ESO 201 A

Lecture # 4

Forms of Energy

Energy can exist in numerous forms such as thermal, mechanical, kinetic, potential, electric, magnetic, chemical, and nuclear, and their sum constitutes the total energy E of a system. The total energy of a system on a unit mass basis is denoted by 'e' and is expressed as

$$e = \frac{E}{m} \tag{4.1}$$

The unit of e is kJ/kg.

Thermodynamics provides no information about the absolute value of the total energy. It deals only with the change of the total energy. Thus the total energy of a system can be assigned a value of zero ($E=0$) at some convenient reference point. The change in total energy of a system is independent of the reference point selected. The decrease in the potential energy of a falling rock, for example, depends on only the elevation difference and not the reference level selected.

In thermodynamic analysis, it is often helpful to consider the various forms of energy that make up the total energy of a system in two groups: macroscopic and microscopic. The macroscopic forms of energy are those a system possesses as a whole with respect to some outside reference frame, such as kinetic and potential energies. The microscopic forms of energy are those related to the molecular structure of a system and the degree of the molecular activity, and they are independent of outside reference frames. The sum of all the microscopic forms of energy is called the internal energy of a system and is denoted by U.

The macroscopic energy of a system is related to motion and the influence of some external effects such as gravity, magnetism, electricity, and surface tension. The energy that a system possesses as a result of its motion relative to some reference frame is called kinetic energy (KE).

When all parts of a system move with the same velocity, the kinetic energy is expressed as

$$KE = m\frac{V^2}{2} \quad (4.2)$$

The unit of KE is kJ.

or, on a unit mass basis,

$$ke = \frac{V^2}{2} \quad (4.3)$$

The unit of ke is kJ/kg.

The 'V' denotes the velocity of the system relative to some fixed reference frame. The kinetic energy of a rotating solid body is given by $\frac{1}{2}I\omega^2$ where I is the moment of inertia of the body and ω is the angular velocity.

The energy that a system possesses as a result of its elevation in a gravitational field is called potential energy (PE) and is expressed as

$$PE = mgz \quad (kJ) \quad (4.4)$$

or on a unit mass basis,

$$pe = gz \quad (kJ/kg) \quad (4.5)$$

where g is the acceleration due to gravity and z is the elevation of the centre of gravity of a system relative to some arbitrarily selected reference level.

The magnetic, electric and surface tension effects are significant only in some specialized cases and are usually ignored. In the absence of such effects, the total energy of a system consists of the kinetic, potential, and internal energies and is expressed as

$$E = U + KE + PE \quad (4.6)$$

$$= U + m\frac{V^2}{2} + mgz$$

or, on a unit mass basis,

$$e = u + ke + pe$$
$$= u + \frac{V^2}{2} + gz \qquad (4.7)$$

Energy Transfer by Heat

Heat is defined as the energy in transit due to temperature difference in a medium or between media. Heat can not be stored.

As a form of energy, heat has energy units, kJ being the most common one. The amount of heat transferred during the process between two states (states 1 and 2) is denoted by Q_{12} or just Q. Heat transfer per unit mass of a system is denoted by q and is determined from

$$q = \frac{Q}{m} \quad (\mathrm{kJ/kg}) \qquad (4.8)$$

Sometimes it is desirable to know the rate of heat transfer (the amount of heat transferred per unit time) instead of the total heat transferred over some time interval. The heat transfer rate is denoted by $\dot{Q}$, where the overdot stands for the time derivative, or "per unit time". The heat transfer rate $\dot{Q}$ has the unit kJ/s, which is equivalent to kW. When $\dot{Q}$ varies with time, the amount of heat transfer during a process is determined by integrating $\dot{Q}$ over the time interval of the process:

$$Q = \int_{t_1}^{t_2} \dot{Q}\, dt \qquad (4.9)$$

When $\dot{Q}$ remains constant during a process, this relation reduces to

~~4.15~~

$$Q = \dot{Q}\, \Delta t \qquad (4.10)$$

where $\Delta t = t_2 - t_1$ is the time interval during which the process takes place.

Energy Transfer by Work

Work like heat, is an energy interaction between a system and its surroundings. More specifically, work is the energy transfer associated with a force acting through a distance.

Work is also a form of energy transferred like heat, and, therefore, has energy units such as KJ. Like heat, work also cannot be stored.

The work done during a process between states 1 and 2 is denoted by W_{12}, or simply W. The work done per unit mass of a system is denoted by w and is expressed as

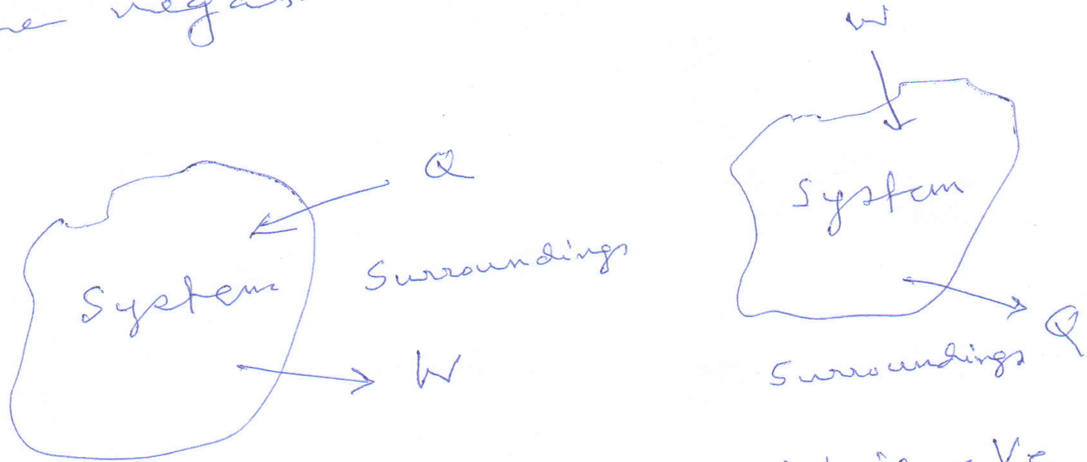
$$w = \frac{W}{m} \quad (KJ/Kg) \qquad (4.11)$$

The work done per unit time is power and is denoted by $\dot{W}$. The unit of power is KJ/s or KW.

The sign convention for heat and work interactions is as follows.

Heat transfer to a system and work done by a system is positive (Fig. 4.1 (a)).

Heat transfer from a system and work done on a system are negative (Fig. 4.1 (b)).



(a) W is +Ve
Q is +Ve

(b) W is −Ve
Q is −Ve

Fig. 4.1 Sign convention for heat and work

Heat and work are energy transfer mechanisms between a system and its surroundings, and there are many similarities between them:

1. Both are recognized at the boundaries of a system as they cross the boundaries. That is, both heat and work are boundary phenomena. They cannot be stored.
2. Systems possess energy, but not heat or work.
3. Both are associated with a process, not a state.
4. Both are path functions (i.e. their magnitudes depend on the path followed during a process as well as the end states).

Path functions have <u>inexact differentials</u> designated by the symbol δ. Therefore, a differential amount of heat or work is represented by δQ or δW, respectively, instead of dQ or dW. Properties, however, are <u>point functions</u> (i.e., they depend on the state only, and not on how a system reaches that state), and they have <u>exact differentials</u> designated by the symbol d. A small change in volume, for example, is represented by $d\forall$ and the total volume change during a process between states 1 and 2 is

$$\int_1^2 d\forall = \forall_2 - \forall_1 = \Delta\forall$$

See Fig. 4.2.

The total work done during the process 1-2, however, is

$$\int_1^2 \delta W = W_{12} \quad (\text{not } \Delta W)$$

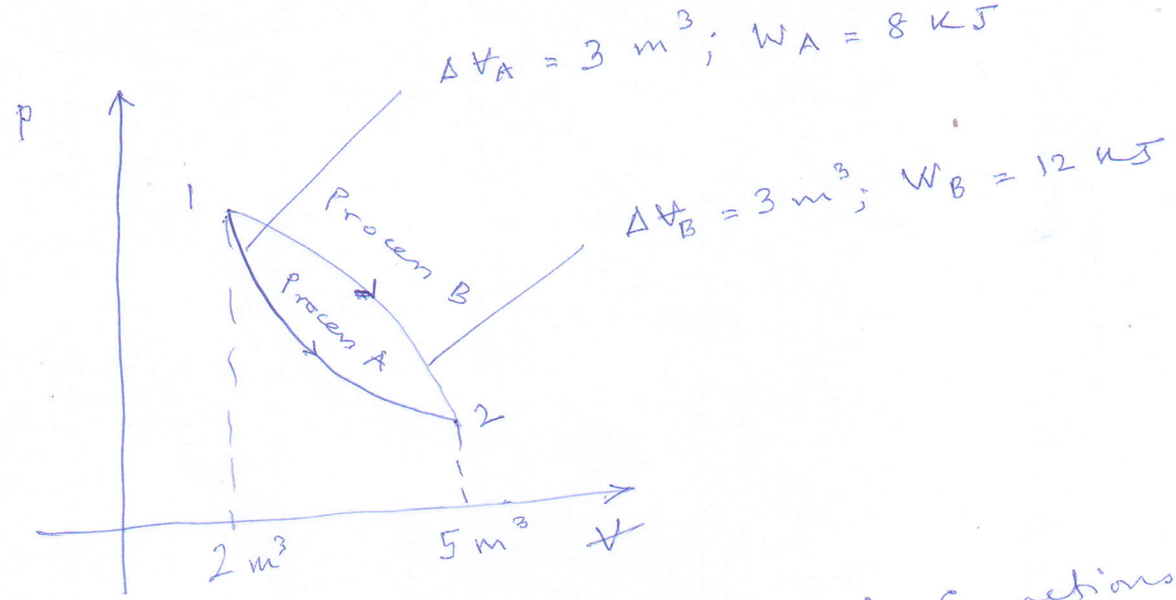


Fig. 4.2 Properties are point functions; but heat and work are path functions (their magnitudes depend on the path followed).

That is, the total work is obtained by following the process path and adding the differential amounts of work (δW) done along the way. The integral of δW is not $W_2 - W_1$, which is meaningless since work is not a property and systems do not possess work at a state.