1. First (1st) Law of Thermodynamics
   1. [History](https://en.wikipedia.org/wiki/First_law_of_thermodynamics#History)

The relationship between heat and work was first suggested in 1798 by Count Rumford (Sir Benjamin Thompson), who when observing the boring of cannons noticed that *heat* (*q*) produced was proportional to the *work* (*w*) performed. This led to the suggestion that heat was an invisible fluid called a *caloric* and that it resided in the constituent particles of the substance. With this temperature was considered to be representation of a determined quantity of caloric gas, where if two bodies of differing temperature were placed against one another would result in the caloric flowing between them. Around 40 years later James Prescott Joule would perform his experiments where work would be performed in a certain quantity of adiabatically contained water and measured the resulting increase in temperature. The resulting proportionality between the work and rise in temperature gave rise to the motion of a *mechanical equivalent of heat* and defined a unit of thermal energy known as a *calorie* (or *15° calorie*). This unit of heat represented a quantity of thermal energy needed to increase the temperature of water from 14.5°C to 15.5°C. Based on this, Joule determined the value of the mechanical equivalent of heat to be 0.241 calories per what we now call a *joule* (J). From this, it can be seen that the transfer of thermal energy (*q*) and the performance of work (*w*) are *processes* that occur on a system and are not intrinsic to it in that when applied to a system they change the properties of a system.

The *first law of thermodynamics* is sometimes considered to be an extension of the *law of conservation of energy*. The law of conservation of energy was discovered in the late seventeenth century for use in mechanical systems but more has been added with introduction of the first law of thermodynamics. Of the four laws, this “first” law of thermodynamics was actually the second to be discovered after the *second law of thermodynamics*. The investigation of the relationship between heat and work first began in the industrial age with the invention of the first engine, whose primary use was to pump water out from the coal mines. Over the course of half a century, the first law was empirically developed where the first statements of the law came in 1850 given by Rudolf Clasius. Rudolf Clasius’ statement of the first law referred to the cyclic thermodynamic process where, **“In all cases in which work is produced by the agency of heat, a quantity of heat is consumed which is proportional to the work done; and conversely, by the expenditure of an equal quantity of work an equal quantity of heat is produced”** (Clasisus, 1850).

* 1. [Summary](https://en.wikipedia.org/wiki/First_law_of_thermodynamics)

The first law introduces the important state variable, *internal energy* (*U*), which is also known as the *thermodynamic potential*. It also argues that energy may be converted from one of its existing forms to another form. Lastly, it introduces the concept that the transfer of thermal energy known as *heat* is a different type of energy that is produced during a process of *work*. Overall, this first law is a modification of the law of conservation of energy to be applicable for thermodynamic systems. The law of conservation of energy states the total energy of an isolated system is constant and that energy cannot be created or destroyed, only transformed from form to another. The first law typically follows the equation and definitions:

Together these variables form to provide an equation where **the change in internal energy of a closed system is equal to the amount of heat supplied by the system minus the work done by the system on its surroundings.** This small equation provides the framework for how heat and work in a system are related while also setting limits such as debasing arguments for perpetual motion machines.

* 1. Basic Mechanics

In the context of basic Newtonian mechanics, *kinetic energy* is conserved in a frictionless system of interacting rigid bodies. A collision between two bodies would result in a transfer of kinetic energy from one body to another and the total kinetic energy of the system would remain unchanged. In the case that the system is under the influence of gravity, the total sum of kinetic energy and potential energy would remain constant. As a result, kinetic energy may be converted to potential energy and potential energy may be converted to kinetic energy, but the total sum of energy in this system must remain constant. The common variant added to this case is when friction occurs in the system; in this case the friction converts the dynamic energy into thermal energy.

* 1. Internal Energy

With the first law of thermodynamics it is important to define a function which depends only on the *internal state* of a body or system. In this case, the internal energy (*U*) is such an applicable function which is related to the system’s capacity to do work. When work is done on an adiabatically contained body of constant potential and kinetic energy, internal energy (*U)* can represent the change in the state of the body. With this, “the work done on (or by) an adiabatically contained body equals the change in the internal energy of the body” (Laughlin & Gaskell, 2017). In other words, the difference between internal energy (*U*) value in its final state and its initial state is equal to the work.

In the description of work, the convention is to assign a **negative sign (-) to work done *on* a body** and a **positive sign (+) to work done *by* a body**. This convention becomes significant since we consider the work done to be *PdV* work. Thus, in this context when the gas expands and does work against an external pressure, the integral of *PdV* denotes a quantity of positive work performed by the system. With this positive quantity of work performed by the system, the internal energy must in turn decrease. The sign convention of work is denoted in this equation:

In describing the transfer of thermal energy, the proper sign convention is to assign a **negative sign (-) to the transfer of thermal energy *out* of a body** and a **positive sign (+) to the transfer of thermal energy *into* a body**. Hence, a negative sign is attributes to an exothermic process while a positive sign attributes to an endothermic process. The sign convention of thermal energy transfer is denoted by this equation:

In both of the previous equations, the prime (i.e. *U’*) indicates the total internal energy of a system while the internal energy symbol (*U*) represents the internal energy *per* mole.

Lastly, we consider the change in the internal energy of a body that simultaneously performs work and absorbs thermal energy. In this case we consider a body that is initially in state *A*, that performs work (*w*) and absorbs energy (*q*), through a thermal gradient and as a result moves to state *B*. The absorption of energy (*q*) *increases* the internal energy of the body by the absorbed amount of energy (*q*). The performance of work (*w*) by the body *decreases* its internal energy by the amount of work (*w*) done by the body. This equates to a total change in the internal energy of the body (Δ*U’*) that is represented by:

This equation is a summary of the first law of thermodynamics for a system of fixed composition and also shows that while internal energy of the system changed by an amount (*q*-*w*), the energy of the surroundings has changed by –(*q*-*w*), leaving the total energy of the universe unchanged.

1. Relevant Examples