

Course Title:

Fundamental of Thermodynamics and Heat Transfer

Chapter 2:

Energy and Energy Transfer

Lecture 1:

Energy and Its Types

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Learning Objective of Lecture:

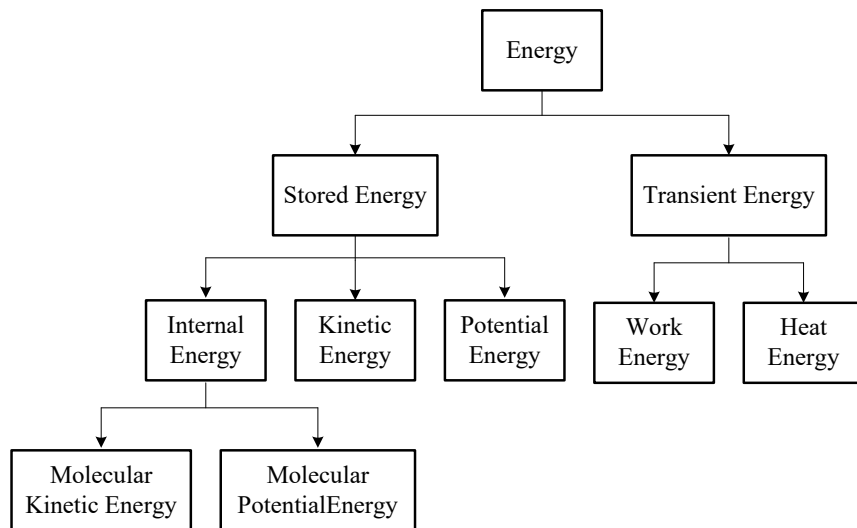
To impart a great deal of knowledge to undergraduate students on the topics:

- ✓ Energy and its Types
- ✓ Stored Energy: Internal Energy (Molecular Kinetic Energy and Molecular Potential Energy), Kinetic Energy and Potential Energy, Total Energy
- ✓ Transient Energy: Work Energy, Heat Energy
- ✓ Energy Transfer: Heat Transfer, Work Transfer

2.0. Energy and Energy Transfer

2.1. Energy and its Types

In mechanics, energy of a system is defined as its capacity for doing work. In other words, energy of a system is the capacity to exert a force through a distance. But in thermodynamics, such a meaning of energy is not adequate. In thermodynamics, energy of a system can be defined as a system quantity that describes a thermodynamic state of the system and its value is changed by transfer to and from the system. Energy has many forms as shown in a chart given below and it can be used in any of these forms.



Energy may be classified into two general types in thermodynamics: *stored energy* in the system and *transient energy* that is energy in transition from one system to another system or surroundings.

2.1.1. Stored Energy and Transient Energy

2.1.1.1. Stored Energy

Stored energy is defined as the energy possessed by a system within its boundaries. This energy remains within the system as an inherent property of

the system. It is a thermodynamic property as it depends on the state of the system. Examples of stored energy are internal energy, kinetic energy and potential energy which are described below.

2.1.1.1.1. Internal Energy

A system has external mechanical energy because of its velocity (kinetic energy) and its position above the reference plane (potential energy). A system has also internal energy resulting from the motion and arrangement of the molecules that constitute the system.

Internal energy of a system is defined as the energy possessed by a system due to its molecular motion (velocity) and arrangement (position) of the molecules. It is therefore the sum of all microscopic forms of energy of a system. Thus, it is the energy needed to create the system and has two major components namely internal (molecular) kinetic energy and internal (molecular) potential energy which are described below.

The internal energy of a system can be changed by heating (or cooling) the system or by doing work on it (or by it). Thermodynamics is chiefly concerned with change in internal energy and not its absolute value. As the molecular activities increase with temperature, internal energy of the system increases with increase in temperature. The change in internal energy denoted by dU depends upon the change in temperature of the system dT as

$$dU = mc_v dT \quad \dots\dots(2.1)$$

where m is the mass of the system and c_v is the specific heat capacity at constant volume.

Internal energy has single value for each equilibrium state and is independent of path. Therefore, it is extensive thermodynamic property.

Internal (Molecular) Kinetic Energy

The *internal or molecular kinetic energy* is the microscopic form of energy due to the motion or velocity of the molecules of the system and it has different forms due to different types of motions of the molecules e.g. translational kinetic energy, rotational kinetic energy, and vibrational kinetic energy of the molecules [1].

When energy is supplied to a system, the motion or velocity of the molecules increases. This implies the internal kinetic energy of the system increases thereby increasing the temperature of the system. Conversely, if the internal kinetic energy decreases by the loss of energy from the system, the motion of the molecules and then temperature of the system will decrease. According to kinetic theory, if the loss of energy from a system continues until the internal kinetic energy reduced to absolute zero (0 Kelvin), then the molecules of the system would become completely motionless.

Internal (Molecular) Potential Energy

Internal or molecular potential energy is defined as the microscopic form of energy due to the molecular separation or configuration. In other words, it is the energy that molecules possess as a result of their position in relative to one another. The greater the degree of molecular separation, the greater is the internal potential energy [2].

When a substance expands or changes its physical state with addition of energy, a rearrangement of molecules takes place that increases the distance between them. In the substance molecules are attracted to one another by forces that tend to pull them together. So, an amount of energy is required to supply to the substance in order to further separate the molecules against this attraction forces. This energy is stored in the substance as an increase in the internal potential energy which is accounted for by the increase in the mean distance between the molecules [2]. Since the mean distance between the molecules of a gas is greater than that of a liquid or a solid, the internal potential energy of a gas is more than that of a liquid or a solid.

Thus, the internal or molecular potential energy is associated with the static constituents of matter, static electric energy of atoms within molecules and the static energy of chemical bonds. Latent energy (energy associated with the phase of a system), chemical energy (energy associated with the atomic bonds in a molecule) and nuclear energy (energy associated with the strong bonds within the nucleus of the atom itself) are also the examples of the molecular potential energy [1].

2.1.1.1.2. Kinetic Energy

Kinetic energy of a system is defined as the energy possessed by a system for doing work due to its mass and velocity of motion. It can also be defined as the work needed to accelerate a system of a given mass from its rest position to a stated velocity. The kinetic energy of a non-rotating system of mass m travelling at a velocity \bar{V} is given by

$$KE = \frac{1}{2} m \bar{V}^2 \quad \dots\dots\dots(2.2)$$

It is a macroscopic form of energy of a system which can be perceived by our sense.

2.1.1.1.3. Potential Energy

Potential energy of a system is defined as the energy possessed by a system for doing work due to its position above the reference plane (ground level). There are various types of potential energy, each associated with a particular type of force. For example, the work of an elastic force is called elastic potential energy; work of the gravitational force is called gravitational potential energy etc. Potential energy of a system for doing work of the gravitational force is

$$PE = mgz \quad \dots\dots\dots(2.3)$$

where m is mass of the system, z is distance through which the system falls and g is acceleration due to gravity. It is also a macroscopic form of energy of a system which can be perceived by our sense.

2.1.1.2. Total Energy

The total energy E of a system is equal to the sum of all forms of energy internal as well as external possessed by a system, i.e. the internal energy, kinetic energy and potential energy of the system. Mathematically,

$$E = U + KE + PE$$
$$\Rightarrow E = U + \frac{1}{2}m\bar{V}^2 + mgz \quad \dots\dots\dots (2.4)$$

For unit mass, the total energy per unit mass of the system, i.e. specific total energy e is given by

$$e = u + \frac{1}{2}\bar{V}^2 + gz \quad \dots\dots\dots(2.5)$$

where u is specific internal energy of the system.

2.1.1.3. Transient Energy

The transient energy is defined as the energy possessed by a system which can cross its boundaries, i.e. it can be transferred from the system to the surroundings and from the surroundings to the system during a thermodynamic process. The transient energy is not a thermodynamic property as it depends upon the path and does not have single value for each equilibrium state. The heat, work and electrical energy are the examples of the transient energy.

2.2. Energy Transfer

Energy transfer is the transfer of the energy from a system to its surroundings and vice versa. In other words, energy transfer is the interactions between a system and its surroundings which bring about changes in the properties of the system, i.e. energy transfer causes the change in the state of the system. There are mainly two ways that these interactions occur:

- Heat transfer
- Work transfer

Thermodynamics mainly studies these energy interactions and the associated property changes of the system.

2.2.1. Heat Transfer

Heat is a disorganized form of energy in transit which can cross the system boundary. It is transferred due to random molecular activities by the mechanisms of conduction, convection and radiation without transfer of mass. Heat transfer occurs always from high temperature system to low temperature system when no device is used. Heat is transient energy which can be identified only during its transfer into or out of a system across the system boundary and therefore not a thermodynamic property of a system. It is a path function denoted by δQ . A precise definition of heat transfer is

Heat transfer is transfer of energy without transfer of mass across the boundary of a system because of a temperature difference between the system and the surroundings [3].

In thermodynamics, heat gain by or supplied to the system is considered as positive and heat lost by or rejected by the system is considered as negative as shown in figure 2.1. Heat transfer is denoted by symbol Q and its unit is Joule (J).

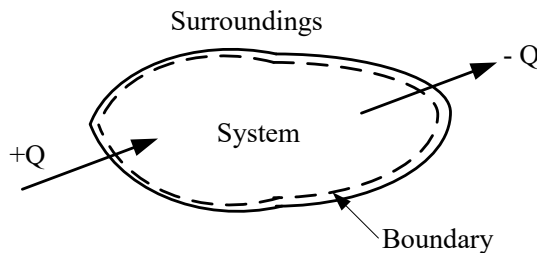


Figure 2.1. Sign convention for heat

2.2.2. Work Transfer

The definition of work in mechanics is the product of a force and distance moved in the direction of the force. Mathematically, the amount of work δW done in moving through a differential distance ds is

$$\delta W = F \cdot ds \quad \dots\dots\dots(2.6)$$

The work done over a finite path between points s_1 and s_2 is

$$W_{1-2} = \int_1^2 F \cdot ds \quad \dots\dots\dots(2.7)$$

This definition of work transfer is not adequate in the study of thermodynamics because during some thermodynamic processes energy is transferred without observable macroscopic displacement. It also fails to provide the distinction between heat and work as we can use heat to cause a force to act through a distance like in steam engine.

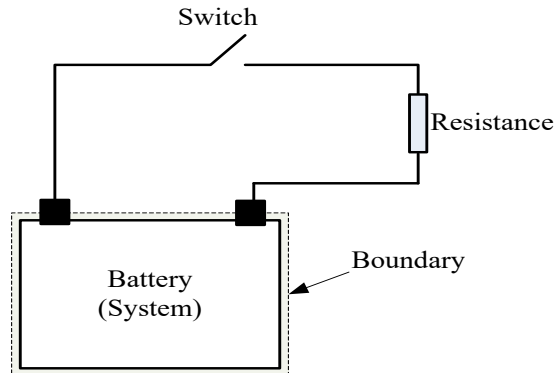


Figure 2.2. Microscopic work interaction due to electric potential difference (without displacement)

For example, let us consider an electric resistance coil connected to an electric storage battery through a circuit as shown in figure 2.2. During any process in this system, energy is transferred from battery to resistance and then from the resistance to the ambient air (surroundings). When switch is turned on, current flows through resistance and its temperature increases and then energy is lost from the resistance coil surface in the form of heat transfer because of temperature difference between the resistance coil surface (system) and the ambient air (surroundings). Again, if we analyze the energy transfer from battery (system) to resistance, it is not dependent upon the temperature and hence it is not the heat transfer. But there has been interaction between the system (battery) and the surroundings (circuit with resistance external to the battery) due to electric potential difference. So, this mode of energy transfer should be work transfer although it does not have any macroscopic displacement. A general definition of work transfer is

Work transfer is transfer of energy without transfer of mass across the boundary of a system because of any property difference other than temperature between the system and the surroundings [3].

This definition is known as thermodynamic definition of work which includes all forms of work and makes a clear distinction between work and heat. Common property differences that can result work transfer are pressure, gravitational potential, electric potential, magnetic potential etc.

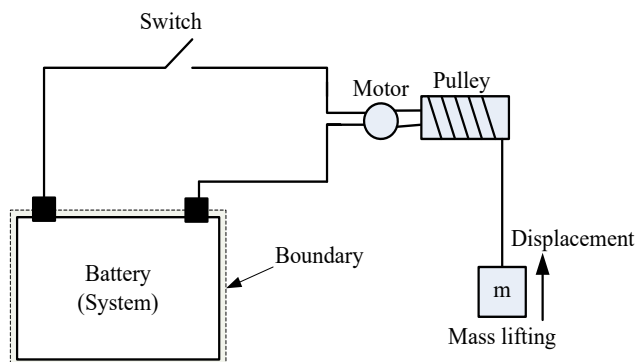


Figure 2.3. Macroscopic work interaction with displacement

Further, the electric resistance can be replaced by a motor – pulley arrangement which can wind a string and thereby lift the suspended mass as shown in figure 2.3. Here, the sole effect external to the system is the lifting of a mass against the gravitational force. Such interaction between system and surroundings is called a work transfer with macroscopic displacement.

Work is a path function like heat and therefore an inexact differential quantity denoted by δW . As the value of work depends not only on the initial and final states of the system but also on the path followed, work is not a thermodynamic property.

In thermodynamics, work done or produced by the system is taken as positive and work done on or consumed by the system is taken as negative as shown in figure 2.4. Work transfer is denoted by symbol W and its unit is also Joule (J).

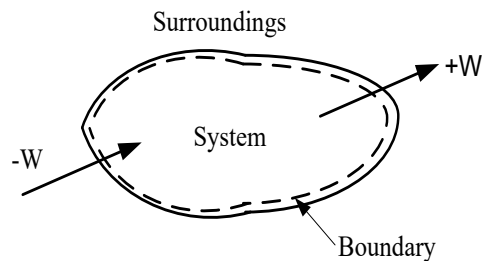


Figure 2.4. Sign Convention for work

Lecture Highlights:

- *Energy*: Energy of a system is defined as its capacity to perform work and it has different forms. The different types of energy in thermodynamics are:
 - (i) *Stored Energy*: Internal energy (Molecular kinetic and potential energy), Kinetic energy and Potential energy
 - (ii) *Transient Energy*: Work energy and Heat energy

- Differences between *stored energy* and *transient energy* are:

Stored Energy	Transient Energy
1. Energy that remains within the system boundary as an inherent property of the system. 2. It can be defined for a state. 3. It is thermodynamic a property. 4. It is called a point or state function. 5. It is not a boundary phenomenon. 6. Its cyclic integral is zero. 7. It has single value for each equilibrium state. 8. For example: internal energy, kinetic energy and potential energy.	1. Energy that can cross the boundary of the system during a thermodynamic process. 2. It can be defined only for a process. 3. It is not a thermodynamic property. 4. It is called a path function. 5. It is a boundary phenomenon. 6. Its cyclic integral is non-zero. 7. It does not have single value for each equilibrium state. 8. For example: work energy, heat energy and electrical energy.

- *Internal energy*: It is defined as the energy of a system due to its molecular activities (motion and position). It is the summation of *molecular potential energy* and *molecular kinetic energy* of the molecules of the system. As the molecular activities increase with temperature, internal energy of the system increases with increase in temperature. The change in internal energy denoted by dU depends upon the change in temperature of the system dT as

$$dU = m c_v dT$$

where m is the mass and c_v is the specific heat capacity at constant volume.

- *Molecular potential energy*: It is the internal or microscopic form of energy that molecules possess as a result of their position in relative to one another. Latent energy, chemical energy and nuclear energy are the examples of the molecular potential energy.
- *Molecular kinetic energy*: It is the internal or microscopic form of energy due to the motion or velocity of the molecules of the system and it has different forms due to different types of motions of the molecules e.g.

translational kinetic energy, rotational kinetic energy, and vibrational kinetic energy of the molecules.

- *Potential Energy*: It is defined as the external or macroscopic form of energy of a system for doing work by virtue of its position or elevation in a gravitational field and is expressed as

$$PE = mgz$$

where m is the mass of the system, g is the acceleration due to gravity and z is the elevation of the system.

- *Kinetic Energy*: It is defined as the external or macroscopic form of energy of a system for doing work by virtue of its velocity and is expressed as

$$KE = \frac{1}{2}m\bar{V}^2$$

where m is the mass of the system and \bar{V} is the velocity of the system.

- *Total energy*: It is the summation of all internal as well as external forms of energy of a system namely internal energy, potential energy and kinetic energy. Mathematically,

$$E = U + PE + KE = U + mgz + \frac{1}{2}m\bar{V}^2$$

- *Energy Transfer*: Energy can be transferred from one system to another or to surroundings in the form of work and heat. Hence, work and heat are energy interactions which bring about changes in the properties of the system. There are mainly two ways that these interactions occur:

- Heat transfer
- Work transfer

- *Heat transfer*: It is transfer of energy without transfer of mass across the boundary of a system because of a temperature difference between the system and the surroundings. In thermodynamics, heat gain by or supplied to the system is considered as positive and heat lost by or rejected by the system is considered as negative.

- *Work Transfer*: Work transfer is transfer of energy without transfer of mass across the boundary of a system because of any property difference other than temperature between the system and the surroundings. In

thermodynamics, work done or produced by the system is taken as positive and work done on or consumed by the system is taken as negative.

References:

- [1] *Thermodynamics: An Engineering Approach*: Cengel Y. A. and Boles M. A., McGraw-Hill Education, New York, 2018.
- [2] *A Textbook of Thermodynamics & Heat Transfer*: Yadav R. D., Pratibha Prakashan, Nepal, 2018.
- [3] *Fundamentals of Thermodynamics & Heat Transfer*: Luintel M.C., Heritage Publishers & Distributors Pvt. Ltd., Kathmandu, Nepal, 2016.