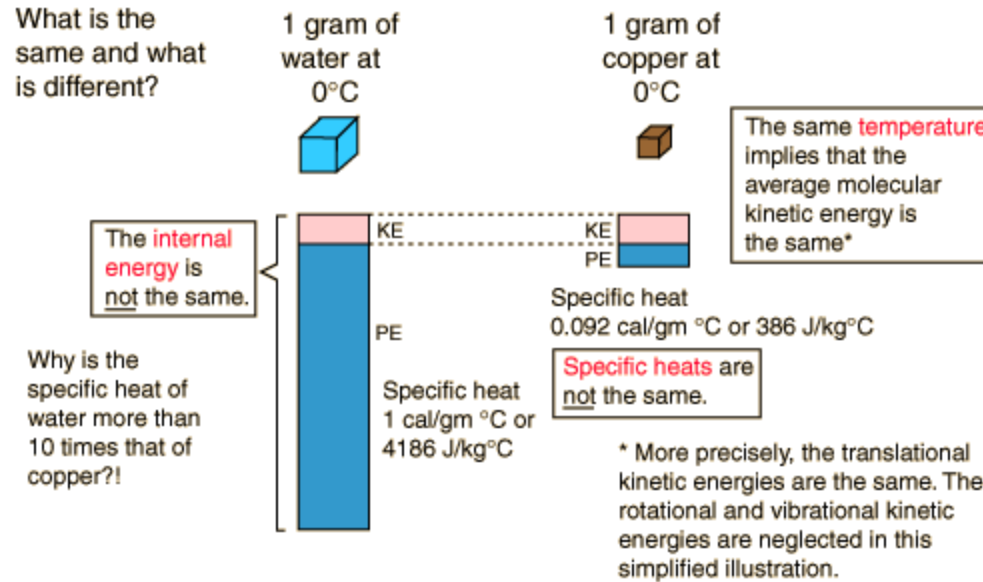
Microscopic kinetic energy is part of internal energy.

Molecular attractive forces are associated with potential energy.

U is the most common symbol used for internal energy.

Related energy quantities which are particularly useful in chemical thermodynamics are [enthalpy](#), [Helmholtz free energy](#), and [Gibbs free energy](#).

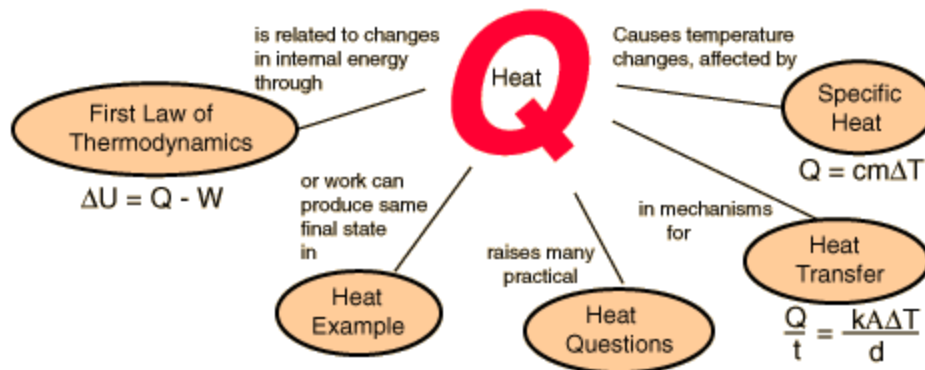
Internal Energy Example



When the sample of water and copper are both heated by 1°C, the addition to the kinetic energy is the same, since that is what temperature measures. But to achieve this increase for water, a much larger proportional energy must be added to the potential energy portion of the internal energy. So the total energy required to increase the temperature of the water is much larger, i.e., its specific heat is much larger.

Heat

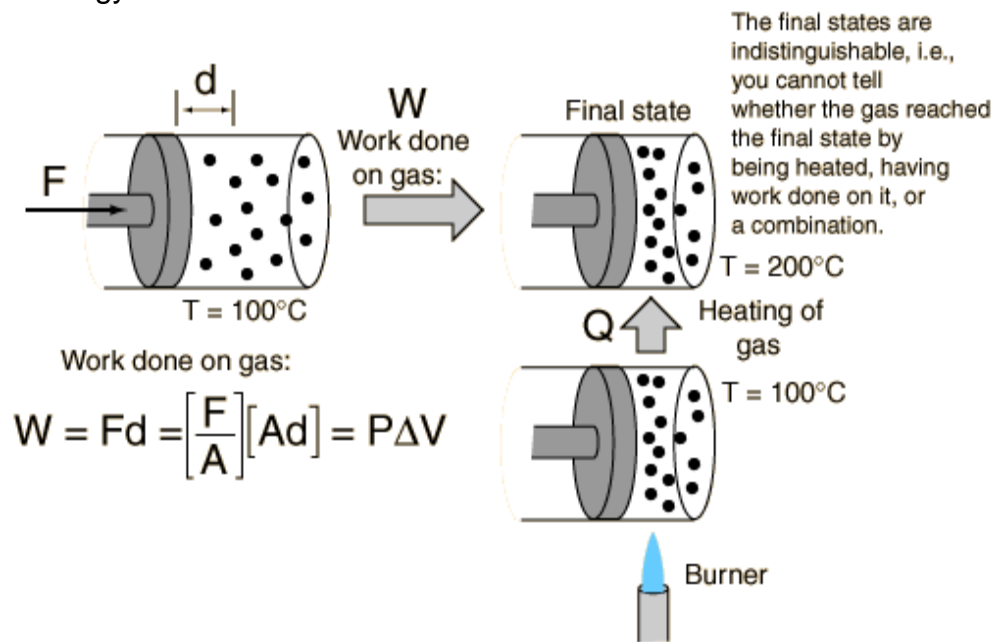
Heat may be defined as energy in transit from a high [temperature](#) object to a lower temperature object. An object does not possess "heat"; the appropriate term for the microscopic energy in an object is [internal energy](#). The internal energy may be increased by transferring energy to the object from a higher temperature (hotter) object - this is properly called heating.



Heat and Work Example

This example of the interchangeability of [heat](#) and [work](#) as agents for adding energy to a system can help to dispel some misconceptions about heat. I found the idea in a little article by Mark Zemansky entitled "The Use and Misuse of the Word 'Heat' in Physics Teaching". One key idea from this example is that if you are presented with a high [temperature](#) gas, you cannot tell whether it reached that high temperature by being heated, or by having work done on it, or a combination of the two.

To describe the energy that a high temperature object has, it is not a correct use of the word heat to say that the object "possesses heat" - it is better to say that it possesses [internal energy](#) as a result of its molecular motion. The word heat is better reserved to describe the process of transfer of energy from a high temperature object to a lower temperature one. Surely you can take an object at low internal energy and raise it to higher internal energy by heating it. But you can also increase its internal energy by doing work on it, and since the internal energy of a high temperature object resides in random motion of the molecules, you can't tell which mechanism was used to give it that energy.



Don't refer to the "heat in a body", or say "this object has twice as much heat as that body". He also objects to the use of the vague term "thermal energy" and to the use of the word "heat" as a verb, because they feed the misconceptions, but it is hard to avoid those terms. He would counsel the introduction and use of the concept of internal energy as quickly as possible.

Zemansky points to the [First Law of Thermodynamics](#) as a clarifying relationship. The First Law identifies both heat and work as methods of energy transfer which can bring about a change in the internal energy of a system. After that, neither the words work or heat have any usefulness in describing the final state of the system - we can speak only of the internal energy of the system.

Mechanical Equivalent of Heat

Heat flow and work are both ways of transferring energy. As illustrated in the heat and work example, the temperature of a gas can be raised either by heating it, by doing work on it, or a combination of the two.

In a classic experiment in 1843, James Joule showed the energy equivalence of heating and doing work by using the change in potential energy of falling masses to stir an insulated container of water with paddles. Careful measurements showed the increase in the temperature of the water to be proportional to the mechanical energy used to stir the water. At that time calories were the accepted unit of heat and joules became the accepted unit of mechanical energy. The British Thermal Unit was also introduced. Their relationship to joules is

$$1 \text{ calorie} = 4.1868 \text{ joules}$$

$$1 \text{ BTU} = 1055.056 \text{ joules}$$

These conversions are the International Steam Table (IT) values, and variations up to 0.5% will be found since they were originally based on energy for temperature changes in water, and that varies a small amount based on what temperature of water is used to establish it. The heat capacity of water does change slightly with temperature.