

The first law of thermodynamics

The first law of thermodynamics may be stated in several ways:

1. The increase in internal energy of a closed system is equal to the heat supplied to the system minus work done by it.

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

2. For a thermodynamic cycle of a closed system, which returns to its original state, the heat Q_{in} supplied to a closed system in one stage of the cycle, minus that Q_{out} removed from it in another stage of the cycle, equals the net work done by the system.

$$\Delta U_{system} (full\ cycle) = 0$$

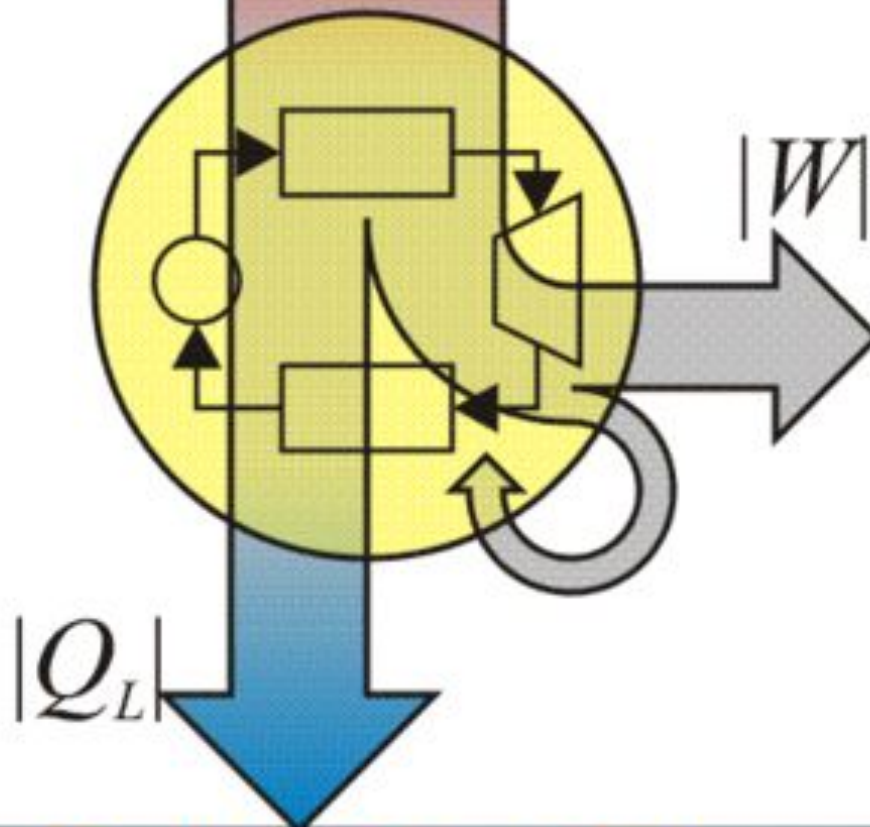
and, consequently $Q = Q_{in} - Q_{out} = W$

3. The increase in internal energy of a closed adiabatic system can only be the result of the net work performed by the system, because $Q = 0$.

$$\Delta U_{system} = U_{final} - U_{initial} = -W$$

Hot source (T_H)

$|Q_H|$



Cold sink (T_L)

Principles of the first law of thermodynamics

1. The law of conservation of energy.

This states that energy can be neither created nor destroyed. However, energy can change forms, and energy can flow from one place to another. The total energy of an isolated system does not change.

2. 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.

- a) If the system is in motion as a whole, it has kinetic energy.
- b) If the system as a whole is in an externally imposed force field (e.g. gravity), it has potential energy relative to some reference point.
- c) Finally, it has internal energy, which is a fundamental quantity for thermodynamics. Beyond the conceptual frame of macroscopic thermodynamics, it can be explained as the sum of the disorganized kinetic energy of microscopic motions of its constituent atoms, and of the potential energy of interactions between them. Other things being equal, the kinetic energy of microscopic motions of the constituent atoms increases as the system's temperature increases. The establishment of the concept of internal energy is the characteristic distinguishing feature of the first law of thermodynamics.

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

3. The flow of heat is a form of energy transfer.

Heating is a natural process of moving energy to or from a system other than by work or the transfer of matter. The heat passes only from a hotter to a colder system.

If the system has rigid walls impermeable to matter, and no external long-range force field affects it, and consequently energy cannot be transferred as work into or out from the system then:

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

where Q denotes the amount of energy transferred into the system as heat.

4. Work is a process of transferring energy to or from a system. Unless otherwise stated, it is customary to treat work as supplied without dissipation in the surroundings. Within the system, in a natural process, some of the transferred work is dissipated.

For example, when a machine lifts a system upwards, some energy is transferred from the machine to the system. The system acquires its energy in the form of gravitational potential energy in this example.

$$-W = \Delta PE_{system}$$

Or in general it can be partitioned to kinetic, potential or internal energy

$$-W = \Delta KE_{system} + \Delta PE_{system} + \Delta U_{system}$$

5. When matter is transferred, its associated internal energy is transferred with it.

$$(u_{external} \Delta M)_{in} = \Delta U_{system}$$

where $u_{external}$ denotes the internal energy per unit mass of the transferred matter, measured when it is still in the surroundings, before transfer; and ΔM denotes the transferred mass.

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 is impossible.

An **isothermal process** is a change of a *system*, in which the temperture remains constant: $\Delta T = 0$. This typically occurs when a system is in contact with an outside thermal reservoir (heat bath = 0). This typically occurs when a system is in contact with an outside thermal reservoir (heat bath), and the change occurs slowly enough to allow the system to continually adjust to the temperature of the reservoir through heat exchange.

In contrast, an **adiabatic process** is where a system exchanges no heat with its surroundings ($Q = 0$). In other words, in an isothermal process, the value $\Delta T = 0$ and therefore $\Delta U = 0$ (only for an ideal gas) but $Q \neq 0$, while in an adiabatic process, $\Delta T \neq 0$ but $Q = 0$.

An ***isobaric process*** is a thermodynamic process in which the pressure stays constant: $\Delta P = 0$. The term derives from the Greek *iso-* (equal) and *baros* (weight). The heat transferred to the system does work, but also changes the internal energy of the system:

An **isochoric process**, also called a **constant-volume process**, is a thermodynamic process, is a thermodynamic process during which the volume, is a thermodynamic process during which the volume of the closed system undergoing such a process remains constant. An isochoric process is exemplified by the heating or the cooling of the contents of a sealed, inelastic container.

A **polytropic process** is a thermodynamic process that obeys the relation:

$$pv^n = C$$

where p is the pressure, v is specific volume, n , the **polytropic index**, is any real number, and C is a constant. The polytropic process equation is particularly useful for characterizing expansion and compression processes which include heat transfer. This equation can accurately characterize a very wide range of thermodynamic processes, that range from $n=0$ to $n=\infty$ which covers, $n=0$ (isobaric), $n=1$ (isothermal), $n=\gamma$ (adiabatic), $n=\infty$ (isochoric) processes. The one restriction is that the process should display an energy transfer ratio of $K=\delta Q/\delta W=\text{constant}$ during that process.