

INTRODUCTION

1.1 PRIME MOVERS

A prime mover may be defined as a device which uses energy from natural sources and converts it into mechanical energy. By 'prime' we mean first and 'movers' mean which moves. Thus, prime movers means a device which moves first. If a horse is used to pull a tonga then horse is the prime mover. Similarly a tractor used to pull a harrow or a trolley is a prime mover. Diesel loco, electric loco, I.C. engine, steam engine, hydraulic turbine etc. are more examples of a prime mover.

1.2 SOURCES OF ENERGY

The sources of energy may be broadly classified as follows:

1. Conventional energy sources

- (i) Fossil fuels — coal, liquid and gaseous fuels
- (ii) Natural gas
- (iii) Water energy
- (iv) Nuclear energy

2. Renewable (non-conventional) energy sources

- (i) Solar energy
- (ii) Wind energy
- (iii) Sea tides
- (iv) Geothermal energy
- (v) Ocean energy
- (vi) Thermo-electric energy
- (vii) Fuel cells
- (viii) Thermionic energy
- (ix) Magneto-hydrodynamic
- (x) Bio-mass

1.2.1 Fossil Fuels

1. Solid fuels

Wood, peat, coal and coke, various solid wastes from processes like sugarcane crushing, municipal waste etc. fall in this category.

2. Liquid fuels

Petrol, diesel oil, paraffins, heavy fuel oil fall in this category. Most of the liquid fuels are obtained from petroleum by fractional distillation.

The advantages of liquid fuels over solid fuels are :

- (i) Higher calorific value.
- (ii) Economy in space.
- (iii) Cleanliness of the surroundings.
- (iv) Easy control of combustion.
- (v) Elimination of grate.

Renewable energy source.

1.3 TYPES OF PRIME MOVERS

The prime movers may be classified based on the source of energy utilized by them.

A. Thermal

(a) Liquid and gaseous fuels.

(i) Heat engines

1. Internal combustion engines

- Petrol engine

- Diesel engine
 - LPG engines
 - 2. External combustion engines
 - Steam engines
- (ii) Turbines
1. Internal combustion turbines
 - Open cycle gas turbine
 2. External combustion engines
 - Steam turbine
 - Closed cycle gas turbine
- (b) Nuclear power plant
- (c) Geothermal power plant
- (d) Bio-gas plant
- (e) Solar energy converter

B. Non-thermal

(a) Hydraulic turbines

- Pelton turbine
- Francis turbine
- Kaplan turbine

(b) Wind mill

(c) Tidal power plant

1.4 SOME DEFINITIONS

Mass — It is the quantity of matter contained in a body. The mass of a body remains constant irrespective of its distance from the sea level. The fundamental unit of mass is kilogram (kg). It is represented by m .

Weight — The weight of a body is the force with which Earth attracts the body. It is equal to the product of mass of the body and the value of acceleration due to gravity (g) at that place. It is represented by W . Thus, $W = mg$. The S.I. units of weight are Newton (N). Since the value of g varies from place to place, therefore the weight of the body also changes accordingly. An average value of g may be taken as 9.81 m/s^2 .

Force — It is the push or pull acting on a body which changes or tends to change the state of rest or of uniform motion of a body. According to Newton's second law of motion, force is directly proportional to the rate of change of linear momentum of the body.

$$\begin{aligned}\text{Force, } F &\propto \frac{d}{dt}(mv), \text{ where } v = \text{velocity} \\ &= k \frac{d}{dt}(mv) \\ &= km \frac{dv}{dt}, \text{ if } m \text{ remains constant} \\ &= kma, \text{ where } a = \text{acceleration.}\end{aligned}$$

If $F = 1$, $m = 1$, $a = 1$, then $k = 1$. Therefore, one unit of force is that which produces an acceleration of 1 m/s^2 in a mass of 1 kg . Thus

$$F = ma$$

and the unit force is called 1 N. Therefore,

$$1N = 1 \text{ kg. m/s}^2$$

The bigger units of force in S.I. units are kilo Newton (kN) and Mega Newton (MN).

Density. It is defined as the ratio of mass to volume of the body. Thus,

$$\text{Density, } \rho = \frac{m}{V}$$

where V = volume

Its S.I. units are kg/m^3 .

Specific weight. It is defined as the multiplication of density and acceleration due to gravity.

Thus,

Specific weight, $\gamma = \rho g$

Its S.I. units are N/m^3 .

Specific Volume. It is defined as the volume per unit mass. Thus,

$$\text{Specific volume, } v = \frac{V}{m} = \frac{1}{\rho}$$

Its units are m^3/kg .

1.5 PRESSURE

The pressure (p) acting on a body is defined as the force exerted normal to unit area of the boundary of the body. Thus,

$$\text{Pressure, } p = \frac{F}{A}$$

where F = force acting on the boundary

A = area on which force is acting

Its S.I. units are N/m^2 or Pascal (Pa).

The bigger units are kilo Pascal (kPa), Mega Pascal (MPa) and Giga Pascal (GPa).

$$1 \text{ kPa} = 10^3 \text{ Pa}$$

$$1 \text{ MPa} = 10^6 \text{ Pa}$$

$$1 \text{ GPa} = 10^9 \text{ Pa}$$

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

Atmospheric pressure — It is the pressure exerted by the atmosphere. The atmospheric pressure varies from place to place. At sea level, the atmospheric pressure is,

$$p_{\text{atm}} = 101.325 \text{ kN/m}^2$$

$$= 1.01325 \times 10^5 \text{ Pa}$$

$$= 1.01325 \text{ bar}$$

Standard atmospheric pressure — It is the pressure of atmospheric air at mean sea level. It is defined as the pressure produced by a column of mercury 760 mm high having a density of $13.59509 \times 10^3 \text{ kg/m}^3$ and acceleration due to gravity being 9.80665 m/s^2 . Thus,

$$\begin{aligned} SAP &= \frac{\rho gh}{10^3} \text{ N/m}^2 \\ &= \frac{13.59509 \times 10^3 \times 9.80665 \times 760}{10^3 \times 10^5} = 1.01325 \text{ bar.} \end{aligned}$$

The mercury barometer is used to measure atmospheric pressure.

Gauge pressure — It is the pressure measured by means of a pressure gauge. It is represented by p_{gauge} . Most of the pressure gauges read the difference between absolute pressure and local atmospheric pressure.

Absolute pressure — It is equal to the sum of atmospheric pressure and gauge pressure. Thus

$$\text{Absolute pressure, } P_{\text{abs}} = P_{\text{atm}} + P_{\text{gauge}}$$

Vacuum pressure — The pressure below atmospheric pressure is called vacuum. It is equal to the difference between atmospheric pressure and absolute pressure below atmosphere. Thus,

$$P_{\text{abs}} = P_{\text{atm}} - P_{\text{vacuum}}$$

The vacuum pressure is measured in Torr.

$$1 \text{ Torr} = 133.33 \text{ Pa}$$

A manometer is used to measure vacuum pressure.

where

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Thus,

Specific weight,

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Specific volume,

Its units are m^3/kg .

$$V = \text{volume}$$

$$\gamma = \rho g$$

$$v = \frac{V}{m} = \frac{1}{\rho}$$

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The relationship between various pressures is shown in Fig. 1.1.

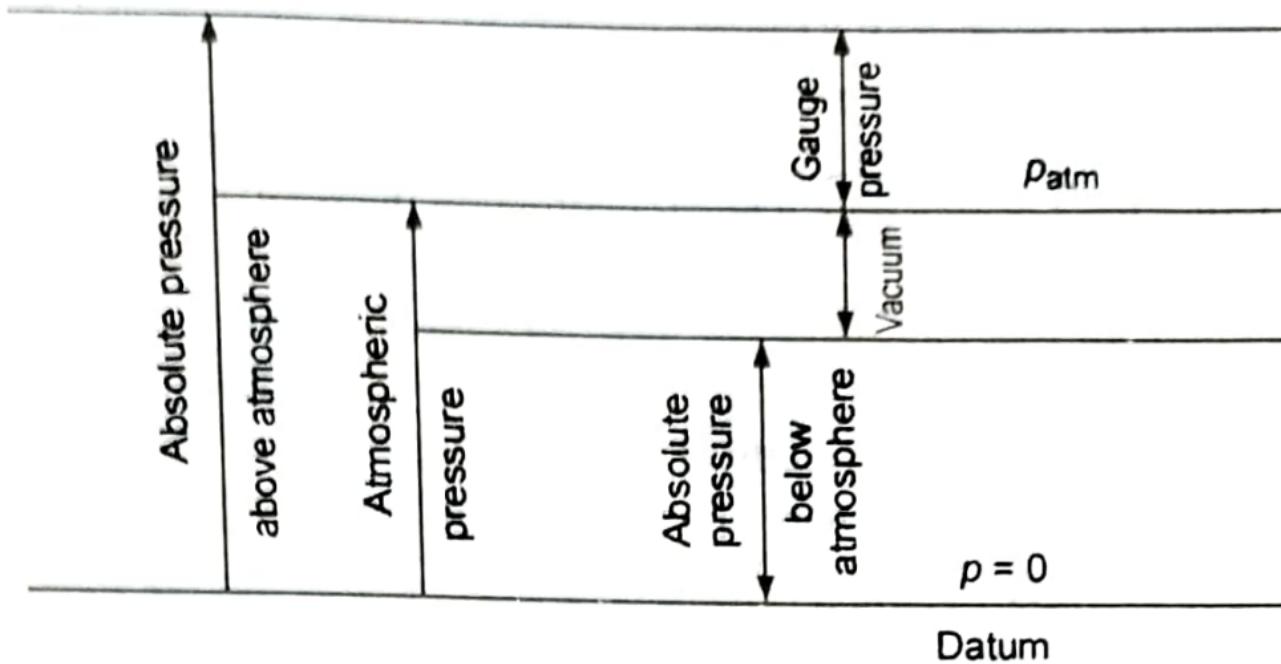


Fig. 1.1 Relationship between various pressures.

H = head, m.

1.7 THERMODYNAMICS

Thermodynamics may be defined as follows:

- The science that deals with energy *interactions between material systems*.

- The science of energy transfer and its effect on the physical properties of substances.
- The science which deals with the transformation of energy of all kinds from one form to another.

The scope of applied thermodynamics is restricted to the study of heat and work and the conversion of one into the other.

The applications of thermodynamics are:

- Design of energy converting devices, such as, steam engines, internal combustion engines, steam and gas turbines, fuel cells, thermoelectric and thermionic generators.
- Design of refrigerators, air-conditioners, heat transfer and reaction equilibrium, etc.

1.8 THERMODYNAMIC SYSTEM

A thermodynamic system is defined as a prescribed region or space of finite quantity of matter surrounded by an envelope called the boundary. The boundary may be a real physical surface, such as the walls of a vessel, internal combustion engine cylinder, etc. or it may be an imaginary surface enclosing some matter, such as, steam, gas, vapour, etc. The boundary may be either fixed or moving.

In a broad sense, a thermodynamic system may be defined as a specified region wherein changes due to transfer of mass or energy or both are studied.

Surroundings (or Environment). The space and matter external to the thermodynamic system is called the surroundings. The system is separated from the surroundings by the system boundary. The thermodynamic system, boundary and surroundings are shown in Fig. 1.3.

Universe. The thermodynamic system and surroundings put together is called the universe.

1.8.1 Types of Systems

There are three types of systems: (1) Closed system, (2) Open system, and (3) Isolated system.

1. Closed System: In a closed system, **the mass within the boundary of the system remains constant and only the energy (heat and work) may transfer across its boundary.** Such a closed system is shown in Fig. 1.4. Some examples of closed system are :

- (i) The gas confined between a piston and cylinder.
- (ii) A rubber balloon filled with air and tightly closed.
- (iii) Pressure cooker.
- (iv) Dash pot.

2. Open System: In an open system, **the mass as well as the energy transfer across its boundaries may take place.** Such an open system is shown in Fig. 1.5. Most of the thermodynamic systems are of the open type. Some examples of open systems are :

- (i) Air compressors
- (ii) Boilers
- (iii) Turbines
- (iv) Water pumps
- (v) Nozzles and diffusers
- (vi) Steam engines, and
- (vii) Internal combustion engines, etc.

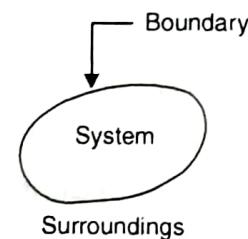


Fig. 1.3 Representation of system, boundary and surroundings.

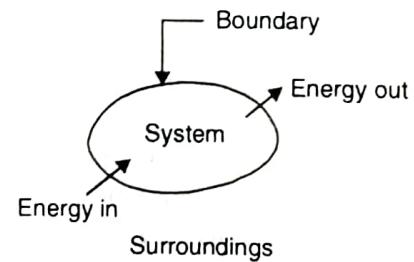


Fig. 1.4 A closed system

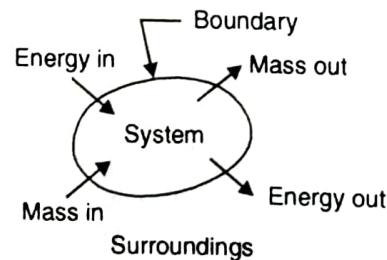


Fig. 1.5 An open system

✓ **3. Isolated System:** In an isolated system, **neither mass nor energy transfer takes place across its boundaries.** Such an isolated system is shown in Fig. 1.6. Some examples of isolated system are:

- (i) Thermos flask
- (ii) The universe.

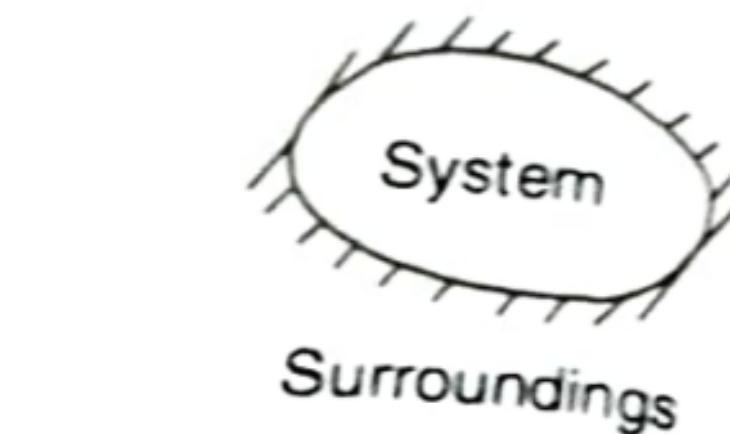


Fig. 1.6 An isolated system

1.9 THERMODYNAMIC PROPERTIES

Every system has some characteristics by which its physical condition may be described. Such characteristics are called properties of the system. A property of a system is its measurable characteristics describing the system, such as, mass, volume, physical composition, temperature, pressure, surface area, velocity, electrical potential, etc. Properties may be classified into two categories: (1) intensive, and (2) extensive.

- 1. Intensive Properties:** These properties are independent of the mass of the system, e.g., temperature, pressure, density, velocity, height, viscosity, etc. specific heat capacity
- 2. Extensive Properties:** These properties are dependent on the mass of the system, e.g., volume, energy (internal, potential and kinetic), surface area, electric charge, magnetization, etc. Enthalpy, heat capacity

1.9.1 Macroscopic and Microscopic Approach

In the macroscopic approach, a certain quantity of matter is considered without taking into account the events occurring at molecular level. This is also called classical thermodynamics point of view. The thermodynamic system is explained by measurable properties, such as, pressure, volume, etc.

In the microscopic approach, the system is considered to be made up of very large number of discrete particles called molecules. This is called statistical thermodynamics point of view.

~~1.10~~ PHASE

If a quantity of matter is homogeneous or uniform throughout in physical structure and chemical composition, it is termed as a phase. The phase may be single, double or multiple constituents. Some examples of a single phase are:

- A piece of iron.
- Gas contained in a vessel.
- A mixture of gases.
- Pure water.
- Liquid oxygen.

- Salt dissolved in water.
- Sugar dissolved in milk.

Some examples of two phases are:

- Ice in water.
- Alcohol in water.
- Steam and water.
- Oil in water.
- Mild steel.

Some examples of three phases are:

- Mixture of solid, liquid and gas.
- Steam, water and ice at the triple point.

On the basis of phases, the system may also be classified as:

(a) **Homogeneous System.** If a system consists of a single phase, it is called a homogeneous system, e.g., mixture of air and water vapour, solution of ammonia in water (aqua ammonia), water and nitric acid, etc.

(b) **Heterogeneous System.** If the system consists of more than one phase, it is called a heterogeneous system, e.g., water and steam, ice and water, water and oil, etc.

1.11 STATE AND PATH

State. The state is a unique condition of the system. It is the condition of a system at an instant of time described by its properties, such as, pressure, volume, temperature, density, etc.

Path. A thermodynamic system passing through a series of states constitutes a path.

In Fig. 1.8, point (p_1, v_1) and point (p_2, v_2) represent the states at 1 and 2. Along curve 12, the states of pressure and volume change from point to point.

1.12 PROCESS AND CYCLE

Process. A process is defined as a transition in which a system changes from one initial state to a final state. In a process, the path is completely specified. In Fig. 1.9, the line 12 is the process that has taken place.

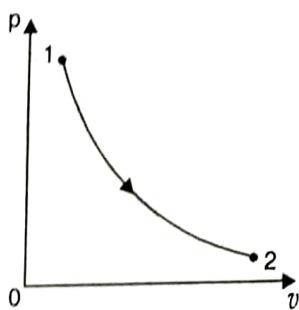


Fig. 1.9 Representation of a process.

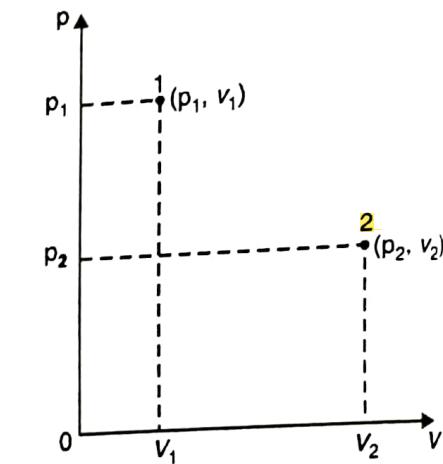


Fig. 1.8 Representation of state.

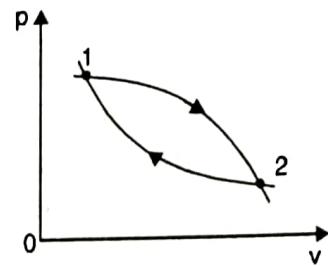


Fig. 1.10 Representation of a cycle.

Cycle. A thermodynamic cycle is defined as a series of state changes such that the final state is identical with the initial state. As shown in Fig. 1.10, 1 – 2 – 1 is the cycle.

1.12.1 Types of Processes

The various types of thermodynamic processes are:

1. Cyclic process.
2. Quasi-static process.
3. Reversible process.
4. Irreversible process.

1. Cyclic Process: In a cyclic process, the system undergoes through a series of processes from one state and ends with the same initial state by forming a complete cycle so that the properties of the system at the end of the cycle are the same as at the beginning. A cyclic process is shown in Fig. 1.11 on $p-v$ diagram.

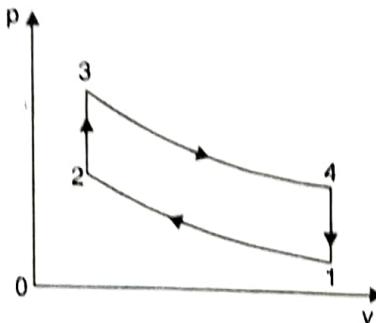


Fig. 1.11 Cyclic process.

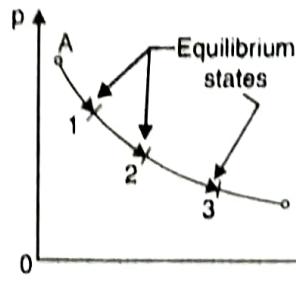


Fig. 1.12 A quasi-static process.

Quasi-static process is a process that happens so slowly (infinitesimal rate of change) that the system remains in equilibrium with itself throughout the entire process

2. Quasi-static Process: A quasi-static process is one in which the system deviates from one equilibrium state by only infinitesimal amounts throughout the entire process. In other words, a process closely approximating to a succession of equilibrium states is called quasi-static process, as shown in Fig. 1.12. It is a nearly reversible process.

3. Reversible Process: A process is said to be reversible if the system passes through a continuous series of equilibrium states. In other words, a process is called reversible if the initial state together with all energies transferred or transformed during the process can be completely restored in both system and environment. A process will be reversible only if there is no friction in the system. A reversible process is shown in Fig. 1.13. Some examples of reversible process are:

- Frictionless adiabatic expansion and compression of a fluid.
- Polytropic expansion and compression of a fluid.
- Isothermal expansion or compression of a gas.
- Extension and compression of a spring.
- Electrolysis.
- Evaporation.

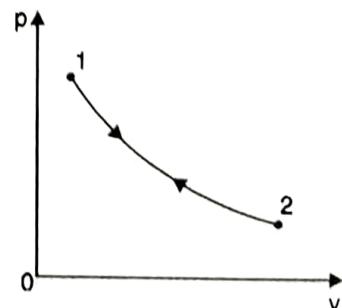


Fig. 1.13 Reversible process.

A reversible process is a thermodynamic process in which the system and surroundings can be restored to their initial state from the final state without producing any changes in the universe. System & surrounding are in equilibrium with each other at each instant of the process.

Conditions for Reversibility

The following conditions must be satisfied by a reversible process:

- There should be no heat transfer across finite temperature difference.
- The process should be carried out with absolute slowness so that the system is always in thermal, chemical and mechanical equilibrium.
- After the process, both the system and surroundings must be restored to the original state, i.e., there should be no friction, viscosity, electrical resistance, inelasticity, etc.

4. Irreversible Process: A process is said to be irreversible if a system passes through a sequence of non-equilibrium states.

An irreversible process will not retrace the reverse path to restore the original state, as shown in Fig. 1.14. Some examples of irreversible process are :

- Mechanical and fluid friction.
- Combustion of air and fuel.
- Throttling.
- Free expansion.

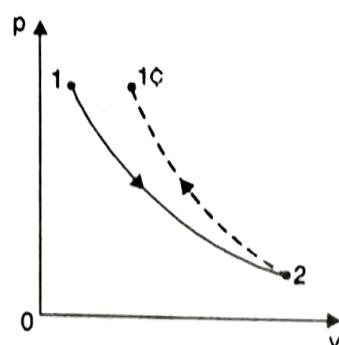


Fig. 1.14 Irreversible process.

1.14 WORK

In thermodynamics, work is defined as **energy transition, i.e. a form of energy interaction between two systems**. Work may be defined as follows:

- Energy transferred, without transfer of mass across the boundary of a system because of an intensive property difference other than temperature that exists between system and surroundings.

- An interaction between a system and surroundings, and is done by a system if the sole effect external to the boundaries of the system could have been the raising of a weight.

The work done is represented by the area under the $p-V$ diagram, as shown in Fig. 1.15. Thus,

Work,

$$W_{1-2} = \int_{V_1}^{V_2} p dV$$

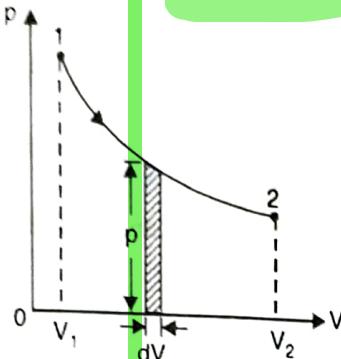


Fig. 1.15 Concept of work.

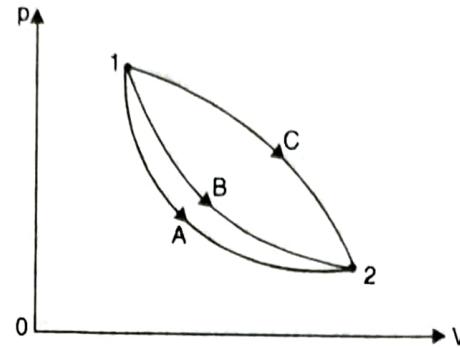


Fig. 1.16 Work a path function.

The SI units of work are N.m or Joule. Work cannot be stored in a system as it exists only during the transition. The work transferred into a system exists as stored energy. Work is a path function, as shown in Fig. 1.16.

$$W_A \neq W_B \neq W_C \text{ as}$$

$$W_{1-2} = \int_1^2 \delta W$$

Sign Convention for Work. Work done by the system is considered positive and work done on the system as negative.

1.15 HEAT

Heat is an energy interaction between the system and the surroundings. Heat may be defined as the energy transferred without transfer of mass across the boundary of a system due to a temperature difference between the system and the surroundings. The SI units of heat are Joule.

A wall which does not allow the flow of heat is called an adiabatic wall whereas a wall which permits the flow of heat is a diathermal wall. The similarities and dis-similarities between work and heat are:

Similarities:

- Energy interaction.
- Transient phenomenon.
- Boundary phenomenon.
- Not the property of the system.
- Path functions.

Dis-similarities :

- Heat transfer is the energy interactions due to temperature difference only. All other energy interactions are work interactions.
- Heat is a low grade energy while work is a high grade energy.

Sign Convention for Heat. Heat received by the system is considered positive and heat rejected by the system as negative.

1.16 THERMODYNAMIC EQUILIBRIUM

A system is said to be in a state of thermodynamic equilibrium if the value of the property is same

at all points in the system. An isolated system always reaches, in due course of time, a state of thermodynamic equilibrium and can never depart from it spontaneously.

A system will be in a state of thermodynamic equilibrium, if the conditions for the following three types of equilibrium are satisfied :

Mechanical Equilibrium. A system is said to be in a state of mechanical equilibrium if there exists no unbalanced force either in the interior of the system or between the system and the surroundings.

Chemical Equilibrium. A system is said to be in a state of chemical equilibrium if there exists no chemical reaction or transfer of matter from one part of the system to another, such as, diffusion or solution.

Thermal Equilibrium. A system is said to be in thermal equilibrium if there exists an uniformity of temperature throughout the system or between system and surroundings.

A system is said to be in thermodynamic equilibrium only if the system is capable of undergoing a spontaneous change in any macroscopic property when it is isolated.

1.18 CONCEPT OF AN IDEAL GAS

Many gases which do not change their phase during a thermodynamic process and obey a set of common rules governing change of their properties are called ideal (or perfect) gases. For an ideal gas, the following conditions are satisfied :

1. Gas law:

$$pv = RT$$

Where

p = pressure, kPa

v = specific volume, m^3/kg

R = 0.287 kJ/kg. K , characteristic gas constant

T = absolute temperature, K

2. The internal energy is a function of temperature alone, i.e.,

$$U = f(T) \text{ only}$$

None of the gases existing in nature are ideal in real sense. However, most of the gases like nitrogen, oxygen, helium, hydrogen, carbon dioxide and air obey gas laws at normal temperature and pressure. These gases may be regarded as ideal or perfect gases for all practical purposes.

1.18.1 Characteristic Gas Equation

By combining Boyle's and Charles' laws a relation between pressure, temperature and volume may be found. The resulting equation is called the equation of state for a gas. Consider one kg mass of a perfect gas undergoing changes from p_1, v_1, T_1 to p_2, v_2, T_2 by the two processes 1 - A and A - 2, as shown in Fig. 1.17. 1 - A is a constant pressure process obeying Charles' law and A - 2 is an isothermal process obeying Boyle's law.

Process 1 - A :

$$\frac{v_1}{T_1} = \frac{v_A}{T_A}$$

or $v_A = \frac{v_1 T_A}{T_1} = \frac{v_1 T_2}{T_1} \quad (\because T_2 = T_A) \quad \dots(1)$

Process A - 2 :

$$p_2 v_2 = p_A v_A = p_1 v_A \quad (\because p_A = p_1)$$

or $v_A = \frac{v_2 p_2}{p_1}$

From Eqs. (1) and (2), we get

$$\frac{v_1 T_2}{T_1} = \frac{v_2 p_2}{p_1}$$

or $\frac{p_1 v_1}{T_1} = \frac{p_2 v_2}{T_2}$

or $\frac{pV}{T} = \text{Const.} = R$

This is called the characteristic gas equation.

For a gas having m kg mass,

$$pV = mRT$$

$$p v_{\text{mole}} = R_0 T$$

$$pV = nR_0 T$$

where

V = total volume of gas, m^3

n = number of kg. moles of a gas = m/M

m = mass of gas, kg

M = molecular weight of a gas, kg/kgmoles

$v_{\text{mole}} = \text{mole volume} = \frac{V}{n}, \text{m}^3/\text{kgmole}$

R_0 = universal gas constant

= 8.314 kJ/kgmole.K

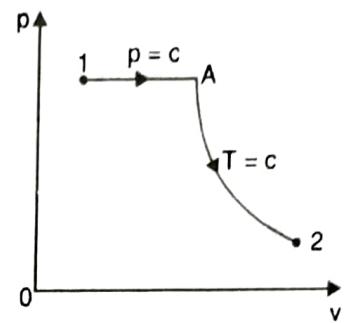


Fig. 1.17 Derivation of characteristic gas equation.

1.19 ZEROTH LAW OF THERMODYNAMICS

This law states that if two systems are each in thermal equilibrium with a third system, they are also in thermal equilibrium with each other.

Let us consider three bodies A, B and C. If body A and body B are in thermal equilibrium with body C then according to this law, bodies A and B will also be in thermal equilibrium with each other.

This law was formulated after the first law of thermodynamics and hence named zeroth law. This law is the basis of concept of temperature as well as of all temperature measurements.

1.22 FIRST LAW OF THERMODYNAMICS

The first law of thermodynamics is the **law of conservation of energy**, i.e., energy can neither be **created nor destroyed but only converted from one form to another**. With reference to thermodynamics, heat and work are transferable.

This law may be stated as follows:

- For a system undergoing a cyclic change the algebraic sum of the work delivered to the surroundings is equal to the algebraic sum of heat taken from the surroundings, i.e.

$$\oint \delta Q = \oint \delta W$$

where

\oint = integral over a cycle

δQ = infinitesimal heat transfer

δW = infinitesimal work transfer

- Heat and work are mutually convertible one into the other.
- The total energy of an isolated system, in all its forms, remains constant.
- It is impossible to construct a perpetual motion machine of the first kind (PMM – 1).
- PMM – 1 is capable of producing energy without corresponding expenditure of the energy.

1.24 ENTHALPY

- 1 The summation of internal energy and flow energy is defined as enthalpy, and is given by:

Enthalpy,

$$H = U + p.dV \quad (\text{or flow work})$$

Specific enthalpy,

$$h = u + p.v$$

Where u = specific internal energy, v = specific volume

Enthalpy is an extensive property, as it depends upon mass.

Now

$$\begin{aligned} dq &= du + p.dv \quad (\text{a variation of the First Law of Thermodynamics}) \\ &= du + d(pv) - v.dp \\ &= d(u + pv) - v.dp \\ &= dh - v.dp \end{aligned}$$

For a constant pressure process, $dp = 0$.

$$\therefore dq = dh$$

3 Thus, enthalpy change is the heat supplied at constant pressure.

4 Also for a perfect gas, $u = c_v T$ and $pV = RT$

$$\therefore h = c_v T + RT$$

where c_v = specific heat at constant volume.

or $h = T(c_v + R) = c_p T$

Where c_p = specific heat at constant pressure
or $H = mc_p T$

1.25 SPECIFIC HEATS

Specific heat of a substance may be defined as the amount of heat required to raise the temperature of a unit mass of the substance through one degree. The SI units of specific heat are J/kg.K. The solids and liquids have only one specific heat whereas gases have two specific heats, i.e., at constant volume, and constant pressure.

1.25.1 Specific Heat at Constant Volume, c_v

The quantity of heat required to raise the temperature of unit mass of the gas through one degree at constant volume is called specific heat at constant volume, c_v .

Consider a quantity of gas contained in a fixed container shown in Fig. 1.29. When heat is supplied to the gas, the volume of the gas will not change because the container is fixed but there will be an increase in pressure and temperature of the gas.

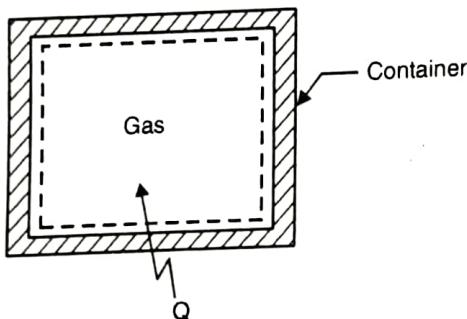


Fig. 1.29 Heating at constant volume.

Let

m = mass of gas

T_1 = initial temperature of gas

T_2 = final temperature of gas

Total heat supplied to gas at constant volume,

$$Q_{1-2} = m c_v (T_2 - T_1) = m c_v dT$$

Since volume of gas remains constant, therefore, no work is done by the gas and the whole heat supplied remains within the mass of gas. This represents the increase in internal energy of the gas.

$$u = c_v \cdot dT$$

$$U = m c_v \cdot dT$$

or

1.25.2 Specific Heat at Constant Pressure, c_p

Consider a quantity of gas contained in a cylinder tightly fitted with a frictionless piston carrying a load W , as shown in Fig. 1.30. When heat is supplied to the gas, there will be increase in temperature of the gas. But the tendency for the pressure to rise will be counter-balanced by the piston rising upward, thereby lifting the load and increasing the volume of gas. Hence the heat supplied increases the temperature and volume of gas and does work.

The specific heat at constant pressure, c_p , may be defined as the quantity of heat required to raise the temperature of unit mass of a gas through one degree.

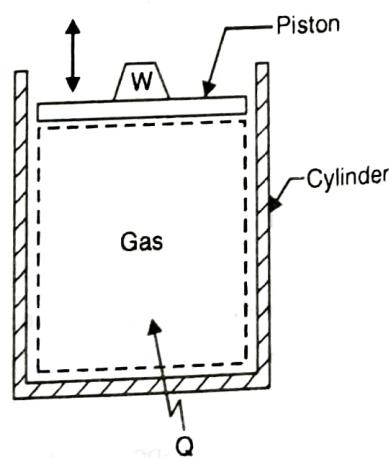


Fig. 1.30 Heating at constant pressure.

Let m = mass of gas

T_1, T_2 = initial and final temperatures respectively

V_1, V_2 = initial and final volumes of gas respectively

$$Q_{1-2} = mc_p(T_2 - T_1)$$

The heat supplied to the gas at constant pressure is utilized for two purposes:

- (i) To increase the internal energy of the gas.
- (ii) For doing external work.

Thus, the heat energy supplied is utilized in increasing the enthalpy of gas.

Now $dU = mc_v(T_2 - T_1)$

And $W_{1-2} = p(V_2 - V_1) = mR(T_2 - T_1)$

1.25.3 Relation Between the Specific Heats

At constant pressure,

$$\begin{aligned} Q_{1-2} &= mc_p \cdot dT = dU + W_{1-2} \\ &= mc_v \cdot dT + p \cdot dV = mc_v \cdot dT + mR \cdot dT \\ &= m(c_v + R) \cdot dT \\ \therefore c_p &= c_v + R \\ \text{or } c_p - c_v &= R \end{aligned}$$

Now $\gamma = \frac{c_p}{c_v}$, is called the adiabatic constant or index.

$$\therefore c_v(\gamma - 1) = R$$

$$\text{or } c_v = \frac{R}{\gamma - 1}$$

The values of c_p , c_v and γ for gases are given in Table 1.3.

Table 1.3 Values of c_p , c_v and γ at 38°C.

Gas	c_p , kJ/kg.K	c_v , kJ/kg.K	γ
Air	1.005	0.718	1.4
Carbon dioxide	0.858	0.67	1.28
Carbon monoxide	1.042	0.745	1.4
Hydrogen	14.343	10.219	1.4
Nitrogen	1.038	0.745	1.4
Oxygen	0.921	0.661	1.39

1.26 PERFECT GAS LAWS

Boyle's Law

When temperature remains constant, the volume of a given mass of a perfect gas is inversely proportional to the absolute pressure. In other words, the product of absolute pressure ' p ' and volume ' V ' of a given quantity of gas is constant when the temperature remains constant.

$$pV = \text{const.}$$

Charles' Law

- If the pressure remains constant, the volume of a given mass of gas varies directly as the absolute temperature ' T '.

$$\frac{V}{T} = \text{const.}$$

- 
- If the volume remains constant, the pressure of a given mass of gas varies directly with absolute temperature.

$$\frac{P}{T} = \text{const.}$$

Where $T = 273.15 + C^\circ$ Kelvin (K).