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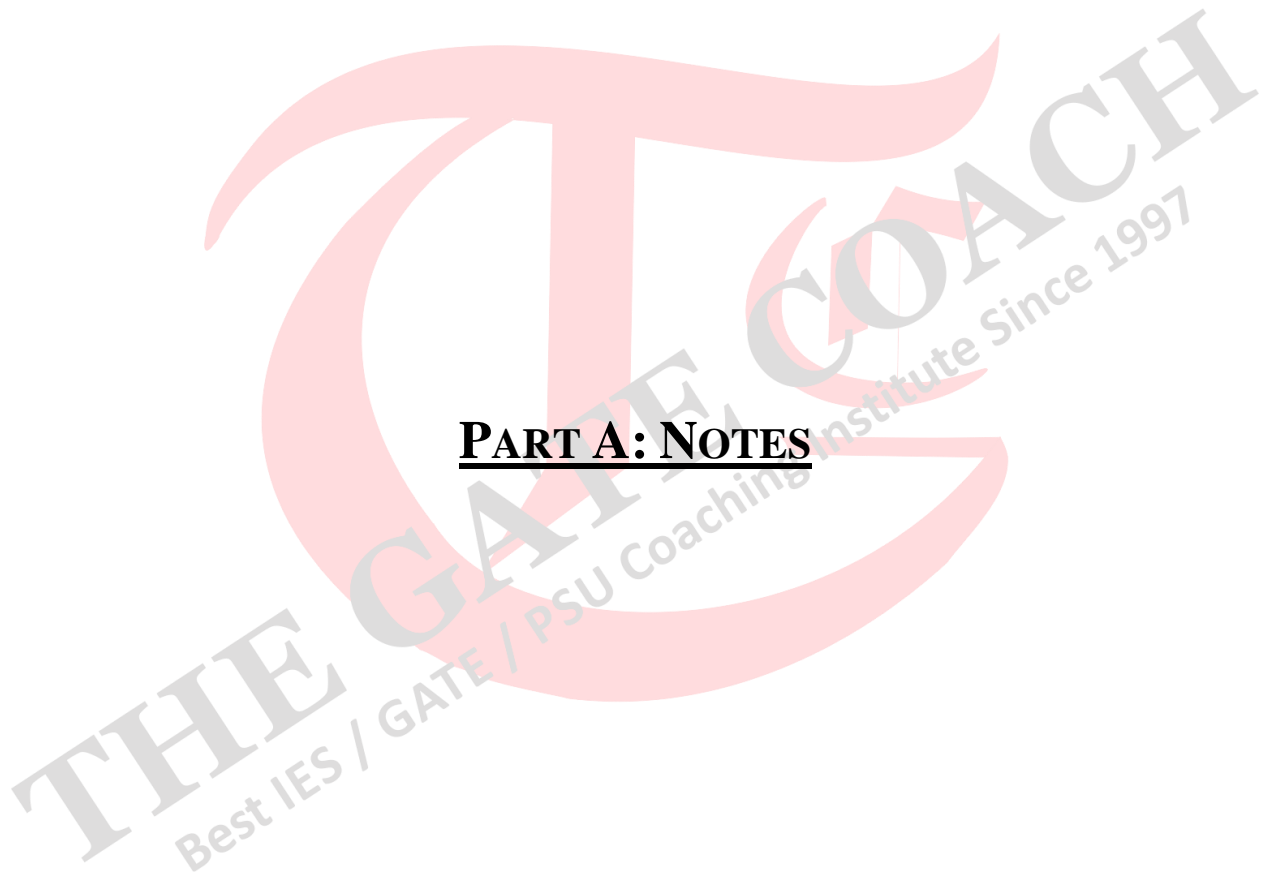
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PART A: NOTES



CHAPTER 1

INTRODUCTION

1.1 INTRODUCTION

Thermodynamics is the branch of science that embodies the principles of energy transformation in macroscopic systems. The general restrictions which experience has shown to apply to all such transformations are known as the laws of thermodynamics. These laws are primitive; they cannot be derived from anything more basic. The first law of thermodynamics states that energy is conserved; that although it can be altered in form and transferred from one place to another, the total quantity remains constant. Thus, the first law of thermodynamics depends on the concept of energy; but, conversely, energy is an essential thermodynamic function because it allows the first law to be formulated. This coupling is characteristic of the primitive concepts of thermodynamics.

- **SYSTEM AND SURROUNDING**

The words system and surroundings are similarly coupled. A system is taken to be any object, any quantity of matter, any region, and so on, selected for study and set apart (mentally) from everything else, which is called the Surroundings. The imaginary envelope which encloses the system and separates it from its surroundings is called the boundary of the system. Attributed to this boundary are special properties which may serve either

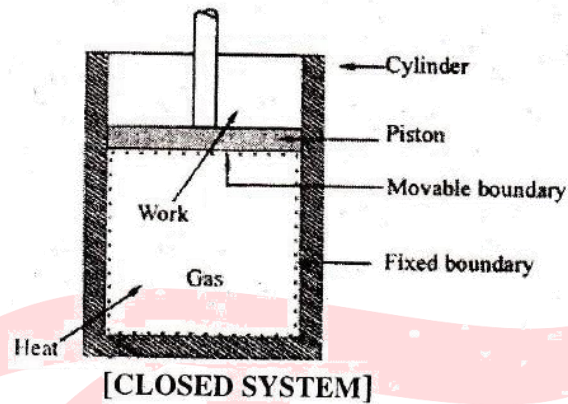
- (1) To isolate the system from its surroundings, or
- (2) To provide for interaction in specific ways between system and surroundings

1.2 CLASSIFICATION OF THERMODYNAMIC SYSTEM

On the basis of mass and energy transfer across / through the system boundaries, a thermodynamic system can be classified as follows:

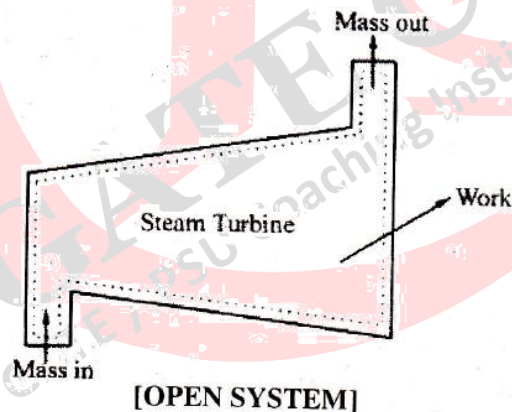
1.2.1 CLOSED SYSTEM (OR NON FLOW SYSTEM)

The system which can exchange energy with surroundings but which cannot transfer matter across the boundaries are known as *closed system*.



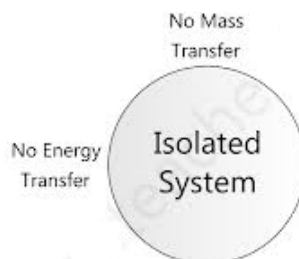
1.2.2 OPEN SYSTEM (OR FLOW SYSTEM)

The system that can exchange both energy and matter with their environment.



1.2.3 ISOLATED SYSTEM

An isolated system exchanges neither matter nor energy with its surroundings.



1.3 BASIC TERMINOLOGIES

There are some basic terms one should know for the good study of thermodynamics:

1.3.1 THERMODYNAMIC EQUILIBRIUM

When a system is in equilibrium with regard to all possible changes in state, the system is in *thermodynamic equilibrium*. For example, if the gas that comprises a system is in thermal equilibrium, the temperature will be the same throughout the entire system.

1.3.2 CONTROL VOLUME

A *control volume* is a fixed region in space chosen for the thermodynamic study of mass and energy balances for flowing systems. The boundary of the control volume may be a real or imaginary envelope. The control surface is the boundary of the control volume.

1.3.3 STEADY STATE

Steady state is that circumstance in which there is no accumulation of mass or energy within the control volume, and the properties at any point within the system are independent of time.

1.3.4 THERMODYNAMIC PROCESS

Whenever one or more of the properties of a system change, a change in the state of the system occurs. The path of the succession of states through which the system passes is called the thermodynamic process. One example of a thermodynamic process is increasing the temperature of a fluid while maintaining a constant pressure. Another example is increasing the pressure of a confined gas while maintaining a constant temperature.

1.3.4.1 CYCLIC PROCESS

When a system in a given initial state goes through a number of different changes in state (going through various processes) and finally returns to its initial values, the system has undergone a cyclic process or cycle. Therefore, at the conclusion of a cycle, all the properties have the same value they had at the beginning. Steam (water) that circulates through a closed cooling loop undergoes a cycle.

1.3.4.2 REVERSIBLE PROCESS

A reversible process for a system is defined as a process that, once having taken place, can be reversed, and in so doing leaves no change in either the system or surroundings. In other words the system and surroundings are returned to their original condition before the process took place. In reality, there are no truly reversible processes; however, for analysis purposes, one uses reversible to make the analysis simpler, and to determine maximum theoretical efficiencies. Therefore, the reversible process is an appropriate starting point on which to base engineering study and calculation. Although the reversible process can be approximated, it can never be matched by real processes.

One way to make real processes approximate reversible process is to carry out the process in a series of small or infinitesimal steps. For example, heat transfer may be considered reversible if it occurs due to a small temperature difference between the system and its surroundings. For example, transferring heat across a temperature difference of 0.00001 °F "appears" to be more reversible than for transferring heat across a temperature difference of 100 °F. Therefore, by cooling or heating the system in a number of infinitesimally small steps, we can approximate a reversible process.

1.3.4.3 IRREVERSIBLE PROCESS

An irreversible process is a process that cannot return both the system and the surroundings to their original conditions. That is, the system and the surroundings would not return to their original conditions if the process was reversed. For example, an automobile engine does not give back the fuel it took to drive up a hill as it coasts back down the hill. There are many factors that make a process irreversible. Four of the most common causes of irreversibility are friction, unrestrained expansion of a fluid, heat transfer through a finite temperature difference, and mixing of two different substances.

These factors are present in real, irreversible processes and prevent these processes from being reversible.

1.3.4.4 ADIABATIC PROCESS

An adiabatic process is one in which there is no heat transfer into or out of the system. The system can be considered to be perfectly insulated.

1.3.4.4 ISENTROPIC PROCESS

An isentropic process is one in which the entropy of the fluid remains constant. This will be true if the process the system goes through is reversible and adiabatic. An isentropic process can also be called a constant entropy process.

1.3.4.5 POLYTROPIC PROCESS

When a gas undergoes a reversible process in which there is heat transfer, the process frequently takes place in such a manner that a plot of the Log P (pressure) vs. Log V (volume) is a straight line. Or stated in equation form $PV^n = \text{constant}$. This type of process is called a polytropic process. An example of a polytropic process is the expansion of the combustion gasses in the cylinder of a water-cooled reciprocating engine.

1.3.4.6 THROTTLING PROCESS

A throttling process is defined as a process in which there is no change in enthalpy from state one to state two, $h_1 = h_2$; no work is done, $W = 0$; and the process is adiabatic, $Q = 0$. To better understand the theory of the ideal throttling process let's compare what we can observe with the above theoretical assumptions.

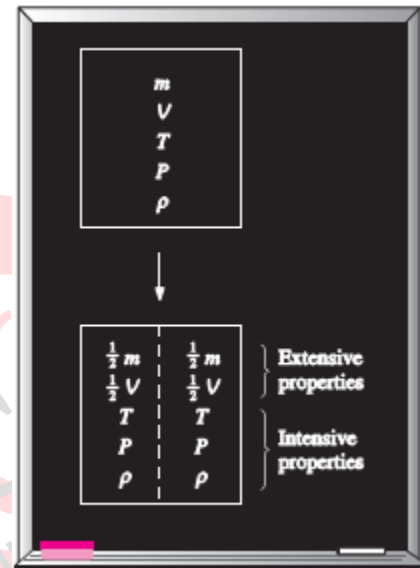
1.4 PROPERTIES OF A SYSTEM

Any characteristic of a system is called a *property*. Some familiar properties are pressure P , temperature T , volume V , and mass m . The list can be extended to include less familiar ones such as viscosity, thermal conductivity, modulus of elasticity, thermal expansion coefficient, electric resistivity, and even velocity and elevation. Properties are considered to be either intensive or extensive.

Intensive properties are those that are independent of the mass of a system, such as temperature, pressure, and density.

Extensive properties are those whose values depend on the size or extent of the system. Total mass, total volume and total momentum are some examples of extensive properties.

An easy way to determine whether a property is intensive or extensive is to divide the system into two equal parts with an imaginary partition, as shown in figure. Each part will have the same value of intensive properties as the original system, but half the value of the extensive properties.



Generally, uppercase letters are used to denote extensive properties (with mass m being a major exception), and lowercase letters are used for intensive properties (with pressure P and temperature T being the obvious exceptions).

Extensive properties per unit mass are called *specific properties*. Some examples of specific properties are specific volume ($v = V / m$) and specific total energy ($e = E / m$).

1.4.1 FUNDAMENTAL PROPERTIES

There are following fundamental properties which are used to describe the state of a system. Such as

1.4.1.1 VOLUME

Space occupied by the system in three dimensions is called *volume* of the system.

Unit: m^3 , litre, c.c.

and $1 \text{ m}^3 = 10^3 \text{ litre} = 10^6 \text{ c.c.}$

1.4.1.2 PRESSURE

Force exerted by the gaseous system perpendicular to the unit surface area.

Mathematically,

$$\text{Pressure} = \frac{\text{Normal Component of Force}}{\text{Area}}$$

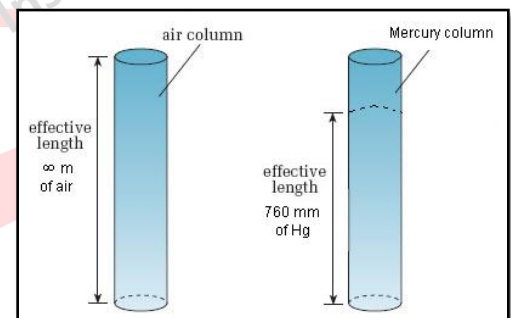
Units: Pa, N / m^2 , atm...

And $1 \text{ atm} = 101.325 \text{ kPa} = 101325 \text{ N} / \text{m}^2$

- **Standard Atmospheric Pressure**

Weight of air column per 1 m^2 area is called standard atmospheric pressure.

Since it is not practically possible to determine the infinite height of air column thus we take a equivalent column of same weight filled with Mercury (Hg), and it comes to be 760 mm height.



Hence we can say that

$$1 \text{ atm} \equiv 760 \text{ mm Hg}$$

Hence,

From Hydrostatic law, we know that

$$P = \rho gh$$

Therefore,

$$1 \text{ atm} \approx$$

$$P = (13595.1) \frac{\text{kg}}{\text{m}^3} \times (9.80665) \frac{\text{m}}{\text{s}^2} \times (0.760) \text{m}$$

$$P = 101325.01 \frac{\text{N}}{\text{m}^2}$$

$$\boxed{P = 101.325 \text{ kPa}}$$

$$\boxed{P = 1.01325 \text{ bar}}$$

$$\text{and } 1 \text{ bar} = 10^5 \text{ Pa}$$

1.4.1.3 FORCE

Force is which may change or tends to change the state of the system.

Force \propto Rate of change of momentum

$$F \propto \frac{d(mv)}{dt}$$

$$F \propto \left(m \frac{dv}{dt} + v \frac{dm}{dt} \right)$$

Since, m is the mass (i.e. matter contained by the body) of the body which is independent of the time and place.

Thus equation becomes:

$$F \propto m \frac{dv}{dt}$$

$$F \propto ma$$

$$\boxed{F = k.ma}$$

Where k is the proportionality constant.

According to Force law:

$$1 \text{ kg}_f = k \times 1 \text{ kg}_m \times 9.81 \frac{\text{m}}{\text{s}^2}$$

$$\Rightarrow k = \frac{1}{9.81} = \frac{1}{g_c} (\text{say})$$

$g_c \rightarrow$ constant value of g everywhere

$$F = \frac{1}{g_c} ma$$

$$\Rightarrow 1\text{kg}_f = \frac{1}{g_c} \times 1\text{kg}_m \times 9.81 \frac{\text{m}}{\text{s}^2}$$

$$\Rightarrow \boxed{g_c = 9.81 \frac{\text{kg}_m \cdot \text{m}}{\text{kg}_f \cdot \text{s}^2}}$$

Example 1.1 A body having mass 70 kg_m weighing 70 kg_f at $g = 9.81 \text{ m / s}^2$, when it placed at a place where the value of g becomes 5 m / s^2 . What will be the weight of the body?

Solution:

From Newton's second law, we know that

$$W = \frac{1}{g_c} mg$$

$$\Rightarrow W = \frac{1}{9.81} \left(\frac{\text{kg}_f \cdot \text{s}^2}{\text{kg}_m \cdot \text{m}} \right) \times 70(\text{kg}) \times 5 \left(\frac{\text{m}}{\text{s}^2} \right)$$

$$\Rightarrow \boxed{W = 36.67788 \text{ kg}_f}$$

- **Relationship between kg_f and N**

$$F = 70 \text{ kg}_m \times 5 \left(\frac{\text{m}}{\text{s}^2} \right) = 350 \frac{\text{kg}_m \cdot \text{m}}{\text{s}^2} = 350 \text{ N}$$

hence, Mathematically

$$F = W$$

$$\Rightarrow 350 \text{ N} = 36.67788 \text{ kg}_f$$

$$\Rightarrow \boxed{1 \text{ kg}_f = 9.81 \text{ N}}$$

1.4.1.4 VELOCITY

Velocity may be defined as the distance travelled by the body per unit time.

$$v = \frac{\text{Distance travelled by the body}}{\text{Time taken by the body to travel that distance}} = \frac{d}{t}$$

1.4.1.5 TEMPERATURE

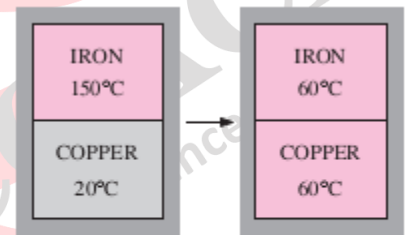
Temperature may be defined as the thermal potential of a system responsible for energy transfer (i.e. heat transfer).

According to science,

Temperature is the measure of average kinetic energy of gases.

- **THERMAL EQUILIBRIUM**

When a body is brought into contact with another body that is at a different temperature, heat is transferred from the body at higher temperature to the one at lower temperature until both bodies attain the same temperature as shown in figure. At that point, the heat transfer stops, and the two bodies are said to have reached thermal equilibrium. The equality of temperature is the only requirement for thermal equilibrium.



Two bodies reaching thermal equilibrium after being brought into contact in an isolated enclosure.

1.4.2 ZEROth LAW OF THERMODYNAMICS

The *zeroth law of thermodynamics* states that if two bodies are in thermal equilibrium with a third body, they are also in thermal equilibrium with each other. It may seem silly that such an obvious fact is called one of the basic laws of thermodynamics. However, it cannot be concluded from the other laws of thermodynamics, and it serves as a basis for the validity of temperature measurement. By replacing the third body with a thermometer, the zeroth law can be restated as two bodies are in thermal equilibrium if both have the same temperature reading even if they are not in contact.

- **TEMPERATURE SCALES**

Temperature scales enable us to use a common basis for temperature measurements. All temperature scales are based on some easily reproducible states

such as the freezing and boiling points of water, which are also called the ice point and the steam point, respectively.

A mixture of ice and water that is in equilibrium with air saturated with vapor at 1 atm pressure is said to be at the *ice point*.

A mixture of liquid water and water vapor (with no air) in equilibrium at 1 atm pressure is said to be at the *steam point*.

The temperature scales used in the SI and in the English system are the Celsius scale and the Fahrenheit scale, respectively.

On the *Celsius scale*, the ice and steam points were originally assigned the values of 0 and 100°C, respectively.

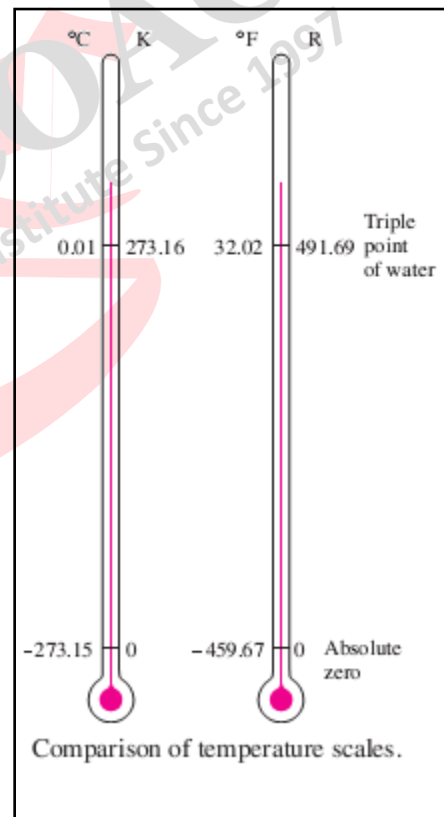
The corresponding values on the *Fahrenheit scale* are 32 and 212°F. These are often referred to as two-point scales since temperature values are assigned at two different points.

In thermodynamics, it is very desirable to have a temperature scale that is independent of the properties of any substance or substances. Such a temperature scale is called a *thermodynamic temperature scale*, which is developed later in conjunction with the second law of thermodynamics.

The thermodynamic temperature scale in the SI is the *Kelvin scale*. The temperature unit on this scale is the kelvin, which is designated by K (not °K). The lowest temperature on the Kelvin scale is absolute zero, or 0 K. Then it follows that only one nonzero reference point needs to be assigned to establish the slope of this linear scale. Using nonconventional refrigeration techniques, scientists have approached absolute zero kelvin (they achieved 0.000000002 K in 1989).

The thermodynamic temperature scale in the English system is the *Rankine scale*, named after William Rankine (1820–1872). The temperature unit on this scale is the rankine, which is designated by R.

The Kelvin scale is related to the Celsius scale by



$$T(K) = T(^{\circ}C) + 273.15$$

The Rankine scale is related to the Fahrenheit scale by

$$T(R) = T(^{\circ}F) + 459.67$$

The temperature scales in the two unit systems are related by

$$T(R) = 1.8 \times T(K)$$

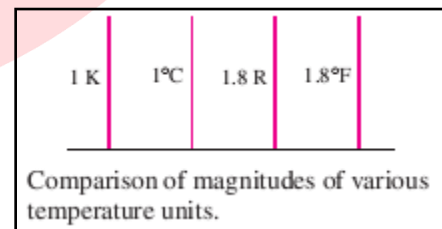
$$T(^{\circ}F) = 1.8 \times T(^{\circ}C) + 32$$

The reference temperature chosen in the original Kelvin scale was 273.15 K (or 0°C), which is the temperature at which water freezes (or ice melts) and water exists as a solid – liquid mixture in equilibrium under standard atmospheric pressure (the ice point).

At the Tenth General Conference on Weights and Measures in 1954, the reference point was changed to a much more precisely reproducible point, the *triple point* of water (the state at which all three phases of water coexist in equilibrium), which is assigned the value 273.16 K.

The Celsius scale was also redefined at this conference in terms of the ideal gas temperature scale and a single fixed point, which is again the triple point of water with an assigned value of 0.01°C. The boiling temperature of water (the steam point) was experimentally determined to be again 100.00°C.

We emphasize that the magnitudes of each division of 1 K and 1°C are identical as shown in figure. Therefore, when we are dealing with temperature differences ΔT , the temperature interval on both scales is the same. Raising the temperature of a substance by 10°C is the same as raising it by 10 K. That is



$$\Delta T(K) = \Delta T(^{\circ}C)$$

$$\Delta T(R) = \Delta T(^{\circ}F)$$

1.5 IDEAL GAS

An ideal gas is one in which the distance between the molecules is so large that the inter molecular forces are negligible and the volume occupied by the molecules is only a negligible fraction of the total volume. It follows from kinetic theory that for such a

fluid the product of pressure and volume varies linearly with temperature. This is stated mathematically by the ideal gas equation $PV = RT$, where R is a constant known as the ideal gas constant and V is the molar volume of the gas.

Mathematically, we write the equation of state, as follows

$$PV = nRT$$

1.6 AVOGADRO'S HYPOTHESIS

It states that,

“All gases contain equal no. of molecules in a given volume at NTP.”

i.e. Avogadro's number , $N_A = 6.023 \times 10^{26}$ no. of molecules
 $N_A = 6.023 \times 10^{23}$ no. of kilo – molecules
in 22.41 m³ Volume

E.g.

$H_2 \rightarrow$ In 22.41 m³ = 6.023×10^{26} no. of molecules of 2 kilogram mass = 1 kgmol

1.7 ENERGY, WORK AND HEAT

Heat and work are the two ways in which energy can be transferred across the boundary of a system. One of the most important discoveries in thermodynamics was that work could be converted into an equivalent amount of heat and that heat could be converted into work.

1.7.1 ENERGY

Energy is defined as the capacity of a system to perform work or produce heat.

1.7.1.1 POTENTIAL ENERGY

Potential energy (PE) is defined as the energy of position. Using English system units, it is defined by Equation

$$PE = \frac{mgz}{g_c}$$

where:

PE = potential energy (N – m)

m = mass (kg_m)

z = height above some reference level (m)

g = acceleration due to gravity (m / s²)

g_c = gravitational constant = 9.81 m – kg_m / kg_f - s²

1.7.1.2 KINETIC ENERGY

Kinetic energy (KE) is the energy of motion. Using English system units, it is defined by Equation

$$KE = \frac{mv^2}{2g_c}$$

where:

KE = kinetic energy (N – m)

v = velocity (m / sec)

1.7.1.3 SPECIFIC INTERNAL ENERGY

The *specific internal energy* (u) of a substance is its internal energy per unit mass. It equals the total internal energy (U) divided by the total mass (m).

i.e.

$$u = \frac{U}{m}$$

where:

u = specific internal energy (J / kg)

U = internal energy (J)

1.7.1.4 SPECIFIC P-V ENERGY

In addition to the internal energy (U), another form of energy exists that is important in understanding energy transfer systems. This form of energy is called $P - V$ energy because it arises from the pressure (P) and the volume (V) of a fluid. It is numerically equal to PV , the product of pressure and volume. Because energy is defined as the capacity of a system to perform work, a system where pressure and volume are permitted to expand performs work on its surroundings. Therefore, a fluid under pressure has the capacity to perform work. In engineering applications, the units of P-V energy, also called flow energy, are the units of pressure times volume (kg_f per square meter times cubic meter), which equals $\text{kg}_f - \text{m}$ ($\text{kg}_f - \text{m}$).

The specific P-V energy of a substance is the P-V energy per unit mass. It equals the total P-V divided by the total mass m , or the product of the pressure P and the specific volume v , and is written as Pv .

$$Pv = \frac{PV}{m}$$

where:

P = pressure (kg_f / m^2)

V = volume (m^3)

v = specific volume (m^3 / kg_m)

1.7.1.5 SPECIFIC ENTHALPY

Specific enthalpy (h) is defined as $h = u + Pv$, where u is the specific internal energy (J / kg) of the system being studied, P is the pressure of the system (kg_f / m^2), and v is the specific volume (m^3 / kg_m) of the system. Enthalpy is usually used in

connection with an "open" system problem in thermodynamics. Enthalpy is a property of a substance, like pressure, temperature, and volume, but it cannot be measured directly. Normally, the enthalpy of a substance is given with respect to some reference value.

For example, the specific enthalpy of water or steam is given using the reference that the specific enthalpy of water is zero at .01°C and normal atmospheric pressure. The fact that the absolute value of specific enthalpy is unknown is not a problem, however, because it is the change in specific enthalpy (Δh) and not the absolute value that is important in practical problems.

1.7.2 WORK

Work is a form of energy, but it is energy in transit. Work is not a property of a system. Work is a process done by or on a system, but a system contains no work. This distinction between the forms of energy that are properties of a system and the forms of energy that are transferred to and from a system is important to the understanding of energy transfer systems.

Work is defined for mechanical systems as the action of a force on an object through a distance. It equals the product of the force (F) times the displacement (d).

$$W = F.d$$

where:

W = work (m – kg_f)

F = force (kg_f)

d = displacement (m)

Work = Unit Force \times Unit Displacement

$$\Rightarrow W = 1 N \times 1 m$$

$$\Rightarrow \boxed{1 J = 1 N - m}$$

1.7.3 HEAT

Heat, like work, is energy in transit. The transfer of energy as heat, however, occurs at the molecular level as a result of a temperature difference. The symbol Q is used to denote heat. In engineering applications, the unit of heat is the calorie (cal).

As with work, the amount of heat transferred depends upon the path and not simply on the initial and final conditions of the system. Also, as with work, it is important to distinguish between heat added to a system from its surroundings and heat removed from a system to its surroundings. A positive value for heat indicates that heat is added to the system by its surroundings. This is in contrast to work that is positive when energy is transferred from the system and negative when transferred to the system. The symbol q is sometimes used to indicate the heat added to or removed from a system per unit mass. It equals the total heat (Q) added or removed divided by the mass (m). The term "specific heat" is not used for q since specific heat is used for another parameter. The quantity represented by q is referred to simply as the heat transferred per unit mass.

$$q = \frac{Q}{m}$$

where:

q = heat transferred per unit mass (cal / kg_m)

Q = heat transferred (cal)

The amount of heat is required to raised the temperature of 1 kg of water (as 15 to 16°C) is the 1 kcal.