

AL - Mustaqpal University
Science College
Dep. Medical Biotechnology

Second Stage

Lec 7

Heat laws
and
conversion of units used in this field

م . م علي سلمان حمادي

Heat laws

Heat is a form of energy related to the movement of atoms and molecules. The amount of heat energy, or the amount of motion of its particles, is measured as temperature. The Second Law of Thermodynamics states that heat energy always moves from a warmer area to a cooler area, and never in the opposite direction.

The **laws of thermodynamics** are a set of scientific laws which define a group of physical quantities, such as temperature, energy, and entropy, that characterize thermodynamic systems in thermodynamic equilibrium. The laws also use various parameters for thermodynamic processes, such as thermodynamic work and heat, and establish relationships between them. They state empirical facts that form a basis of precluding the possibility of certain phenomena, such as perpetual motion. In addition to their use in thermodynamics, they are important fundamental laws of physics in general and are applicable in other natural sciences.

Traditionally, thermodynamics has recognized three fundamental laws, simply named by an ordinal identification, the first law, the second law, and the third law. A more fundamental statement was later labelled as the zeroth law after the first three laws had been established.

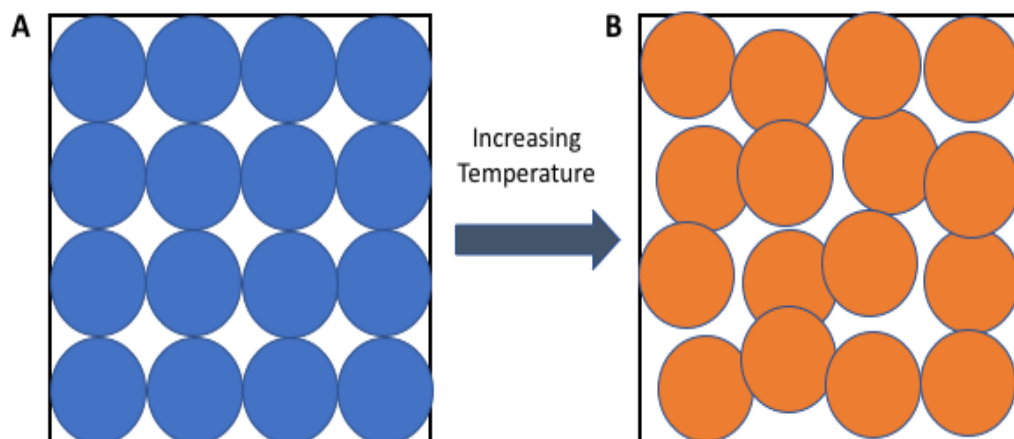
The **zeroth law** of thermodynamics defines thermal equilibrium and forms a basis for the definition of temperature: **If two systems are each in thermal equilibrium with a third system, then they are in thermal equilibrium with each other.**

The **first law** of thermodynamics states that : **When energy passes into or out of a system (as work, heat, or matter), the system's internal energy changes in accordance with the law of conservation of energy.**

The **second law** of thermodynamics states that : **In a natural thermodynamic process, the sum of the entropies of the interacting thermodynamic systems never decreases. A common corollary of the statement is that heat does not spontaneously pass from a colder body to a warmer body.**

The **third law** of thermodynamics states that : **A system's entropy approaches a constant value as the temperature approaches absolute zero. With the exception of non-crystalline solids (glasses), the entropy of a system at absolute zero is typically close to zero.**

The first and second laws prohibit two kinds of perpetual motion machines, respectively: the perpetual motion machine of the first kind which produces work with no energy input, and the perpetual motion machine of the second kind which spontaneously converts thermal energy into mechanical work.



Zeroth law

The zeroth law of thermodynamics provides for the foundation of temperature as an empirical parameter in thermodynamic systems and establishes the transitive relation between the temperatures of multiple bodies in thermal equilibrium. The law may be stated in the following form:

If two systems are both in thermal equilibrium with a third system, then they are in thermal equilibrium with each other.

Though this version of the law is one of the most commonly stated versions, it is only one of a diversity of statements that are labeled as "the zeroth law". Some statements go further, so as to supply the important physical fact that temperature is one-dimensional and that one can conceptually arrange bodies in a real number sequence from colder to hotter.

These concepts of temperature and of thermal equilibrium are fundamental to thermodynamics and were clearly stated in the nineteenth century. The name 'zeroth law' was invented by Ralph H. Fowler in the 1930s, long after the first, second, and third laws were widely recognized. The law allows the definition of temperature in a non-circular way without reference to entropy, its conjugate variable. Such a temperature definition is said to be 'empirical'.

First law

The first law of thermodynamics is a version of the law of conservation of energy, adapted for thermodynamic processes. In general, the conservation law states that : **The total energy of an isolated system is constant**; energy can be transformed from one form to another, but can be neither created nor destroyed.

In a closed system (i.e. there is no transfer of matter into or out of the system), the first law states that the change in internal energy of the system (ΔU_{system}) is equal to the difference between the heat supplied to the system (Q) and the work (W) done by the system on its surroundings.

To define (W) as the work done on the system by its surroundings:

$$\Delta U_{\text{system}} = Q - W$$

For processes that include the transfer of matter, a further statement is needed. When two initially isolated systems are combined into a new system, then the total internal energy of the new system, U_{system} , will be equal to the sum of the internal energies of the two initial systems, U_1 and U_2 :

$$U_{\text{system}} = U_1 + U_2$$

The First Law encompasses several principles:

- Conservation of energy, which says that energy can be neither created nor destroyed, but can only change form. A particular consequence of this is that the total energy of an isolated system does not change.
- The concept of internal energy and its relationship to temperature. If a system has a definite temperature, then its total energy has three distinguishable components, termed kinetic energy (energy due to the motion of the system as a whole), potential energy (energy resulting from an externally imposed force field), and internal energy. The establishment of the concept of internal energy distinguishes the first law of thermodynamics from the more general law of conservation of energy.

$$E_{\text{total}} = KE_{\text{system}} + PE_{\text{system}} + U_{\text{system}}$$

- Work is a process of transferring energy to or from a system in ways that can be described by macroscopic mechanical forces acting between the system and its surroundings. The work done by the system can come from its overall kinetic energy, from its overall potential energy, or from its internal energy.

For example, when a machine (not a part of the system) lifts a system upwards, some energy is transferred from the machine to the system. The system's energy increases as work is done on the system and in this particular case, the energy increase of the system is manifested as an increase in the system's gravitational potential energy. Work added to the system increases the potential energy of the system.

- When matter is transferred into a system, the internal energy and potential energy associated with it are transferred into the new combined system.

$$(u \Delta M)_{\text{in}} = \Delta U_{\text{system}}$$

where u denotes the internal energy per unit mass of the transferred matter, as measured while in the surroundings; and ΔM denotes the amount of transferred mass.

- The flow of heat is a form of energy transfer. Heat transfer is the natural process of moving energy to or from a system, other than by work or the transfer of matter. In a diathermal system, the internal energy can only be changed by the transfer of energy as heat:

$$\Delta U_{\text{system}} = Q$$

Combining these principles leads to one traditional statement of the first law of thermodynamics: It is not possible to construct a machine which will perpetually output work without an equal amount of energy input to that machine.

Or more briefly: A perpetual motion machine of the first kind is impossible.

Second law

The second law of thermodynamics indicates the irreversibility of natural processes, and in many cases, the tendency of natural processes to lead towards spatial homogeneity of matter and energy, especially of temperature.

It can be formulated in a variety of interesting and important ways. One of the simplest is the statements: **That heat does not spontaneously pass from a colder to a hotter body.**

- **It is not possible to build a perpetual motion machine without supplying it with external energy.**
- **There is no spontaneous change of state that can transfer heat from a cold body to a hot body.**
- **Any process that occurs on its own is irreversible.**
- **All mixing processes are irreversible.**

It implies the existence of a quantity called the entropy of a thermodynamic system. In terms of this quantity it implies that:

When two initially isolated systems in separate but nearby regions of space, each in thermodynamic equilibrium with itself but not necessarily with each other, are then allowed to interact, they will eventually reach a mutual thermodynamic equilibrium. The sum of the entropies of the initially isolated systems is less than or equal to the total entropy of the final combination. Equality occurs just when the two original systems have all their respective intensive variables (temperature, pressure) equal; then the final system also has the same values.

The second law is applicable to a wide variety of processes, both reversible and irreversible. According to the second law, in a reversible heat transfer, an

element of heat transferred, δQ , is the product of the temperature (T), both of the system and of the sources or destination of the heat, with the increment (dS) of the system's conjugate variable, its entropy (S):

$$\delta Q = TdS$$

While reversible processes are a useful and convenient theoretical limiting case, all natural processes are irreversible. A prime example of this irreversibility is the transfer of heat by conduction or radiation. It was known long before the discovery of the notion of entropy that when two bodies, initially of different temperatures, come into direct thermal connection, then heat immediately and spontaneously flows from the hotter body to the colder one.

Entropy may also be viewed as a physical measure concerning the microscopic details of the motion and configuration of a system, when only the macroscopic states are known. Such details are often referred to as *disorder* on a microscopic or molecular scale, and less often as *dispersal of energy*. For two given macroscopically specified states of a system, there is a mathematically defined quantity called the 'difference of information entropy between them'. This defines how much additional microscopic physical information is needed to specify one of the macroscopically specified states, given the macroscopic specification of the other – often a conveniently chosen reference state which may be presupposed to exist rather than explicitly stated. A final condition of a natural process always contains microscopically specifiable effects which are not fully and exactly predictable from the macroscopic specification of the initial condition of the process. This is why entropy increases in natural processes – the increase tells how much extra microscopic information is needed to distinguish the

initial macroscopically specified state from the final macroscopically specified state. Equivalently, in a thermodynamic process, energy spreads.

Third law

The third law of thermodynamics can be stated as:

A system's entropy approaches a constant value as its temperature approaches absolute zero.

And that means:

“ The temperature cannot be reached absolute zero ”

At absolute zero temperature, the system is in the state with the minimum thermal energy, the ground state. The constant value (not necessarily zero) of entropy at this point is called the residual entropy of the system. With the exception of non-crystalline solids (e.g. glass) the residual entropy of a system is typically close to zero. However, it reaches zero only when the system has a unique ground state (i.e., the state with the minimum thermal energy has only one configuration, or microstate). Microstates are used here to describe the probability of a system being in a specific state, as each microstate is assumed to have the same probability of occurring, so macroscopic states with fewer microstates are less probable. In general, entropy is related to the number of possible microstates according to the Boltzmann principle

$$S = K_B \ln \Omega$$

where S is the entropy of the system, k_B is the Boltzmann constant, and Ω the number of microstates. At absolute zero there is only 1 microstate possible ($\Omega = 1$ as all the atoms are identical for a pure substance, and as a result all orders are identical as there is only one combination)

and $\ln(1) = 0$

conversion of units used in Heat laws

Generally, heat is anything that provides warmth but scientifically, heat is the flow of energy from a warmer object to a cooler object in comparison till both the objects attain equilibrium. Every matter on earth has some amount of heat energy stored in it. Heat energy flows due to the difference in temperature of the two bodies. In this lecture, students will learn about the units and conversions of heat energy but first let's look at a few terms, definitions and concepts.

Heat: Scientifically, heat is defined as the energy that is spontaneously transferred from one object to another due to differences in temperatures. Heat transfer occurs until the bodies attain equilibrium.

Temperature: Temperature is defined as the kinetic energy of molecules of a body.

Internal Energy: The total energy of all the molecules of a body is the internal energy within the object.

Specific Heat: Specific heat, also known as heat capacity, is the amount of energy required to produce a unit change in its temperature.

Difference between Temperature, Heat, and Internal Energy

- **Temperature** is the kinetic energy of the molecules of a body. The average kinetic energy of individual molecules is termed temperature.
- The total energy of all the molecules is the internal energy within the object. Internal energy is an extensive property.
- Heat is defined as the energy that is spontaneously transferred from one body to another due to its temperature difference.

For example, if a 5 kg of steel, at 100°C, is placed in contact with a 500 kg of steel at 20°C, heat flows from the cube at 300°C to the cube at 20°C, even though, the internal energy of the 20°C cubes is much greater because there is so much more of it. Mathematically heat can be expressed as:

$$C = \frac{Q}{m \times \Delta T}$$

Where m = mass of the body,

C = specific heat,

ΔT = temperature difference. $m \times \Delta T$

Q = heat

SI Unit of Heat: As all the energy is represented in Joules (J), therefore, heat is also represented in Joules. Hence, the SI unit of heat is Joules. Joules can be defined as the amount of energy required to raise the temperature of a given mass by one degree. To increase the temperature of one unit weight of water by one degree, we require 4.184 joules of heat.

Other Heat Units:

Other heat units are:

BTU

Calorie

Joules

BTU:

BTU is a British thermal unit. It is the amount of energy required to raise the temperature of one pound of water by 10 F at sea level.

Conversion:

1 BTU = 1055.06 J = 2.931×10^{-4} kWh = 0.252 kcal = 778.16 ft lbf = 1.055 $\times 10^{10}$ ergs = 252 cal = 0.293 watt-hours

Calorie:

The amount of energy required to raise the temperature of one gram of water by 1° C.

Conversion: 1 kcal = 4186.8 J = 426.9 kp m = 1.163×10^{-3} kWh = 3.088 ft lbf = 3.9683 BTU = 1000 cal

Joule:

Joule is the SI unit of heat.

Conversion:

1 J = 0.1020 kpm = 2.778×10^{-7} kcal = 0.7376 ft lb = 1 kg m² / s² = 1 watt second = 1 Nm = 9.478×10^{-4} BTU

Conversion Table:

Units of Heat		
Calorie	1 cal	4184 J
Joules	1 J	0.000239006 kcal / 0.000947817 Btu
BTU	1 Btu	1055.06 J

Temperature Conversion :

Celsius to Kelvin

$$K = C + 273$$

For Example: $100^{\circ}\text{C} = 100 + 273 = 373 \text{ K}$

Kelvin to Celsius

$$C = K - 273$$

For Example: $273 \text{ K} = 273 - 273 = 0^{\circ}\text{C}$

Celsius To Fahrenheit

$$^{\circ}\text{F} = \frac{9}{5} (^{\circ}\text{C}) + 32$$

Kelvin to Fahrenheit

$$^{\circ}\text{F} = \frac{9}{5} (K - 273) + 32$$

Fahrenheit to Celsius

$$^{\circ}\text{C} = \frac{5}{9} (^{\circ}\text{F} - 32)$$

Fahrenheit to Kelvin

$$K = \frac{5}{9} (^{\circ}\text{F} - 32) + 273$$

Example 1:

An electric kettle contains 1.5 kg of water. The specific heat capacity of water is $4180 \text{ J kg}^{-1} \text{ K}^{-1}$.

Calculate the amount of energy required to raise the temperature of the water from 15°C to 100°C .

Solution:**Given:**

Specific heat (C) = $4180 \text{ J kg}^{-1} \text{ K}^{-1}$

$T_1 = 15^\circ\text{C} = 15 + 273 = 288 \text{ K}$

$T_2 = 100^\circ\text{C} = 100 + 273 = 373 \text{ K}$

$m = 1.5 \text{ kg}$

$$Q = m \times \Delta T \times C$$

$$Q = 1.5 \times 4180 \times (373 - 288) = 533 \text{ kJ}$$

Example 2:

Calculate the energy needed to raise the temperature of the water from 20°C to 90°C ?