

Specific Heat Capacity

The definition of Specific Heat Capacity is:

“The amount of energy required to raise the temperature of 1 kg (a unit mass) of the substance by 1 K (a unit temperature rise)”

Symbol: c

Unit: $\text{Jkg}^{-1}\text{K}^{-1}$

$$c = \frac{\Delta Q}{m\Delta\theta}$$

Where:

ΔQ is the heat energy added (or removed)

m is the mass of the substance you are heating (or cooling)

$\Delta\theta$ is the change in temperature.

Note: the value of ‘ c ’ for ice isn’t the same as that for water or for steam. However, the value of ‘ c ’ is the same if you are cooling rather than heating the substance. In other words you get as much heat back out of the substance when you cool it as you put in when you heated it.

Heat Capacity:

The product of mass times specific heat capacity defines the heat capacity of a body.

$$C = mc$$

The concept of heat capacity is useful when a body consists of a number of parts of different specific capacities, such as, for example, a metal tank containing water. If the tank has mass M and specific heat capacity c and the water has mass m and specific heat capacity c^1 , then the heat capacity of the water tank is

$$C = Mc + m c^1$$

Knowing the heat capacity of the water tank allows us to say that, if a quantity of thermal energy Q is given to the water tank, then the rise in temperature ΔT will be found from

$$Q = C \Delta T$$

Specific latent heat capacity of fusion (L_f):

The thermal energy required to melt a unit mass of material at its melting point is called *the specific latent heat of fusion*, L_f . $Q = m L_f$

Specific latent heat of vaporization (L_v):

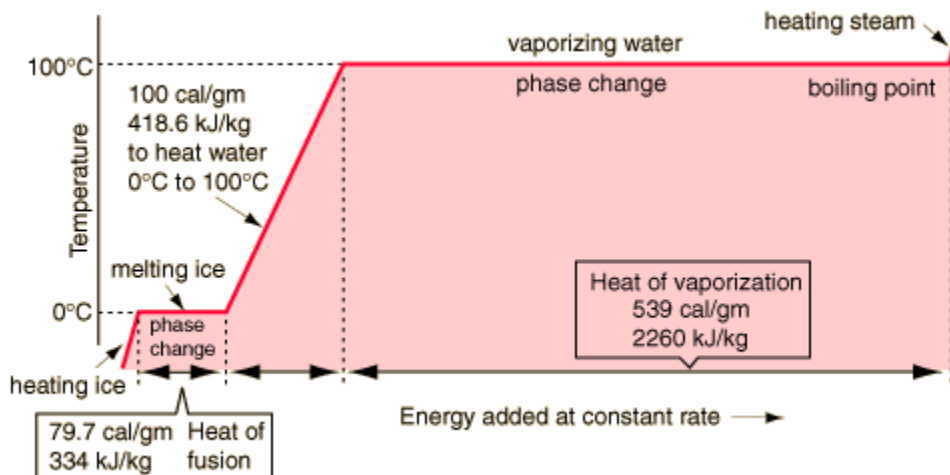
The thermal energy required to boil a unit mass of material at its boiling point is called *the specific latent heat of vaporization*, L_v .

Thus to vaporize a quantity of mass m , we require a quantity of thermal energy

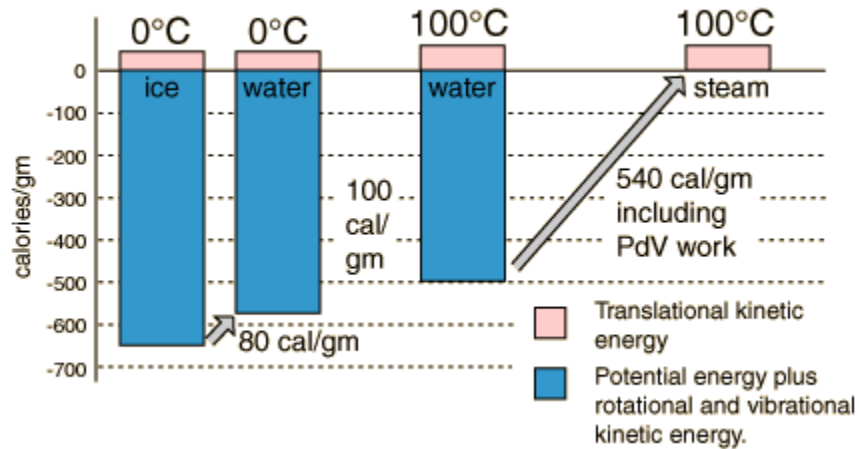
$$Q = m L_v$$

Phase changes

Transitions between solid, liquid, and gaseous phases typically involve large amounts of energy compared to the **specific heat**. If heat were added at a constant rate to a mass of ice to take it through its phase changes to liquid water and then to steam, the energies required to accomplish the phase changes (called the latent **heat of fusion** and latent **heat of vaporization**) would lead to plateaus in the temperature vs time graph. The graph below presumes that the pressure is one **standard atmosphere**.



Energy Involved in the Phase Changes of Water

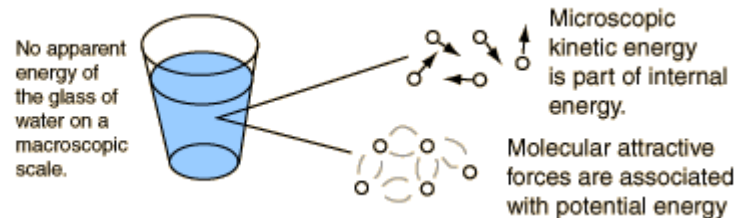


The data for the vaporization phase change presumes that the pressure is one **standard atmosphere**.

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.

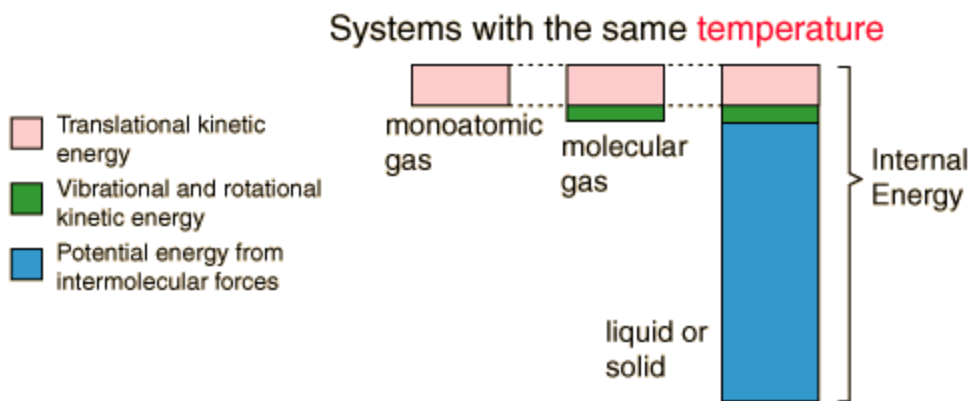
Does a glass of water sitting on a table have any energy?



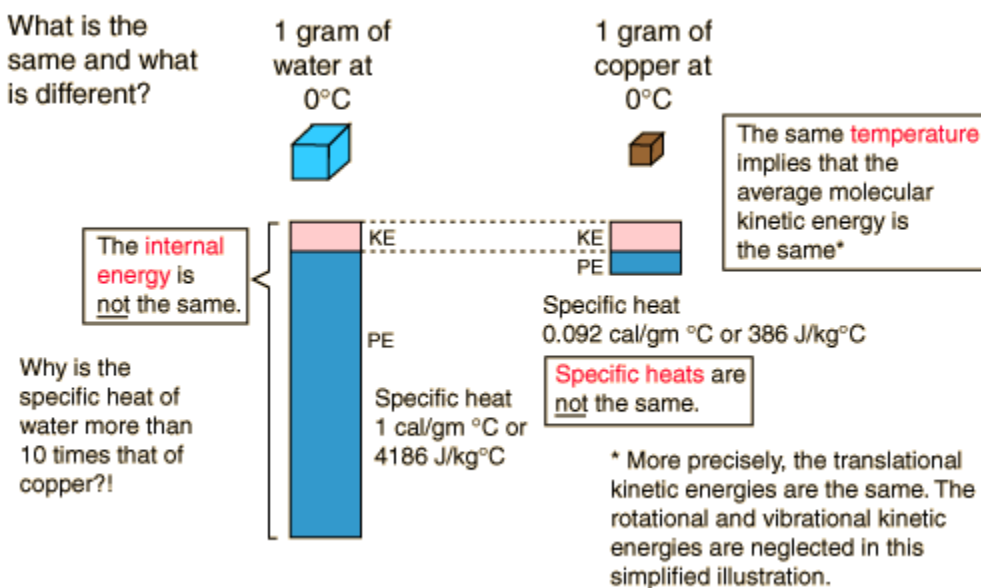
U is the most common symbol used for internal energy.

Microscopic Energy

Internal energy involves energy on the microscopic scale. For an ideal monoatomic gas, this is just the translational kinetic energy of the linear motion of the "hard sphere" type atoms, and the behavior of the system is well described by kinetic theory. However, for polyatomic gases there is rotational and vibrational kinetic energy as well. Then in liquids and solids there is potential energy associated with the intermolecular attractive forces. A simplified visualization of the contributions to internal energy can be helpful in understanding phase transitions and other phenomena which involve internal 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.

First Law of Thermodynamics

The first law of thermodynamics is the application of the **conservation of energy** principle to heat and thermodynamic processes:

The change in internal energy of a system is equal to the heat added to the system minus the work done by the system.

$$\Delta U = Q - W$$

Change in internal energy Heat added to the system Work done by the system

The first law makes use of the key concepts of **internal energy**, **heat**, and **system work**. It is used extensively in the discussion of **heat engines**.

It is typical for chemistry texts to write the first law as $\Delta U = Q + W$. It is the same law, of course - the thermodynamic expression of the conservation of energy principle. It is just that W is defined as the work done on the system instead of work done by the system. In the context of physics, the common scenario is one of adding heat to a volume of gas and using the expansion of that gas to do work, as in the pushing down of a piston in an internal combustion engine. In the context of chemical reactions and process, it may be more common to deal with situations where work is done on the system rather than by it.