

Energy Balance and 1st Law of Thermodynamics for Closed Systems

Energy balance for any system undergoing any kind of process was expressed as:

Net energy transfer by heat, work, and mass = Change in internal, kinetic, potential, etc., energies

$$\left(\begin{array}{c} \text{Total energy} \\ \text{entering the system} \end{array} \right) - \left(\begin{array}{c} \text{Total energy} \\ \text{leaving the system} \end{array} \right) = \left(\begin{array}{c} \text{Change in the total} \\ \text{energy of the system} \end{array} \right)$$

$$E_{in} - E_{out} = \Delta E_{system} \quad (kJ)$$

$$\text{Rate form} \quad \dot{E}_{in} - \dot{E}_{out} = dE_{system} / dt \quad (kJ/s), (kW)$$

$$\text{Per unit mass} \quad e_{in} - e_{out} = \Delta e_{system} \quad (kJ/kg)$$

$$\text{As } e = \frac{E}{m} \quad (kJ/kg)$$

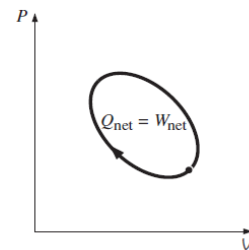
$$\text{Differential form} = \begin{cases} \delta E_{in} - \delta E_{out} = dE_{system} \\ \delta e_{in} - \delta e_{out} = de_{system} \end{cases}$$

The symbol δ means (path function and inexact differentials) such as heat, work, and mass transfer, while the symbol d is used for properties that are (point functions and have exact differentials) such as internal energy, enthalpy, etc.

For a closed system undergoing a cycle, the initial and final states are identical, and thus $\Delta E_{system} = E_2 - E_1 = 0$ and $E_{in} - E_{out} = 0$, $E_{in} = E_{out}$

A closed system does not involve any mass flow across its boundaries; the energy balance for a cycle can be expressed in terms of heat and work interactions as:

$$W_{net,out} = Q_{net,in} \text{ or } W_{net,in} = Q_{net,out} \text{ for a cycle}$$



That is, the net work output during a cycle is equal to net heat input

$$\Delta E = 0, \text{ thus } Q = W$$

The first law of thermodynamics, also known as the conservation of energy principle states that:

“During an interaction between a system and its surroundings, the amount of energy gained by the system must be exactly equal to the amount of energy lost by the surroundings”

The First Law of Thermodynamics (energy balance relation) in that case for a closed system becomes:

$$Q_{net,in} - W_{net,out} = \Delta E_{system} \quad \text{or} \quad Q - W = \Delta E$$

Various forms of the first-law relation for closed systems are given as:

General	$Q - W = \Delta E$
Stationary systems	$Q - W = \Delta U$
Per unit mass	$q - w = \Delta e$
Differential form	$\delta q - \delta w = de$

The change in the total energy of a system during a process is the sum of the changes in its internal (U), kinetic (KE), and potential (PE) energies and can be expressed as:

$$\Delta E = \Delta U + \Delta KE + \Delta PE$$

$$\Delta e = \Delta u + \Delta ke + \Delta pe$$

$$\Delta U = m(u_2 - u_1)$$

$$\Delta KE = \frac{1}{2} m (V_{e2}^2 - V_{e1}^2)$$

$$\Delta PE = mg(z_2 - z_1)$$

When the initial (1) and final (2) states are specified, the values of the specific internal energies u_1 and u_2 can be determined directly from the property tables (vapors, liquid and mixture) or thermodynamic property relations (ideal gas).

For closed and stationary systems, no change in kinetic and potential energies, thus:

$$\Delta KE = 0, \quad \Delta PE = 0$$

$$\Delta E = \Delta U$$

$$Q - W = \Delta U \quad (\text{kJ})$$

$$q - w = \Delta u \quad (\text{kJ/kg})$$

In words:

Net heat supplied + Net work input = Gain in internal energy

Specific Heats

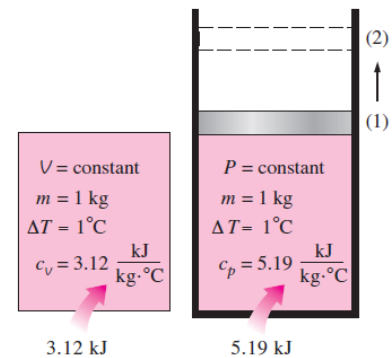
The specific heat is defined as the energy required to raise the temperature of a unit mass of a substance by one degree.

In thermodynamics, we are interested in two kinds of specific heats: specific heat at constant volume C_v and specific heat at constant pressure C_p .

$C_p > C_v$ always because at constant pressure the system is allowed to expand and the energy for this expansion work must also be supplied to the system

$C_v = \left(\frac{\partial u}{\partial T} \right)_v$ = The change in internal energy with temperature at constant volume (kJ/kg.°C, kJ/kg.K)

$C_p = \left(\frac{\partial h}{\partial T} \right)_p$ = The change in enthalpy with temperature at constant pressure (kJ/kg.°C, kJ/kg.K)



Both the specific heat depend on the properties, therefore they must be properties themselves.

Internal Energy, Enthalpy, and Specific Heats of Ideal Gases

Joule's law states that the internal energy of a perfect gas is a function of the absolute temperature only.

$$u = u(T)$$

Using the definition of enthalpy:

$$\left. \begin{aligned} h &= u + Pv \\ Pv &= RT \end{aligned} \right\} h = u + RT$$

Since R is constant and $u = u(T)$, therefore, the enthalpy is a function of temperature only too.

$$h = h(T)$$

Since u and h depend only on temperature for an ideal gas, the specific heats C_v and C_p also depend, at most, on temperature only.

$$C_v = C_v(T)$$

$$C_p = C_p(T)$$

Also,

$$du = C_v(T) dT$$

$$dh = C_p(T) dT$$

Integrating above equations, we get:

$$\Delta u = u_2 - u_1 = \int_1^2 C_v(T) dT$$

$$\Delta h = h_2 - h_1 = \int_1^2 C_p(T) dT$$

By using constant average specific heats, the above equations become:

$$\Delta u = u_2 - u_1 = C_v(T_2 - T_1)$$

$$\Delta h = h_2 - h_1 = C_p(T_2 - T_1)$$

For ideal gases,

$$h = u + RT$$

by differentiating:

$$dh = du + R dT$$

$$C_p dT = C_v dT + R dT$$

$$C_p = C_v + R, \quad C_p > C_v$$

The ratio of specific heat at constant pressure to specific heat at constant volume is called specific heat ratio or adiabatic index k

$$k = \frac{C_p}{C_v}, \quad k > 1$$

Some useful relations can be derived as:

$$C_v = \frac{R}{k-1}$$

$$C_p = \frac{k R}{k-1}$$