Chapter -11

THERMODYNAMICS

WE LEARN ABOUT

Brief introduction

sadi carnot was a French physicist, who gave the first successful theoretical account of heat engines, now known as the Carnot cycle, there by laying the foundations of the second law of thermodynamics. He is often described as the " Father of thermodynamics", being responsible for such concepts as Carnot efficiency, Carnot theorem, Carnot heat engine, and others. In 1824 the principle of conservation of energy was still immature and controversial, and an exact formulation of the first law of thermodynamics was still more than a decade away. The mechanical equivalent of heat was not identified for another two decades. The prevalent theory of heat was the caloric theory, which regarded heat as a sort of weightless, invisible fluid that flowed when out of equilibrium.

11.1.1 INTRODUCTION

In previous chapter we have studied thermal properties of matter. In this chapter we shall study laws that govern thermal energy. We shall study the processes where work is converted into heat and vice versa. In winter, when we rub our palms together, we feel warmer; here work done in rubbing produces the 'heat'. Conversely, in a steam engine, the 'heat' of the steam is used to do useful work in moving the pistons, which in turn rotate the wheels of the train.

In physics, we need to define the notions of heat, temperature, work, etc. more carefully. Historically, it took a long time to arrive at the proper concept of 'heat'. Before the modern picture, heat was regarded as a fine invisible fluid filling in the pores of a substance. On contact between a hot body and a cold body, the fluid (called caloric) flowed from the colder to the hotter body ! This is similar to what happens when a horizontal pipe connects two tanks containing water up to different heights. The flow continues until the levels of water in the two tanks are the same. Likewise, in the 'caloric' picture of heat, heat flows until the 'caloric levels' (i.e., the temperatures) equalise.

In time, the picture of heat as a fluid was discarded in favour of the modern concept of heat as a form of energy. An important experiment in this connection was due to Benjamin Thomson (also known as Count Rumford) in 1798. He observed that boring of a brass cannon generated a lot of heat, indeed enough to boil water. More significantly, the amount of heat produced depended on the work done (by the horses employed for turning the drill) but not on the sharpness of the drill. In the caloric picture, a sharper drill would scoop out more heat fluid from the pores; but this was not observed. A most natural explanation of the observations was that heat was a form of energy and the experiment demonstrated conversion of energy from one form to another–from work to heat.

Thermodynamics is the branch of physics that deals with the concepts of heat and temperature and the inter-conversion of heat and other forms of energy. Thermodynamics is a macroscopic science. It deals with bulk systems and does not go into the molecular constitution of matter. In fact, its concepts and laws were formulated in the nineteenth century before the molecular picture of matter was firmly established. Thermodynamic description involves relatively few macroscopic variables of the system, which are suggested by common sense and can be usually measured directly. A microscopic description of a gas, for example, would involve specifying the co-ordinates and velocities of the huge number of molecules constituting the gas. The description in kinetic theory of gases is not so detailed but it does involve molecular distribution of velocities. Thermodynamic description of a gas, on the other hand, avoids the molecular description altogether. Instead, the state of a gas in thermodynamics is specified by macroscopic variables such as pressure, volume, temperature, mass and composition that are felt by our sense perceptions and are measurable⊛.

The distinction between mechanics and thermodynamics is worth bearing in mind. In mechanics, our interest is in the motion of particles or bodies under the action of forces and torques. Thermodynamics is not concerned with the motion of the system as a whole. It is concerned with the internal macroscopic state of the body. When a bullet is fired from a gun, what changes is the mechanical state of the bullet (its kinetic energy, in particular), not its temperature. When the bullet pierces a wood and stops, the kinetic energy of the bullet gets converted into heat, changing the temperature of the bullet and the surrounding layers of wood. Temperature is related to the energy of the internal (disordered) motion of the bullet, not to the motion of the bullet as a whole.

⊛ *Thermodynamics may also involve other variables that are not so obvious to our senses e.g. entropy, enthalpy, etc., and they are all macroscopic variables.*

11.1.1 HEAT AND TEMPERATURE

Heat is a form of energy called thermal energy.

When we bring a hot body in thermal contact with a cold body, then temperature of the hot body decreases and temperature of the cold body increases. The energy of the hot body decreases and that of the cold body increases. This transfer of energy takes place until the two bodies attain a common temperature, showing that the two bodies are at the same thermal condition.

So heat flows from a body at higher temperature to a body at lower temperature. This leads to the definition of heat.

Definition of Heat: The energy which is transferred between a system and its surroundings, as a I result of temperature difference only, is called heat.

Unit: joule in SI, Calorie in CGS system.

Dimensional formula: [ML^2T^{-2}].

Note:

Heat is certainly energy, but it is the energy in transit.

Definition of Temperature:

Temperature is defined as the thermal condition of a body which determines its ability to transfer heat to other bodies.

(a) Temperature is the measure of degree of hotness or coldness of a body.

(b) Temperature determines the direction of net flow of heat.

Unit: Kelvin (K) in absolute scale.

Dimensional formula [K] or [θ].

11.1.2 UNITS OF HEAT

Various units of heat are obtained by choosing different units of mass and temperature. The most commonly used unit is calorie.

Calorie : It is the unit of measuring heat. Calorie is generally defined as the amount of heat required by one gram of water to rise its temperature by 1 °C.

In fact, one gram of water does not require die same amount of heat to rise the temperature by 1°C at all ranges of temperature. When one gram of water is heated from 0°C to 100°C, the mean value of the amount of heat required for 1°C rise of temperature is found to be the same as it requires between 14.5°C and 15.5°C. Hence mean calorie is defined as

Def: "The amount of heat required by one gram of water to rise, its temperature from 14.5°C to 15.5°C (or 58.1°F to 59.9°F) at a pressure of 1 atm."

11.1.3 THERMODYNAMIC SYSTEMS

To study the relation between the heat and mechanical work, a system is required. This relation can be studied by studying the changes in the properties of a system due to exchange of energy in the form of work.

A system is defined as a region in space. The region may contain a collection of large number of atoms or molecules with in a real or imaginary boundary.

Ex: A gas or vapour enclosed in a cylinder having a movable piston.

Surroundings:

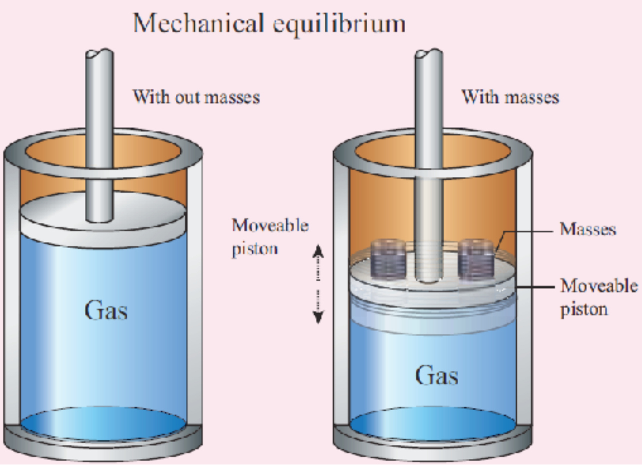
The medium or matter or vacuum that surrounds the system which may participate in the process of exchange of matter or energy or both with the system.

11.1.4 THERMAL EQUILIBRIUM

When a hot cup of coffee is kept in the room, heat flows from coffee to the surrounding air. After some time the coffee reaches the same temperature as the surrounding air and there will be no heat flow from coffee to air or air to coffee. It implies that the coffee and surrounding air are in thermal equilibrium with each other. Two systems are said to be in thermal equilibrium with each other if they are at the same temperature, which will not change with time.

Mechanical equilibrium:

Consider a gas container with piston as shown in Fig. When some mass is placed on the piston, it will move downward due to downward gravitational force and after certain humps and jumps the piston will come to rest at a new position. When the downward gravitational force given by the piston is balanced by the upward force exerted by the gas, the system is said to be in mechanical equilibrium. A system is said to be in mechanical equilibrium if no unbalanced force acts on the thermodynamic system or on the surrounding by thermodynamic system.



Mechanical equilibrium

Chemical equilibrium:

If there is no net chemical reaction between two thermodynamic systems in contact with each other then it is said to be in chemical equilibrium.

Thermodynamic equilibrium:

If two systems are set to be in thermodynamic equilibrium, then the systems are at thermal, mechanical and chemical equilibrium with each other. In a state of thermodynamic equilibrium the macroscopic variables such as pressure, volume and temperature will have fixed values and do not change with time.

11.1.5 THERMODYNAMIC STATE VARIABLES

In mechanics velocity, momentum and acceleration are used to explain the state of any moving object. In thermodynamics, the state of a thermodynamic system is represented by a set of variables called thermodynamic variables.

Examples: Pressure, temperature, volume and internal energy etc.

The values of these variables completely describe the equilibrium state of a thermodynamic system. Heat and work are not state variables rather they are process variables.

There are two types of thermodynamic variables: Extensive and Intensive

Extensive variable depends on the size or mass of the system.

Example: Volume, total mass, entropy, internal energy, heat capacity etc.

Intensive variables do not depend on the size or mass of the system.

Example: Temperature, pressure, specific heat capacity, density etc.

11.1.6 EQUATION OF STATE:

The equation which connects the state variables in a specific manner is called equation of state. A thermodynamic equilibrium is completely specified by these state variables by the equation of state. If the system is not in thermodynamic equilibrium then these equations cannot specify the state of the system

An ideal gas obeys the equation $PV = NkT$ at thermodynamic equilibrium. Since all four macroscopic variables (P,V,T and N) are connected by this equation, we cannot change one variable alone. For example, if we push the piston of a gas container, the volume of the gas will decrease but pressure will increase or if heat is supplied to the gas, its temperature will increase, pressure and volume of the gas may also increase.

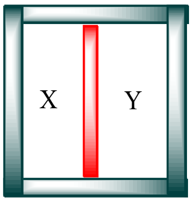
There is another example of equation of state called van der Waals equation. Real gases obey this equation at thermodynamic equilibrium. The air molecules in the room truly obey van der Waals equation of state. But at room temperature with low density we can approximate it into an ideal gas.

11.2.1 ZEROTH LAW OF THERMODYNAMICS

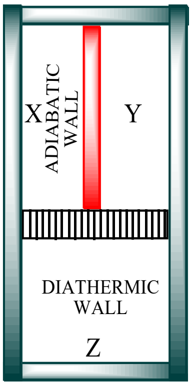
Def: If two systems are in thermal equilibrium with a third system separately then they must be in thermal equilibrium with each other.

Consider two systems X and Y separated from each other by an adiabatic wall which does not allow the heat energy to pass through. The systems are not influenced by each other or influenced by the surroundings. So they are in separate state of equilibrium with different temperatures T_1 and T_2, Fig (i). When this combination is arranged on another system Z with a diathermic wall which allows the heat energy to pass through as shown in the fig (ii). Now the two systems X and Y are in thermal contact with Z separately. Due to exchange of heat energy between systems X and Z they attain thermal equilibrium showing that these two are at same temperature T. Similarly system Y and system Z attain thermal equilibrium due to exchange of heat energy. The systems Y and Z attain the same temperature T.

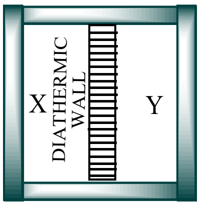
The adiabatic wall between the systems X and Y is replaced by another diathermic wall and the combination is seperated from the system Z as shown in fig (iii). It is observed that no further changes in their temperatures showing that there is no exchange of heat energy between the systems X and Y. As they are at same temperature the two systems X and Y are said to be in thermal equilibrium.



Fig(i)



Fig(ii)



Fig(iii)

Thus the temperature is defined as "the property which determines whether a system is in thermal equilibrium or not with the neighbouring system". If the two systems are at different temperature they are not in thermal equilibrium.

The property of thermal equilibrium is used in the measurement of temperature.

i) Zeroth law of thermodynamics, states that 'two systems in thermal equilibrium with a third system separately are in thermal equilibrium with each other'. R.H. Fowler formulated this law in 1931 long after the first and second I thermodynamics were stated and so numbered.

ii) The Zeroth Law clearly suggests that when two systems A and B are in thermal equilibrium, there must be a physical quantity that has the same value for both. The thermodynamic variable whose value is equal for two systems in thermal equilibrium is called temperature (T). Thus, if A and B are separately in equilibrium with C, $T_A = T_C$ and $T_B = T_C$.This implies that $T_A = T_B$ i.e. the systems A and B are also in thermal equilibrium.

11.2.2 JOULES LAW AND MECHANICAL]EQUIVALENT OF HEAT

Statement of Joule's Law :

Whenever work is done against friction, heat is generated. The amount of heat generated (H) is directly proportional to the work done (W).

If w is the amount of mechanical work done and H is the amount of heat produced, then

$W \alpha H$ or $W = JH$

Here J is constant of proportionality which is known as Joule's constant or Mechanical equivalent of heat If H= 1, than J = W

So, Joule’s mechanical equivalent of heat is defined as the amount of work required to be done to produce a unit quantity of heat.

Significance:

Whenever heat is gained or lost by a system during some process, the gain or loss of heat can be taken to be equivalent to some amount of mechanical work done. In this way heat is included as a form of energy (thermal energy) in the solutions of the law of conservation of energy.

If H is measured in calories and W in joules, then J = 4.186 J cal^{-1} If, W is measured in ergs then J = 4,186 x 10^7 erg cal^{-1}

J is not a physical quantity. It is a conversion factor to convert units of work from one system to another.

The value of J= 1 in SI where both W& H are measured in Joules.

In SI we write H = W since H and work are measured in Joules

Problem

A person of mass 60 kg wants to lose 5kg by going up and down a 10 m high stairs. Assume he burns twice as much fat while going up than coming down. If 1 kg of fat is burnt on expending 7000 kilo calories, how many times must he go up and down to reduce his weight by 5 kg?

Work done in moving up = mgh

Work done in moving down = mgh/2

Total work done, = $\frac{3}{2}$ mgh = $\frac{3}{2}$ x 60 x 10 x 10

N number of trips

Sol:

$$5 \times 7000 \times 10^3 \times 4.2\ \text{J} = 60 \times 15 \times 10 \times N$$

$$N = \frac{21 \times 7 \times 10^6}{9000} \times 10^3 = 16.3 \times 10^3 \text{ times}$$

Problem:

A lead bullet of mass 21 g travelling at a speed of 100 ms^{-1} comes to rest in a wooden block. If no heat is taken away by the wood, find the rise in temperature of the wood.

(Specific heat of lead = 0.03 calorie/g °C.)

Sol:

kinetic energy of the bullet = heat gained by the bullet $\frac{1}{2}$ $mv^2 = ms\Delta t$

$$\Delta t = \frac{v^2}{2s} = \frac{(100)^2}{2 \times 0.03 \times 4.2 \times 1000} = 39.68°c.$$

11.2.3 HEAT, INTERNAL, ENERGY AND WORK:

The three terms ; heat, internal energy and work will be used very commonly in this chapter. Therefore, it is important to understand them.

As we all know, heat is energy that flows from higher temperature to lower temperature. This flow is only because of difference in temperatures. We may thus define Heat as a form of energy in transit from hot to cold.

Every bulk system consists of a large number of molecules.

Internal energy of a system is the total energy possessed by the system due to molecular motion and molecular configuration.

It is represented by U.

The energy due to molecular motion is called internal kinetic energy U_k. The motion includes translational, rotational and vibrational motion of the molecules. The energy due to molecular configuration is called internal potential energy, U_p. Thus,

$$\boxed{U = U_k + U_p}$$

In case of an ideal gas or perfect gas, we assume that intermolecular forces are zero. Therefore, no work is done, whatsoever , in changing the distance between the molecules. Thus, $U_p = 0$. Hence, internal energy of an ideal gas consists only of internal kinetic energy- which depends only on temperature of the gas.

In a real gas, intermolecular forces are not negligible. Therefore, a definite amount of work has to be done in changing the distance between the molecules. Thus, internal energy of a real gas is sum of internal kinetic energy and internal potential energy of the molecules of the gas. It would obviously, depend on both, the temperature and volume of the gas.

The following points regarding internal energy should be clearly understood :

(i) Internal energy of a system is a macroscopic state variable of the system. The internal energy of a given mass of gas depends on its state described by specified values of pressure, volume and temperature.

(ii) Internal energy of a system does depend on state of the system, but not on how that state was achieved or not on history of achieving the state, i.e., not on the path taken to arrive at that state.

(iii) In thermodynamics, the kinetic energy of the system as a whole is not relevant. It means if the container of the gas is moving as a whole with some velocity, the kinetic energy of the container is not to be included in internal energy.

(iv) The internal energy of a real gas can be changed by changing any of the macroscopic state variables of the gas.

Two ways of changing the state and hence internal energy of a gas are :

(a) Heating the cylinder containing the gas or keeping the cylinder in contact with a body at higher temperature. Some heat flows from hotter body to the gas on account of temperature difference. Therefore, internal energy of the gas increases.

(b) Pushing the piston of the cylinder containing gas in the upward direction. Some work is done by the gas, due to which the internal energy of the gas decreases.

Note that both these things could happen in the reverse direction too. For example, if the gas were in contact with a body at lower temperature, some heat would flow from gas to the surroundings resulting in fall in temperature and hence decrease in its internal energy. Similarly, if the piston were pushed downwards, work will be done on the gas. Therefore, internal energy of the gas would increase.

We observe that heat and work are two different modes of changing state of a thermodynamical system and hence changing internal energy of the system.

Hence, we conclude that heat and work in thermodynamics are not state variables. They are two modes of energy transfer to or from a system resulting in change in internal energy of the system- which is tt state variable.

NOTE:

1. Out of heat, internal energy and work, only internal energy of the system is a state variable. The other two, i.e., heat and work are two modes of energy transfer to or from a system. This energy transfer would result in the change of internal energy of the system.
2. Internal energy of an ideal gas consists only of internal K.E. – which depends only on temperature of the gas.
3. Internal energy of a real gas is sum of internal K.E. and internal P.E. of the molecules of the gas. It would therefore, depend on the temperature and volume of the gas.

11.2.4 THERMODYNAMIC STATE VARIABLES AND EQUATION OF STATE

Themodynamic state variables of a system are the parameters which describe equilibrium states of the system. For example, equilibrium state of a gas is completely specified by the values of pressure, volume, temperature, mass and composition (if there is a mixture of gases). The various state variables of a system are not necessarily independent.

The Equation of State represents the connection between the state variables of a system. For example, the equation of state of an ideal/perfect gas is represented as

$$PV = \mu RT$$

where μ is number of moles of the gas and R is gas constant for one mole of the gas.

For a fixed amount of gas (i.e., for given p), there are only two independent variables of ideal gases, say (P and V) or (T and V). The real gases may have more complicated equations of state.

Further, the thermodynamic state variables are of two types :

(i) Extensive state variables and

(ii) Intensive state variables.

Extensive variables change with the size of the system, but intensive variables do not.

For example, imagine a gaseous system in thermal equilibrium divided into two equal parts. The state variables, volume V, mass M and internal energy U which get halved in each part are extensive variables. The other state variables like pressure P, temperature T and density ρ which remain unchanged for each part are intensive variables.

11.2.5 QUASI STATIC PROCESS

A process or a change in the state variables of a thermodynamical system, which is, infinitely slow is called quasi – static process (meaning–nearly static). The system undergoes changes in its variables like P, V, T so slowly that it continues to be in thermal and mechanical equilibrium with its surroundings throughout all regions of the system and at all times.

Quasi – static process is only an idealised process, in which we imagine that at every stage, the system is an equilibrium state. This visualization is necessary because it is difficult to deal with non equilibrium states of the system.

In a quasi – static process, the difference in the pressure of the system and the external pressure is Infinitesimally small. Similarly, the difference between the temperature of the system and temperature of the surroundings is infinitesimally small. Therefore, a quadristatic process must he infinitely slow.

For example, to take a gas from the state (P, T) to another state (P', T'), via a quasi-static process, we change the external pressure/temperature by a very small amount and allow the system to equalise its pressure/temperature with the surroundings. Continue the process infinitely slowly till the final state (P'/ T') is attained.

Obviously, a quasi – static process is a hypothetical construct. The process must be infinitely slow, should not involve large temp, differences or accelerated motion of the piston of the container.

In our discussion further, we shall assume all thermodynamic processes to be quasi – static, unless stated otherwise.

11.3.1 THERMODYNAMIC PROCESSES

A thermodynamical process is said to take place when some changes occur in the state of a thermodynamic system, i.e,, the thermodynamic parameters of the system change with time.

A system can interact with its surroundings in several ways. The heat energy and work that come into play always obey the first law of thermodynamics.

Following are some of the important thermodynamic processes :

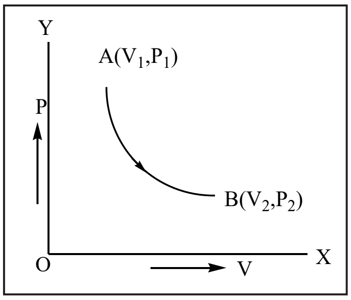
1. Isothermal Process is that which occurs at a constant temperature.
2. Adiabatic Process is that in which no heat enters or leaves the thermodynamic system during the change.
3. Isobaric Process is that which occurs at constant pressure.
4. Isochoric or Isovolumic Process is that which occurs at constant volume.

Summarises the four special thermodynamic processes.

S. No	Type of process	Feature
1.	Isothermal	Temperature constant
2.	Adiabatic	No heat flow between the system and the surroundings, i.e., $\Delta Q = 0$
3.	Isobaric	Pressure constant
4.	Isochoric	Volume constant

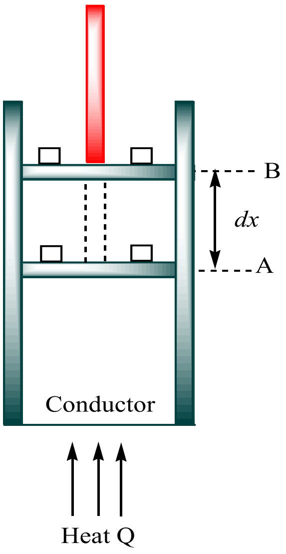
11.3.2 Indicator Diagram or P – V Diagram

The indicator diagram or P-V diagram represents the variation of volume V of a system with the pressure P of the system. Usually, we plot volume V along OX and pressure P along OY, as shown in Fig. The shape of PV diagram shall depend upon the nature of thermodynamical process, the system undergoes. In Fig, point A (V_1, P_1) represents initial state and the point B (V_2, P_2) represents final state of the system. The various points on the curve AB represent the intermediate states of the system in going from A to B.



11.3.3 First Law of Thermodynamics

Statement: "If the quantity of heat supplied to a system is capable of doing work, then the quantity of heat absorbed by the system is equal the sum of the increase in the internal energy of the system, and the external work done by it".



Consider some amount of gas enclosed in a cylinder having thermally insulated walls fixed with a frictionless movable piston made of thermally insulated material. But it has a conducting bottom.

Let 'A' be the initial position of the piston and ${}_1$ be the initial internal energy of the gas. Let an amount of heat 'dQ' be added to the gas through the conducting bottom

The total initial thermal energy of the gas = U_1+ dQ

The gas absorbs heat energy and expands the pushing the piston to a new position 'B'. The work done by the gas in displacing the piston through a small distance dx is dW.

The internal energy of the system increases to U_2 by absorbing the heat 'dQ'. Some of the internal energy is utilised in performing external work done (dW) in pushing the piston away. The internal energy changes to U_3.

The total energy after pushing = (U_3 + energy converted into work) j

According to law of conservation of energy

$U_1 + dQ = U_2 = U_3 + dW$

$\therefore \quad dQ = (U_3 - U_1) + dW$

$(U_3 - U_1)$ is the change in internal energy dU.

If an infinitesimally small amount of heat 'dQ' is added to the system (gas), corresponding change in internal energy is 'dU' and the work done by the gas is 'dW'.

Then dQ = dU + dW.

This is the mathematical statement of first law of thermodynamics.

Sign convention

If the heat energy dQ is added to the system (gas) it is taken as positive (+dQ).

If the heat is given out (released) by the system it is taken as negative (– dQ).

If the work is done by the system (gas) it is taken as positive (+ dW).

If the work is done on the system (gas) by surroundings it is taken as negative (–dW).

The increase in internal energy is taken as positive (+dU).

The decrease in internal energy is taken as negative (– dU).

Significance and limitations

1) This law introduces the concept of internal energy.

2) This law is applicable to all the three phases of matter (i.e., solid, liquid and gas).

3) This law is applicable to any process in nature.

4) This law does not give any information about the direction of heat flow, between the hot and cold body, when they are in thermal contact.

5) This law tells that it is impossible to get work from any machine without giving it an equivalent amount of heat energy.

NOTE:

- Heat and work are two distinct modes of energy transfer to a system that results in change in its internal energy, (a) Heat is energy transfer due to temperature difference between the system and the surroundings, (b) Work is energy transfer brought about by means (e.g. moving the piston by raising or lowering some weight connected to it) that do not involve such a temperature difference.
- The internal energy U of a system can change through two modes of energy transfer : heat and work.
- First Law of Thermodynamics is simply the general law of conservation of energy applied to any system in which the energy transfer from or to the surrounding is taken into account.
- U is a state variable, ΔU depends only on the initial and final states and not on the path taken by the gas to go from one to the other. However, ΔQ and ΔW will, in general, depend on the path taken to go from the initial to final states. Form the First Law of Thermodynamics, it is clear that the combination, is however $\Delta Q - \Delta W$, path independent.

11.3.4 APPLICATIONS OF THE FIRST LAW OF THERMODYNAMICS

(a) Isothermal Process:

In an isothermal process, temperature remains constant. If the system is an ideal gas, whose internal energy U depends only on temperature, the internal energy shall remain constant, i.e., $dU = 0$. The first law of thermodynamics implies ....

$$dQ = dU + dW = 0 + dW = dW,$$

i.e. $dQ = dW$

i.e., heat supplied in an isothermal process is used entirely to do work against the external surroundings.

(b) Adiabatic Process:

In an adiabatic process, no heat energy enters or leaves the system as it is well insulated, i.e., $dQ = 0$. The first law of therodynamics implies

$$dQ = dU + dW = 0 \text{ or } \quad dU = -dW$$

When a gas expands adiabatically, dW is positive. Therefore, dU must be negative, i.e., internal energy of the system would decrease and the gas will be cooled. The reverse is also true.

(c) Isochoric Process:

In an isochoric process, volume of a system (such as a gas) is kept constant, $dW = P\,(dV) = 0$. The first law of thermodynamics implies

$$dQ = dU + dW$$

$$dQ = dU$$

If heat is absorbed by a system at constant volume, its internal energy increases by the same amount and vice-versa.

(d) Isobaric Process:

In an isobaric process, pressure is kept constant. For example, when a liquid boils, its pressure and temperature remain constant till whole of the liquid is converted into vapours.

If V_i = initial volume of liquid

V_f = final volume of vapours

P = constant pressure at which liquid converts into vapours, then work done, $dW = P.dV = P\,(V_f - V_i)$

If m is mass of liquid and L is latent heat of vaporisation of liquid, then heat supplied for converting entire liquid into vapours, $dQ = mL$

From first law of thermodynamics

$$dQ = dU + dW$$

or $dU = dQ - dW = mL - P\,(V_f - V_i)$

This is the change in internal energy of the liquid on vaporisation.

(e) Cyclic Process:

In such processes, the system returns to its initial state after any number of changes.

In that case, no intrinsic property of the system would change. As internal energy of a system is a state variable and final state of the system is same as the initial state, therefore there is no change in internal energy of the system, i.e., $dU = 0$.

The first law of thermodynamics implies $dQ = dU + dW$

$dQ = dW \qquad (\because dU = 0)$

$\therefore$ Net work done during a cyclic process exactly equals the amount of heat energy transferred.

In a cyclic process, the system returns to its initial state after any number of changes. Therefore, P-V diagram of a cyclic process will be a closed loop of any arbitrary shape.

(f) Melting Process:

When a solid melts into a liquid, its internal energy increases, which can be calculated from first law of thermodynamics.

Let m = mass of a solid, L = latent heat of the solid.

Amount of heat absorbed during the melting process, $dQ = mL$

When a substance melts, the change in its volume (dV) is very small and therefore, it can be neglected.

$\therefore \quad dW = P(dV) = P \times 0 = 0$

According to first law of thermodynamics, $dQ = dU + dW$

$\therefore \quad mL = dU + 0$ or $dU = mL$

Hence, the internal energy during the melting process increases by (mL).

Note that increase in internal energy during melting process is at constant temperature. Therefore, it must be due to increase in internal potential energy (U_p). The internal kinetic energy (U_k) remains constant.

(g) Boiling process

Using first law of thermodynamics, we can calculate increase in internal energy of a liquid in the process of boiling. In boiling, a liquid changes into vapours at constant temperature and pressure. Suppose

m = mass of a liquid at its boiling point, V_1 = volume of the liquid,

P = constant pressure at which boiling process occurs,

V_2 = Volume of the same liquid in the vapour form under the same pressure,

L = Latent heat of vaporisation of the liquid.

$\therefore$ Work done in expansion, $dW = P(dV) = P(V_2 - V_1)$

Heat absorbed by the liquid in the boiling process, $dQ = mL$

According to first law of thermodynamics,

$$dQ = dU + dW$$

$$\therefore \quad dU = dQ - dW$$

$$dU = mL - P(V_2 - V_1)$$

Knowing m, L, P, V_1 and V_2, we can calculate dU, i.e., increase in internal energy of the liquid in the boiling process. This is again due to increase in internal potential energy.

(h) Free Expansions

The free expansions consist of adiabatic processes in which no transfer of heat occurs between the system and its surroundings and no work is done on or by the system, i.e., $dQ = 0$ and $dW = 0$

According to first law of thermodynamics, change in internal energy of the system

$$dU = dQ - dW = 0$$

Note that free expansions cannot be eluted slowly and in a controlled way. Therefore, we cannot plot P – V diagram for such processes.

11.3.5 SPECIFIC HEAT OF A GAS AT CONSTANT PRESSURE : C_P

At constant pressure, the quantity of heat necessary to increase the temperature of unit mass of a gas through one degree is called specific heat of the gas at constant pressure.

If 'ΔQ' is the quantity of heat required to rise the temperature of "m" mass of gas through 'Δt' at constant pressure.

$$C_p = \frac{\Delta Q}{m\Delta t}$$ CGS unit : Cal / g °C

SI unit: J $kg^{-1}K^{-1}$

Molar specific heat of a gas at constant pressure: C_P

At constant pressure, the quantity of heat necessary to increase the temperature of one mole of a gas through one degree is catted molar specific heat or molar heat capacity of the gas at constant pressure, C_p)

Let 'ΔQ' is the quantity of heat required to rise the temperature of 'n' moles of gas through 'Δt' at constant pressure. Then

$$C_p = \frac{\Delta Q}{n\Delta t}$$

If 'm' is the mass of gas considered and 'M' is its molecular weight then n = m/M.

$$C_p = \frac{M\Delta Q}{m\Delta t} = M\,c_p$$

C.G.S unit : Cal / mole °C

SI unit: J $mole^{-1}$ K^{-1}

11.3.6 SPECIFIC HEAT OF A GAS AT CONSTANT VOLUME : C_V

At Constant volume, the quantity of heat necessary to increase the temperature of unit mass of a gas through one degree is called the specific heat of the gas at constant volume. It is denoted by c_v. If 'ΔQ' is the quantity of heat required to rise the temperature of 'm' mass of gas at constant volume through 'Δt' then

$$c_v = \frac{\Delta Q}{m\Delta t}$$

C.G.S unit : Cal / g°C

SI unit: J kg^{-1} K^{-1}

Molar specific heat of a gas at constant volume: C_V.

At constant volume, the quantity of heat necessary to increase the temperature of one mole of a gas through one degree is called "Molar specific heat" or molar heat capacity of the gas at constant volume C_V

If 'ΔQ' is the quantity of heat required to rise the temperature of 'n' moles of a gas through 'At' at constant volume.

$$C_V = \frac{\Delta Q}{n\Delta T}$$

If 'm' is the mass of the gas and 'M' is its molecular weight then

$$n = \frac{m}{M} \text{ then } C_V = \frac{M\Delta Q}{m\Delta T} = M\,c_V$$

C.G.S unit : Cal / mole °C

SI unit: J $mole^{-1}$ K^{-1}

11.3.7 LIMITS OF SPECIFIC HEAT OF A GAS

Consider mass 'm' of a gas. Let ΔQ units of heat raise the temperature of the gas through ΔT. Then the specific heat of the gas is given by

$$c = \frac{\Delta Q}{n\Delta T} \qquad [\because \Delta Q = mc\,\Delta T]$$

Consider a gas enclosed in a cylinder fitted with an air-tight and frictionless piston.

i) Let the gas be suddenly compressed. In this case, no heat is supplied to the gas. But there is an increase in the temperature of the gas.

$$c = \frac{\Delta Q}{n\Delta T} \qquad [\because \Delta Q = 0]$$

ii) Let the gas be heated and allowed to expand. Suppose the ‘fall in temperature due to expansion' is equal to the ‘rise in temperature due to supplied’.

$$\therefore \quad c = \frac{\Delta Q}{m\times 0} = \infty \qquad [\because \Delta T = 0]$$

iii) Let the gas be heated and allowed to expand suppose, in this case, the ‘fall in temperature due to expansion’ is less than the ‘rise in temperature due to heat supplied’. The net effect will be a rise in temperature of the gas.

So, ΔT is positive. Thus $c = \left(\frac{\Delta Q}{m\Delta T}\right)$ is positive.

iv) Let the gas be heated and allowed to expand such that the ‘fall in temperature due to expansion is more than the ‘rise in temperature due to heat supplied’. The net effect will be a decrease in the temperature of the gas.

So ΔT is negative. Thus, $c = \left(\frac{\Delta Q}{m\Delta T}\right)$ is negative.

11.3.8. C_P IS GREATER THAN C_V

When a gas is heated at constant volume, no heat is spent in the expansion of the gas. The whole amount of heat is used to increase the temperature of the gas; increasing the internal energy.

If $(dQ)_v$ is heat supplied to n moles of a gas at Mont volume.

dU_1 is the increase in internal energy

dT_1 is the rise in temperature

n is the number of moles

C_v is the molar specific heat at constant volume

Then $(dQ)_v - dU_1 + 0 = n\,C_v\,dT_1$ ....(1)

When a gas is heated at constant pressure, the temperature of the gas increases and the volume also increases due to expansion.

Therefore a part of the heat supplied, to the gas is utilised to increase the temperature, increasing the internal energy and the other part is utilised in performing work against the external pressure during its expansion (dW)

If $(dQ)_p$ is the heat supplied to the gas at constant pressure

dU_2 is increase in internal energy

dT_2 is the rise in temperature

dW is the work done during the expansion

n is the number of moles

C_p is the molar specific heat at constant pressure

$(dQ)_p = n\,C_p dT_2 = dU_2 + dW$ ...(2)

If the rise in temperature is same $dT_1 = dT_2 = dT$

so internal energies $dU_1 = dU_2$

From equations 1 and 2

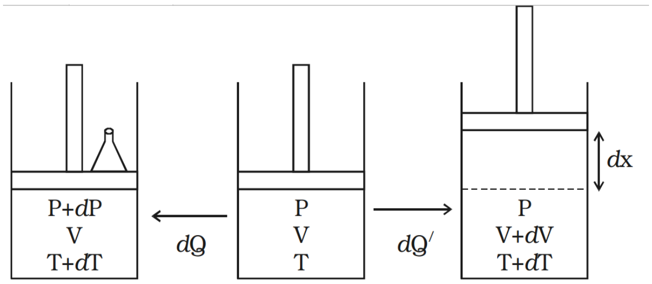
$nC_p dT = nC_v\,dT + dW$

$$C_p = C_v + \frac{dW}{ndT} \qquad [\therefore C_p > C_v]$$

Relation between C_p and C_v (Meyer's relation)

Let us consider one mole of an ideal gas enclosed in a cylinder provided with a frictionless piston of area A. Let P, V and T be the pressure, volume and absolute temperature of gas respectively (Fig.).

A quantity of heat dQ is supplied to the gas. To keep the volume of the gas constant, a small weight is placed over the piston. The pressure and the temperature of the gas increase to P + dP and T + dT respectively. This heat energy dQ is used to increase the internal energy dU of the gas. But the gas does not do any work (dW = 0).



Meyer's relation

$\therefore \quad dQ = dU = 1 \times C_v \times dT$ ... (1)

The additional weight is now removed from the piston. The piston now moves upwards through a distance dx, such that the pressure of the enclosed gas is equal to the atmospheric pressure P. The temperature of the gas decreases due to the expansion of the gas.

Now a quantity of heat dQ' is supplied to the gas till its temperature becomes T + dT. This heat energy is not only used to increase the internal energy dU of the gas but also to do external work dW in moving the piston upwards.

$\therefore \quad dQ' = dU + dW$

Since the expansion takes place at constant pressure,

$dQ' = C_p dT$

$\therefore \quad C_p dT = C_v dT + dW$ ... (2)

Work done, dW = force × distance

$= P \times A \times dx$

dW = P dV (since A × dx = dV, change in volume)

$\therefore \quad C_p dT = C_v dT + P\,dV$ ... (3)

The equation of state of an ideal gas is

$PV = RT$

Differentiating both the sides

$PdV = RdT$ ... (4)

Substituting equation (4) in (3),

$C_p dT = C_v dT + RdT$

$C_p = C_v + R$

$\therefore \quad C_p - C_v = R$

This equation is known as Meyer's relation

11.4.1 SPECIFIC HEAT OF GASES IN TERMS OF UNIVERSAL GAS CONSTANT R

Temperature is the measure of average kinetic energy of the molecules of a gas. The kinetic energy of a molecule is the sum of translational kinetic energy, vibrational kinetic energy and rotational kinetic energy. According to the law of equipartition of energy, the total energy is divided equally among the degrees of freedom of a molecule.

The average kinetic energy per molecule per degree $= \frac{1}{2}$KT.

where K is the Boltzmann's constant T is the average temperature.

Consider one mole of gas containing N_A number (Avogadro number) of molecules.

The average kinetic energy per one mole per degree of freedom $= \frac{1}{2}$RT.

(i) For a mono – atomic gas the degree freedom is 3.

∴ The internal energy $U = \frac{3}{2}RT$

∴ The change in internal energy

∴ $dU = \frac{3}{2}RdT$ ...(1)

Let a small quantity of heat dQ is transferred to the gas at constant volume so that only the internal energy is increased by dU. Hence dQ = dU.

Let the corresponding rise in temperature be dT.

∴ Molar specific heat at constant volume

$$C_V = \frac{dQ}{1 \times dT} = \frac{dU}{dT}$$

from equation(1)

∴ $C_V = \frac{dU}{dT} = \frac{3}{2}\frac{RdT}{dT} = \frac{3}{2}R$ ...(2)

But

∴ $C_P - C_V = R$

∴ $C_P = R + C_V$

∴ $C_P = R + \frac{3}{2}R = \frac{5}{2}R$ ...(3)

For a monoatomic gas =

$$\gamma = \frac{C_P}{C_V} = \frac{\frac{5}{2}R}{\frac{3}{2}R} = \frac{5}{3} = 1.67$$

(ii) For a diatomic gas the degrees of freedom is 5

Hence the internal energy for n mole of gas

$$U = \frac{5}{2}nRT$$

∴ The change in internal energy

$$dU = \frac{5}{2}RdT \qquad ...(4)$$

$$C_V = \frac{dU}{ndT} = \frac{\frac{5}{2}nRdT}{ndT} = \frac{5}{2}R$$

But $C_P = C_V + R$

∴ $C_P = \frac{5}{2}R + R = \frac{7R}{2}$ ...(6)

For a diatomic gas =

$$\gamma = \frac{C_P}{C_V} = \frac{\frac{7}{2}R}{\frac{5}{2}R} = \frac{7}{5} = 1.4$$

(iii) For a triatomic gas the degree of freedom may be 6 or 7.

∴ The internal energy

$$U = \frac{6nRT}{2} = 3nRT \Rightarrow dU = 3nRdT$$

$$CV = \frac{dU}{nRT} = \frac{3nRdT}{ndT} = 3R$$

& $C_P = 3R + R = 4R$

For a triatomic or polyatomic gas

$$\gamma = \frac{C_P}{C_V} = \frac{4R}{3R} = \frac{4}{3} = 1.33$$

EXPRESSION OF C_P AND C_V IN TERMS OF ADIABATIC CONSTANT γ

Consider the relation $C_P - C_V = R \rightarrow (1)$

$\frac{C_P}{C_V} = \gamma$ ∴ $C_P = \gamma C_V \rightarrow (2)$

From (1) & (2) $\gamma C_V - C_V = R$

∴ $C_V = \frac{R}{\gamma - 1} \rightarrow (3)$

from (2) and (3) $C_P = \frac{\gamma R}{\gamma - 1}$

ISOTHERMAL OPERATION/CHANGE

A change in pressure and volume of a gas without any change in its temperature, is called an isothermal change. In such a change, there is a free exchange of heat between the gas and its surroundings.

When a gas in an enclosure is compressed, work is done on the gas. Some heat is produced and its temperature tends to rise. If temperature of the gas is to remain constant, heat produced must be lost to the surroundings. Similarly, when the pressure on the gas is decreased, it expands, work is done by the gas, some heat energy is spent. Therefore, the temperature of the gas tends to fall. To keep the temperature of the gas constant, it must gain heat from its surroundings. Obviously, for an isothermal change, there should be a free exchange of heat between the gas and its surroundings.

Hence, the two essential conditions for a perfect isothermal change are :

(i) the walls of the container must be perfectly conducting, to allow free exchange of heat between the gas and its surroundings,

(ii) the process of compression or expansion should be slow so as to provide time for the exchange of heat between the gas and its surroundings.

As both these conditions are difficult to be realised strictly in practice, therefore perfect isothermal changes are practically impossible

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Some Examples of isothermal changes

1. Melting process is an isothermal change, because temperature of a substance remains constant during melting.
2. Similarly, boiling process is also an isothermal operation.
3. When a gas in a cylinder with perfectly conducting walls and perfectly conducting frictionless piston is compressed or allowed to expand so slowly that heat produced in compression or heat spent in expansion is exchanged with the surroundings, the changes are isothermal changes.

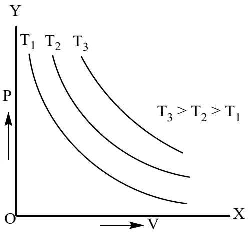
As in all isothermal operations, temperature is kept, constant, pressure (P) and volume (V) are related to each other by Boyle's Law; i.e.,

$$PV = \text{constant}^*$$

or

$$P_2V_2 = P_1V_1 \qquad ...(1)$$

This is known as the equation of an isothermal change.



If we plot a graph between V and P keeping temperature constant, we get a curve called an isothermal curve or an isotherm for the given mass of gas at a given temperature.

In Fig., isotherms for three different temperatures T_1, T_2, and T_3 are shown. From Fig, we find that the isotherms move away from the origin at higher temperatures.

NOTE:

- In an isothermal change, T = constant; $\Delta T = 0$

  As $\quad c = \dfrac{\Delta Q}{m(\Delta T)}$

  therefore, $c = \infty$

  i.e., during an isothermal change, specific heat of gas is infinite.

- For an ideal gas in an isothermal process, there is no change in internal energy of the gas, i.e., $dU = 0$. This is because T = constant, i.e., $dT = 0$.

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WORK DONE IN AN ISOTHERMAL EXPANSION

Consider one gram mole of an ideal gas enclosed in a cylinder with perfectly conducting walls and fitted with a perfectly frictionless and conducting piston, Fig. Let P_1,V_1,T be the initial pressure, volume and temperature of the gas. Let the gas expand to a volume V_2, when pressure reduces to P_2. The temperature remains constant at T as equivalent amount of heat spent in expansion is gained from the surroundings. At any instant during expansion, let the pressure of the gas be P.

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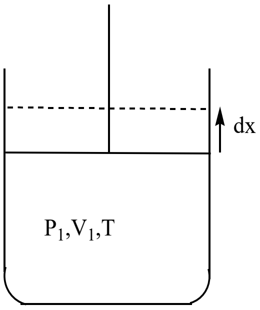
*From standard gas equation, PV = RT. As T is constant, therefore, PV = constant

If A is area of cross section of the piston, then force exerted by the gas on the piston is F = P x A.



If we assume that pressure of the gas during an infinitestimally small outward displacement dx of the piston remains constant, then small amount of work done during expansion

$$dW = F \times dx = P \times A \times dx$$

dW = P (dV) where dV = A (dx) = small increase in volume of the gas.

Total work done by the gas in expansion from initial volume V_1 to final volume V_2 is $W = \int_{V_1}^{V_2} P(dV)$

From standard gas equation,

$$PV = RT \qquad \text{or} \qquad P = RT/V$$

$$\therefore \quad W = \int_{V_1}^{V_2} \frac{RT}{V} dV = RT \int_{V_1}^{V_2} \frac{1}{V} dV$$

$$= RT[\log_e V]_{V_1}^{V_2}$$

$$W = RT[\log_e V_2 - \log_e V_1] = RT \log_e \frac{V_2}{V_1}.$$

$$W = 2.3026\ RT \log_{10} \frac{V_2}{V_1}$$

As $P_1V_1 = P_2V_2 \qquad \therefore \qquad \frac{V_2}{V_1} = \frac{P_1}{P_2}$

$$\therefore \quad W = 2.3026\ RT \log_{10} \frac{P_1}{P_2}$$

Amount of heat spent in isothermal expansion (in calories) is

$$H = \frac{W}{J} = 2.3026 \frac{RT}{J} \log_{10}\left(\frac{V_2}{V_1}\right)$$

This much heat must have been gained by the gas from the surroundings. THat is why temperature of the gas remains constant at T.

11.4.3 Adiabatic operation/change

A change in pressure and volume of a gas when no heat is allowed to enter into or escape from the gas is called an adiabatic change. Obviously, temperature of gas will also change in adiabatic operations.

When a gas in an enclosure is compressed suddenly, some heat is produced. If this heat is not allowed to escape, the temperature of the gas would rise. Further, if the gas is allowed to expand suddenly, some of heat energy is spent in increasing the volume of the gas. If this heat is not compensated from the surroundings, the temperature of the gas would fall.

Obviously, for an adiabatic change, the exchange of heat between the gas and the surroundings must be prevented.

Hence, the two essential conditions for a perfect adiabatic change are :

(i) the walls of the container must be perfectly non-conducting in order to prevent any exchange of heat between the gas and its surroundings, and

(ii) the process of compression or expansion should be sudden so that there is no time for the exchange of heat.

Both these conditions are difficult to be realised strictly in practice. Therefore, perfect adiabatic changes are practically impossible.

Some examples of nearly perfect adiabatic changes

1. Sudden compression or expansion of a gas in a container with perfectly non conducting walls.
2. Sudden bursting of the tube of a bicycle tyre,
3. Propagation of sound waves in air and other gases,
4. Expansion of steam in the cylinder of a steam engine,
5. Expansion of hot gases in an internal combustion engine.

A curve which represents relative variations of pressure and volume of a given mass of a gas for an adiabatic change is called an adiabatic curve. The shape of adiabatic curves is similar to the shape of isothermal curves, Fig.

The adiabatic relation between P and V for a perfect gas is

$$PV^{\gamma} = K, \text{ a constant} \quad ...(2)$$

where, $\gamma = \dfrac{\text{specific heat of the gas at constant pressure}}{\text{specific heat of the gas at constant volume}}$

Equation (2) can be rewritten as

$$P_1V_1^{\gamma} = P_2V_2^{\gamma}$$

From standard gas equation

$$PV = RT \qquad \text{or} \qquad P = RT/V$$

Putting in (2), we get

$$\frac{RT}{V}V^{\gamma} = K$$

$$T\,V^{(\gamma-1)} = \frac{K}{R} \text{ another constant.}$$

$\therefore \quad T\,V^{(\gamma-1)} = \text{constant} \quad ...(4)$

or $\quad T_1V_1^{(\gamma-1)} = T_2V_2^{(\gamma-1)} \quad ...(5)$

This is the adiabatic relation between P and T for a perfect gas. Again, from standard gas equation V = RT/P

Putting in (2), we get

$$P\frac{R^{\gamma}T^{\gamma}}{P^{\gamma}} = K$$

$$P^{(\gamma-1)}T^{\gamma} = \frac{K}{R^{\gamma}} \text{ another constant.}$$

i.e., $\quad P^{(\gamma-1)}T^{\gamma} = \text{constant.}$

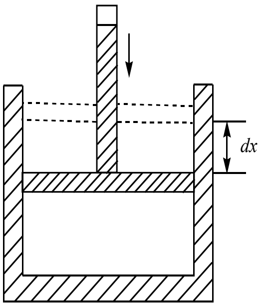
or $\quad P_1^{(\gamma-1)}T_1^{\gamma} = P_2^{(\gamma-1)}T_2^{\gamma}$

This is the adiabatic relation between P and T for a perfect gas.

11.4.4 RELATIONS BETWEEN P, V AND T IN AN ADIABATIC PROCESS:

(i) Relation between P and V

Consider one mole of an ideal gas contained in a thermally insulated cylinder fitted with a frictionless, movable and thermally insulated piston.



Let P, V and T be the pressure, volume and temperature at any instant.

Let the gas in the cylinder be compressed adiabatically so that the piston moves inwards through an infinitesimally small distance "dx".

If "A" is the area of the piston, the force acting on the piston F = PA.

Work done $\quad dW = \text{Force X distance}$

$$dW = P.A.dx = P.dV \quad ...(1)$$

where dV is decrease in volume = A.dx

The heat generated due to compression causes a rise in temperature dT. This heat energy is equal to $C_v dT$, where C_v is molar specific heat of the gas at constant volume. THis is equal to the change in internal energy dU.

$$dU = C_v.dT \quad ....(2)$$

In an adiabatic process, there is no exchange of heat.

$\therefore \quad dQ = 0 \quad ....(3)$

Accdring to first law of thermodynamics

$$dQ = dU + dW$$

Substituting the values from (1), (2), & (3) in above equation

$$O = C_v.dT + P,dV \quad .....(4)$$

From the ideal gas equation, we have PV = RT. On differentiating both sides

$$P.dV + V.dP = R.dT \qquad dT = \frac{PdV + VdP}{R}$$

Substituting the above value in equation (4), we have

$$C_v\left(\frac{PdV + VdP}{R}\right) + PdV = 0$$

$$C_v(PdV + VdP) + RPdV = 0$$

$$C_v(PdV + VdP) + (C_p - C_v)PdV = 0$$

$$(\text{But } C_p - C_v = R)$$

$$C_vPdV + C_vVdP + C_pPdV - C_vPdV = 0$$

dividing the above equation throughout by (C_vPV), we get

$$\frac{dP}{P} + \frac{C_p}{C_v}\frac{dV}{V} = 0$$

$$\frac{dP}{P} + \gamma\frac{dV}{V} = 0 \qquad \left[\because \gamma = \frac{C_p}{C_v}\right]$$

Let the initial pressure and volume br P_i and V_i respectively and the final pressure and volume be P_f and V_f respectively. Then on integration with in the limits,

$$\int_{V_i}^{V_f} \gamma \frac{dV}{V} = -\int_{P_i}^{P_f} \frac{dP}{P}$$

$$\gamma[\log_e V]_{V_f}^{V_f} = [\log_e P]_{P_i}^{P_f}$$

$$\gamma \log_e\left(\frac{V_f}{V_i}\right) = \log_e\left(\frac{P_f}{P_i}\right)$$

$$\log_e\left(\frac{V_f}{\frac{P_f}{P_i}\frac{P_f}{P_i}\frac{P_f}{P_i}}\right)^{\gamma} = \log_e\left(\frac{P_f}{P_i}\right)$$

$$\therefore \frac{V_f^{\gamma}}{V_i^{\gamma}} = \frac{P_i}{P_f} \qquad \therefore P_i V_i^{\gamma} = P_f V_f^{\gamma}$$

$$PV^{\gamma} = \text{Constant.} \qquad ...(5)$$

This is the relation between P and V during an adiabatic process.

(ii) Relation between V and T

From the gas equation PV = RT and P = RT/v

Substituting this value in equation

$$PV^{\gamma} = \text{Constant}$$

$$\therefore \quad \left(\frac{RT}{V}\right)V^{\gamma} = \text{Constant}$$

$$RTV^{\gamma-1} = \text{Constant}$$

$$\therefore \quad TV^{\gamma-1} = \text{Constant}$$

This is the relation between V and T during and adiabatic process.

Work done in an adiabatic change

Consider one mole of an ideal gas contained in a cylinder having non conducting walls and provided with a non conducting and frictionless piston. Let the pressure, and volume of the gas be P and V at any instant. Suppose the piston moves through an infinitesimally small distance 'dx' at constant pressure 'P'.

The work done dW = P.A.dx = P.dV

Where 'A' is the area of the piston and dV is the infinitesimally small change in volume.

The system changes from the initial state (P_1, V_1) to the final state (P_2, V_2).The temperature changes from T_1 to T_2.

The work done during the change

$$W = \int dW = \int_{V_1}^{V_2} P.dV$$

For an adiabatic change

$$PV^{\gamma} = k \text{ (a constant)} \qquad P = \frac{K}{V^{\gamma}} = KV^{-\gamma}$$

$$\therefore \quad W_{adia} = \int_{V_1}^{V_2} KV^{\gamma}\, dv$$

$$W_{adia} = \int_{V_1}^{V_2} V^{\gamma}\, dv = K = \int_{V_1}^{V_2} KV^{\gamma}\, dv$$

$$W_{adia} = \frac{K}{1-\gamma}[V_2^{1-\gamma} - V_1^{1-\gamma}]$$

$$W_{adia} = \frac{1}{1-\gamma}[KV_2^{1-\gamma} - KV_1^{1-\gamma}]$$

$$\text{But} \quad P_1V_1^{\gamma} = P_2V_2^{\gamma}$$

$$\therefore \quad W_{adia} = \frac{1}{1-\gamma}[P_2V_2^{\gamma}V_2^{1-\gamma} = P_1V_1^{\gamma}V_1^{1-\gamma}]$$

$$W_{adia} = \frac{1}{1-\gamma}[P_2V_2 - P_1V_1]$$

But from gas equation

$$P_1V_1 = RT_1 \text{ and } P_2V_2 = RT_2$$

$$\therefore \quad W_{adia} = \frac{1}{1-\gamma}(RT_2 - RT_1)$$

$$W_{adia} = \frac{R}{1-\gamma}[T_2 - T_1]$$

$$\text{or} \quad W_{adia} = \frac{R}{\gamma-1}[T_1 - T_2]$$

This expression gives the work done for a 1 mole of gas.

For 'n' moles of gas

$$W_{adia} = \frac{nR}{\gamma-1}(T_1 - T_2)$$

If 'm' is the mass of gas considered and M is its molecular weight then

$$W_{adia} = \frac{mR}{M(\gamma-1)}(T_1 - T_2)$$

Work done if 1 gram of gas is considered

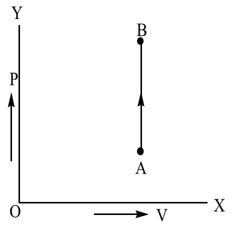
$$W_{adia} = \frac{R}{M(\gamma-1)}(T_1 - T_2)$$

But R/M = r, the specific gas constant.

$$\therefore \quad W_{adia} = \frac{r}{\gamma-1}(T_1 - T_2)$$

11.4.5 ISOCHORIC AND ISOBARIC PROCESSES:

In an isochoric process, volume V is constant. No work is done on or by the gas. Heat absorbed by the gas goes entirely to change its internal energy and its temperature. The change in temperature for a given amount of heat is determined by specific heat of the gas at constant volume.



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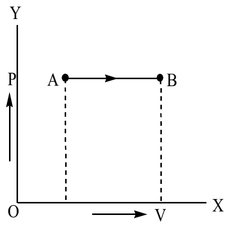


ⓑ

The pressure-volume plot for an isochoric process is a vertical straight line parallel to pressure axis as shown in Fig(a). No work is done one or by the gas (i.e., W = 0). The area under the vertical line is zero. Heat absorbed changes directly the internal energy or temperature of the gas.

For an isobaric process, pressure-volume plot is a horizontal straight line parallel to volume axis as shown in Fig. (b).

In an isobaric process, pressure P is fixed. Work done by the gas is

$$W = P(V_2 - V_1) = \mu R(T_2 - T_1)$$
$$= \text{area ABCDA}$$

As temperature changes, so does the internal energy of the gas.

The heat absorbed increases partially the internal energy and is spent partially in doing work. The change in temperature for a given amount of heat is determined by the specific heat of the gas at constant pressure.

11.4.6 Limitations of the First Law of Thermodynamics

The first law of thermodynamics establishes the essential equivalence between the heat energy and mechanical work and says that the two can be converted into eachother. Further, 4.18 joule of mechanical work are required o produce one calorie of heat and vice-versa. However, this law has the following limitations :

1. The first law does not indicate the direction in which the change can occur

For example (i) when two bodies at different temperatures are put in thermal contact with each other, heat flows from the body at higher temperature to the body at lower temperature. We now know that heat cannot flow from the body at lower temperature to the body at higher temperature, although first law of thermodynamics is not violated, (ii) When a moving car is stopped by applying brakes, work done against friction is converted into heat. When the car cools down, it does not start moving with the conversion of all its heat energy into mechanical work, (iii) When a bullet strikes a target, kinetic energy of the bullet is converted into heat energy. But heat energy developed in the target cannot be converted back into mechanical energy of the bullet enabling it to fly back.

2. The first law gives no idea about the extent of change

Our observations and experience tell that there appears to be no restriction on conversion of mechanical work into heat. But there are severe restrictions on the reverse process, i.e., conversion of heat energy into mechanical energy.

We know that heat is not converted into mechanical energy all by itself. An external agency called the heat engine is required for the purpose.

No heat engine can convert all the heat energy received from the source into mechanical energy. The first law of thermodynamics is silent about all this.

3. The first law of thermodynamics gives no information about the source of heat, i.e., whether it is a hot or a cold body.

11.5.1 Cyclic and Non-cyclic Processes

A cyclic process consists of a series of changes which return the system back to its initial state.

For example, suppose a system consists of a gas at pressure P, volume V and temperature T. Let the system be subjected to a number of changes including isothermal and adiabatic expansions and compressions. In the final state, if the system has the same pressure P, the same volume V and the same temperature T, the succession of changes involved are said to form a cyclic process.

When the system is restored to its initial state, no intrinsic property of the system – including its internal energy can change, i.e,, dU = 0. According to first law of thermodynamics, dQ = dU + dW = dW, i.e., net work done during the process = net amount of heat energy transferred. Internal energy of the system remains unchanged.

On the contrary, in a non-cyclic process, the series of changes involved do not return the system back to its initial state.

11.5.2 WORK DONE DURING A NON-CYCLIC PROCESS

Work done in a non-cyclic process is calculated in Art 8.10. It is equal to the area under the curve on the PV diagram.

Conventionally, the work done is taken as positive, when the gas is expanding, and it is taken as negative, when the gas is compressed.

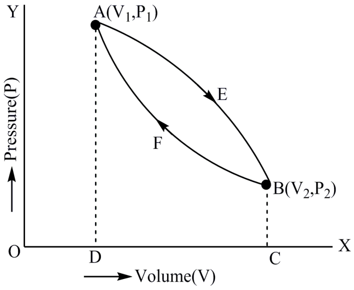
As the two points representing the initial and final state of a system can be joined by any number of j different curves, and the area enclosed by each curve is different, therefore in such cases, work done will be j different. Hence we conclude that the work done in a non cyclic process depends upon the path chosen or | the series of changes involved.

11.5.3 WORK DONE DURING A CYCLIC PROCESS

Let A (V_1, P_1) represent the initial state of a gas, Fig. Let it be subjected to a number of changes in volume and pressure such that it acquires-the final state B (V_2, P_2), via the path AEB.

∴ Work done by the gas in going from A to B, along AEB is W_1 = area AEBCDA.

Let the system be now subjected to another succession of changes in volume and pressure, which return the system back to the initial state A (V_1, P_1), along the path BFA. The total number of changes involved obviously constitute a cyclic process.



∴ Work done on the gas in going from B to A via the path BFA is $W_2 = -$ area BFADCB

Net work done by the gas in the cyclic process

$$W_{A\rightarrow B\rightarrow A} = W_1 + W_2 = \text{area AEBCDA} - \text{area BFADCB}$$
$$= \text{area AEBFA}$$

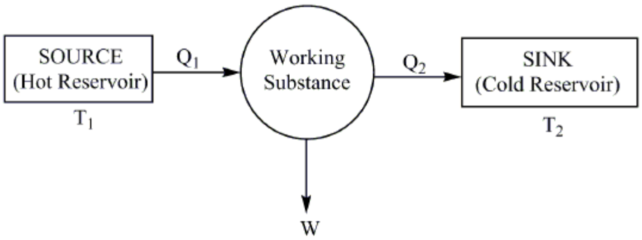
Hence, we conclude that work done per cycle is numerically equal to the area enclosed by the loop representing the cycle.

In Fig., cyclic path has been traced in the clockwise direction. The expansion curve AEB lies above the compression curve BFA. The area enclosed by the loop AEBFA comes out to be positive, indicating that net work is done by the system.

If cyclic path were traced in the anticlockwise direction, the expansion curve would be below the compression curve. The area of the loop would come out to be negative, indicating that net work will be done on the system.

11.5.4 HEAT ENGINE

A heat engine is a device which converts heat energy into mechanical energy. For this, a system is made to undergo a cyclic process.



A heat engine essentially consists of three parts :

1. Source of heat at higher temperature,
2. Working substance,
3. Sink of heat at lower temperature.

In a steam engine, steam acts as the working substance. In a diesel engine, a mixture of fuel vapours and air acts as the working substance. Heat is supplied to working substance at a relatively high input temperature from a place called the source/hot reservoir. It then undergoes a cyclic change consisting of several processes. It actually does the work. The remaining of input heat is ejected to the sink/cold reservoir.

The work done by the working substance of the system in a cycle, is transferred to the environment via some arrangement. For example, when the working substance is in a cylinder with moving piston, the piston transfers mechanical energy to the wheels of a vehicle via a shaft. The cycle is repeated again and again to obtain the desired useful work for some purpose.

Suppose, Q_1 = amount of heat absorbed by the working substance from the source at T_1K, in one complete cycle, Fig.

Q_2 = amount of heat rejected to the sink at T_2K, in the cycle.

W = net amount of external work done by the working substance, on the environment in the cycle.

∴ Net amount of heat absorbed in one cycle,

$$dQ = Q_1 - Q_2$$

As the working substance returns to its initial state, the change in its internal energy = 0 i.e.,dU = 0

According to first law of thermo – dynamics,

$$dQ = dU + dW \quad \therefore \quad dQ = dW$$

i.e., net amount of heat absorbed = external work done by the engine

or $Q_1 - Q_2 = W$

Thermal efficiency of a heat engine is defined as the ratio of net work done per cycle by the engine to the total amount of heat absorbed per cycle by the working substance from the source. It is denoted by η.

Thus,

$$\eta = \frac{\text{Network done/cycle}(W)}{\text{Total amount of heat absorbed/cycle}(Q_1)} = \frac{Q_1 - Q_2}{Q_1}$$

$$\eta = 1 - \frac{Q_2}{Q_1}$$

As some heat is always rejected to the sink, $Q_2 \neq 0$. Therefore, efficiency is always less than 1, i.e., thermal efficiency of a heat engine is always less than 100%.

For $Q_2 = 0$, $\eta = 1 = 100\%$, i.e., the engine will have 100% efficiency in converting heat into work. Thus, first law of thermodynamics does not rule out a heat engine with 100% efficiency. However, practical experience shows that efficiency is less than 100% even if we eliminate all sorts of losses associated with actual heat engines.

It turns out that there is a fundamental limit on the efficiency of a heat engine set by an independent principle of nature, called the second law of thermodynamics.

Types of heat engines

In practice, heat engines are of two types :

1. External combustion engine, in which heat is produced by burning the fuel in a chamber outside the main body (i.e., cylinder and piston arrangement) of the engine.

For example, steam engine is an external combustion engine.

The theoretical value of efficiency of steam engine at 20 atmosphere pressure and 250°C is about 32%. However, practically thermal efficiency of a steam engine varies from 12% to 16%.

2. Internal combustion engine, in which heat is produced by burning the fuel inside the main body of the engine. Petrol engine and diesel engine are internal combustion engines. They are also called four stroke engines.

The theoretical value of thermal efficiency of an internal combustion engine is about 52 %.

Practically, efficiency of a petrol engine is about 26 % and the efficiency of a diesel engine is about 40%.

As petrol engine is light in weight and small in size, it is used in aeroplanes, cars, scooters etc.

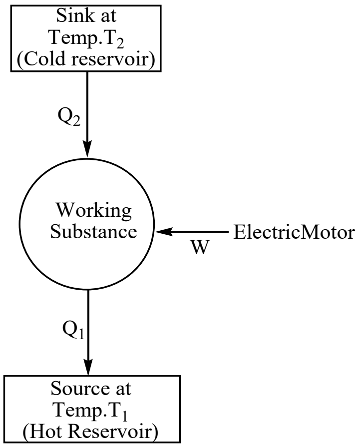
A diesel engine is heavy and hence it is used in heavy vehicles like bus, truck, tractor, railway engine, etc.

11.5.5 PRINCIPLE OF A REFRIGERATOR (OR HEAT PUMP)

A refrigerator, as we all know, is a device used far cooling things. It is also called a heat pump.

As is known, the natural tendency of heat is to flow from higher temperature to lower temperature. However, if energy is supplied, heat can be made to flow from lower to higher temperatures, against its natural tendency. Refrigerators, air conditioners and heat pumps do just that.

An ideal refrigerator can be regarded as a heat engine working in the reverse direction. Thus, in a refrigerator, the working substance would absorb a certain quantity of heat Q_2 from the sink at lower temperature T_2 and reject a larger amount of heat Q_1 to the source at higher temperature T_1 with the help of an external agency supplying the necessary energy to the system, Fig.



In a refrigerator, the working substance (usually, in gaseous form) goes through the following steps :

(a) Sudden expansion of the gas from high pressure to low pressure, resulting in cooling of the gas converting it into a vapour- liquid mixture.

(b) This mixture absorbs heat from the region to be cooled and gets converted into vapours.

(c) The vapours get heated on account of external work done on the system by the supply of electric power to the refrigerator.

(d) The heated vapours then release the heat to the surroundings, bringing them to the initial state after completing one cycle.

Note that water and other food stuff to be cooled in the refrigerator serve as sink at lower temperature. The source in this case is the atmosphere or surrounding air at room temperature T_1

In the cycle of changes, as the working substance returns to its initial state, therefore, there is no change in its internal energy, i.e., $dU = 0$.

If Q_2 = amount of heat extracted per cycle from the cold reservoir at lower temperature T_2, Fig.

W = Work done/cycle on the system-the refrigerant.

Q_1 = amount of heat released/cycle to the source (surrounding air at higher temperature T_1), then net amount of heat absorbed, $dQ = Q_2 - Q_1$

Work done on the system $dW = -W$

According to first law of thermodynamics,

$dQ = dU + dW$

$Q_2 - Q_1 = 0 - W$ or $W = Q_1 - Q_2$ (*)

Coefficient of performance of a refrigerator (β) is defined as the ratio of quantity of heat removed per cycle from the contents of the refrigerator (Q_2) to the energy spent per cycle (W) to remove this heat.

i.e.,

$$\beta = \frac{Q_2}{W}$$

using (*)

$$\beta = \frac{Q_2}{Q_1 - Q_2}$$

$$\beta = \frac{\frac{Q_2}{Q_1}}{\frac{Q_1}{Q_1} - \frac{Q_2}{Q_1}} = \frac{\frac{Q_2}{Q_1}}{1 - \frac{Q_2}{Q_1}}$$

In a Carnot cycle, $\frac{Q_2}{Q_1} = \frac{T_2}{T_1}$ $\quad \therefore \beta = \frac{\frac{T_2}{T_1}}{1 - \frac{T_2}{T_1}}$ ..(*1)

or

$$\beta = \frac{T_2}{T_1 - T_2} \quad .....(1)$$

This relation shows that to have large coefficient of performance of refrigerator, $(T_1 - T_2)$ should be small, i.e., the object to be cooled must have temperature close to the temperature of surroundings (/i.e., room temperature) before placing it in the refrigerator.

If η is efficiency of Carnot cycle, then

$$1-\frac{T_2}{T_1}=\eta \qquad \therefore 1-\eta=\frac{T_2}{T_1}$$

From (*1), $$\beta=\frac{1-\eta}{\eta}$$

1. Coefficient of performance of a refrigerator stands for efficiency of the refrigerator. Thus, higher the value of p, more efficient is the refrigerator.
2. Eqn.(1) shows that smaller is the value of (T_1-T_2), greater is the value of β, i.e., smaller is the difference in temperatures of atmosphere and food stuff to be cooled, more efficient will be the refrigerator.

As the refrigerator works, T_2 goes on decreasing due to formation of ice. T_1 is almost constant. Therefore, β decreases. When the refrigerator is defrosted, T_2 increases and hence β increases. Thus, defrosting is necessary for better working of the refrigerator.

3. A refrigerator cannot work without some external work done on the system, i.e., W is never zero. Hence, $\beta = Q_2/W$ cannot be infinite i.e., coefficient of performance of a refrigerator can never be infinite.

11.5.6 SECOND LAW OF THERMODYNAMICS

We know that first law of thermodynamics is the principle of conservation of energy. A critical study shows that there are many conceivable processes that are perfectly allowed by the first law and yet they are never observed in practice. For example, the first law does not prohibit heat flowing from a cold body to a hot body. Still, when a hot body is placed in contact with a cold body, hot body has never been observed to become hotter and the cold body colder. Heat flows spontaneously from hot body to cold body.

Again, first law of thermodynamics does not prohibit a heat engine to convert 100% of heat energy into mechanical energy with no waste heat. But Carnot found that even an ideal heat engine would generate some waste heat. The real engines would generate even more waste heat.

Clearly, some additional basic principle of nature forbids certain phenomena, even though the first law is not violated. This principle which disallows certain phenomena consistent with the First Law of thermodynamics is known as the Second law of Thermodynamics, According to second law of thermodynamics, heat flows spontaneously from a substance at higher temperature to another at lower temperature. Heat does not flow spontaneously in the reverse direction.

The other two important forms of second law of thermodynamics are

(a) Kelvin Planck Statement

It is impossible to construct a heat engine which would absorb heat from a reservoir and convert 100% of the heat absorbed into work.

The law implies that a perfect heat engine with 100% efficiency can never be constructed. Infact, presence of sink is essential for continuous conversion of heat into work. If temperature of source T_1 were equal to temperature of sink T_2, the efficiency of heat engine $\left(\eta=1-\frac{T_2}{T_1}\right)$ becomes zero.

(b) Clausius Statement

It is impossible to design a self acting machine unaided by any external agency, which would transfer heat from a body at a lower temperature to another body at a higher temperature.

We know that in a refrigerator, the working substance extracts heat from a colder body (sink) and rejects a larger amount of heat to a hotter body (source) with the help of an external agency, i.e., the electric supply to the refrigerator. No refrigerator can ever work without external supply of electric energy to it.

Both the statements of second law of thermodynamics are equivalent, as one can be obtained from the other.

Note that the statements of second law start from 'it is impossible'. Thus, the second law is negative in form. Hence, it cannot be proved directly. However, the law is accepted universally and has been applied to a wide variety of phenomena like transfer of heat, expansion of a gas, diffusion of gases, etc.

NOTE:

The second law of thermodynamics puts a fundamental limitation on

(i) efficiency of a heat engine,

(ii) coefficient of performance of a refrigerator.

The law emphasises that efficiency of a heat engine can never be 100%. It implies that heat released to the cold reservoir (/. e., sink) can never be made zero.

For refrigerator, the second law emphasises that its coefficient of performance can never be infinite. It implies that some external work has to be done to transfer heat from lower temperature to higher temperature.

11.5.7 REVERSIBLE AND IRREVERSIBLE PROCESSES

A thermodynamical process taking a system from initial state i to final state f is reversible, if the process can be turned back such that both, the system and the surroundings return to their original states, with no other change anywhere else in the universe.

For a process to be reversible, the following conditions must be satisfied :

1. The process should proceed at an extremely slow rate, i.e., process is quasi-static so that the system is in equilibrium with surroundings at every stage, i.e.,

(i) the system remains in mechanical equilibrium, i.e., there is no unbalanced force, (ii) the system remains in thermal equilibrium, i.e., all parts of the system and the surroundings remain at the same temperature, (iii) the system remains in chemical equilibrium, i.e., the internal structure of the system does not change.

2. The system should be free from dissipative forces like friction, inelasticity, viscosity, etc. This is because energy spent against such forces cannot be recovered.

As all the conditions mentioned above are of an idealized nature, no process in nature is truly reversible.

Infact, reversibility is an idealized concept which can never be attained. It can at best be approximated.

Some of the examples of approximately reversible processes are :

(i) An ideal gas allowed to expand slowly and then compressed slowly in a cylinder fitted with frictionless movable piston.

(ii) Electrolysis can be taken as a reversible process provided resistance offered by electrolyte is zero.

(iii) Slow compression and expansion of a spring can also be treated as a reversible process.

A process, which does not satisfy even one of the conditions for reversible process is called an irreversible process.

Infact, all spontaneous processes of nature are irreversible processes. For example, transfer of heat from a hot body to a cold body, ordinary expansion of a gas, diffusion of gases, stopping of a moving body through friction etc. are all irreversible processes.

Irreversibility arises mainly from two causes :

(i) Many processes like free expansion or an explosive chemical reaction take the system to non equilibrium states, (ii) most processes involve friction, viscosity and other dissipative effects.

As the dissipative effects are present everywhere, and they can be minimised only and cannot be fully eliminated, therefore, most processes we deal with are irreversible processes.

Some additional examples of irreversible processes are :

(i) Production of heat on passing an electric current through a wire. The same amount of heat is produced in a given time, even when the direction of current is reversed.

(ii) Most of the chemical reactions are irreversible, because they involve changes in the internal structure of the constituents.

(iii) The decay of organic matter is an irreversible process.

(iv) Rusting of iron is an irreversible process.

(v) Adiabatic (sudden) compression and expansion of a gas are irreversible processes.

(vi) Mixing of two different substances like water and alcohol.

(vii) The formation of solution of a solid in water etc.

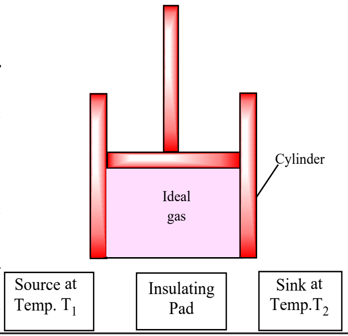
The significance of the concept of reversibility can be understood as follows :

The main concern of thermodynamics is the efficiency with which heat can be converted into mechanical work. The second law of thermodynamics ruled out the possibility of a perfect heat engine with 100% efficiency. It turned out that a heat engine based on idealised reversible processes achieves the highest possible efficiency. All other engines involving irreversibility of some kind have much lower efficiency. Therefore, reversibility is an important concept in thermodynamics.

11.5.8 CARNOT CYCLE

Sadi Carnot devised an ideal cycle of operation for a heat engine, in the year 1824. This cycle came to be known as Carnot cycle. The machine used for realising this ideal cycle of operation is called an ideal heat engine or Carnot heat engine.

Construction



The essential parts of an ideal heat engine or Carnot heat engine are shown in Fig.

(i) Source of heat. The source is maintained at a fixed higher temperature T_1, from which the working substance draws heat. The source is supposed to possess infinite thermal capacity and as such any amount of heat can be drawn from it without changing its temperature.

(ii) Sink of heat. The sink is maintained at a fixed lower temperature T_2, to which any amount of heat can be rejected by the working substance. It has also infinite thermal capacity and as such its temperature remains constant at T_2, even when any amount of heat is rejected to it.

(iii) Working substance. A perfect gas acts as the working substance. It is contained in a cylinder with non-conducting sides but having a perfectly conducting-base. This cylinder is fitted with perfectly non-conducting and frictionless piston.

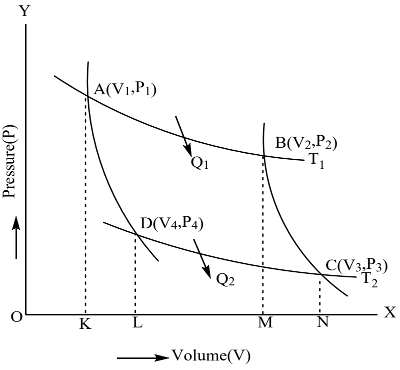
Apart from these essential parts, there is a perfectly insulating stand or pad on which the cylinder can be placed. It would isolate the working substance completely from the surroundings. Hence, the gas can undergo adiabatic changes.

Theory

The Carnot cycle consists of the following four stages :

1. Isothermal expansion,
2. Adiabatic expansion,
3. Isothermal compression,
4. Adiabatic compression.

The cycle is carried out with the help of the Carnot engine as detailed below:



Consider one gram mole of an ideal gas enclosed in the cylinder. Let V_1, P_1, T_1 be the initial volume, pressure and temperature of the gas. The initial state of the gas is represented by the point A on P – V diagram, Fig. We shall assume that all the four processes are quasi-staticand non dissipative, the two conditions for their reversibility. Now, the four processes are :

1. Isothermal expansion

Let the cylinder be placed on the source and the gas be allowed to expand by slow outward motion of the piston. Since it absorbs the requisite amount of heat from the source through the conducting base of the cylinder, therefore, temperature of the gas remains constant. This operation is called isothermal expansion and is represented by the isothermal curve AB. Let the amount of heat energy absorbed in the process be Q_1 and W_1 be the corresponding amount of work done by the gas in expanding isothermally from A (V_1, P_1) to B (V_2, P_2)

$$\therefore \quad Q_1 = W_1 = \int_{V_1}^{V_2} P.dV = RT_1 \log_e \frac{V_2^{\,*}}{V_1}$$

$$= \text{area ABMKA} \qquad \ldots(1)$$

NOTE:

In an ideal heat engine, all processes are reversible and there occurs absolutely no wastage of energy due to friction and turbulence, etc.

2. Adiabatic Expansion

The cylinder is now removed from the source and is placed on the perfectly insulating pad. The gas is allowed to expand further from B (V_2, P_2) to C (V_3, P_3). Since, the gas is thermally insulated from all sides, no heat can be gained from the surroundings. Temperature of gas falls to T_2, the expansion is adiabatic and is represented by the adiabatic curve BC. Let W_2 be the work done by the gas in expanding adiabatically from B (V_2, P_2) to C (V_3, P_3)

$\therefore$

$$W_2 = \int_{V_2}^{V_3} P.dV = \frac{R(T_2 - T_1)^{*}}{(1-\gamma)} = \text{area BCNMB} \qquad \ldots(2)$$

3. Isothermal Compression

The cylinder is now removed from the insulating pad and is placed on the sink at a temperature T_2. The piston is moved inwards slowly so that the gas is compressed until its pressure is P_4 and volume is V_4. This state is represented by the point D. Since, the base of cylinder is conducting, heat produced in compression will pass to the sink so that the temperature of the gas remains constant at T_2. This process is, therefore, isothermal and is represented by an isothermal curve, CD. Let Q_2 be the amount of heat energy rejected to the sink and W_3 be the amount of work done on the gas in compressing it isothermally from a state C (V_3, P_3) to D (V_4, P_4).

$$\therefore Q_2 = W_3 = \int_{V_3}^{V_4} -P.dV = -RT_2 \log_e \frac{V_4}{V_3} = RT_2 \log_e \frac{V_3}{V_4}$$

$$= -\text{ area CDLNC} \qquad \ldots(3)$$

4. Adiabatic Compression

The cylinder is again placed on the insulating pad. The piston is further moved downwards so that the gas is compressed to its initial volume V_1 and pressure P_1. As the gas is insulated from all sides, heat produced raises the temperature of the gas to T_1. This change is adiabatic and is represented by the adiabatic curve DA. Let W_4 be the work done on the gas in compressing it adiabatically from a state D (V_4, P_4) to the initial state A (V_1, P_1).

$$\therefore \quad W_4 = \int_{V_4}^{V_1} -P.dV = \frac{-R(T_2 - T_1)^{*}}{(1-\gamma)}$$

$$= -\text{ area DAKLD} \qquad \ldots.(4)$$

Work done by the engine per cycle:

Total work done by the gas (in steps 1 and 2) = $W_1 + W_2$

Total work done on the gas (in steps 3 and 4) = $W_3 + W_4$

$\therefore$ Net work (W) done by the gas in a complete cycle, $W = W_1 + W_2 - (W_3 + W_4)$ ....(5)

From, $W_2 = W_4$, in magnitude,** $\therefore W = W_1 - W_3$

$$W = Q_1 - Q_2 \qquad \ldots(6)$$

In terms of area, we can rewrite as

W = area ABMKA + area BCNMB – area CDLNC – area DAKLD

$$W = \text{area ABCDA} \qquad \ldots.(7)$$

Hence, in Carnot heat engine, net work done by the gas per cycle is numerically equal to the area of the loop representing the cycle.

EFFICIENCY OF CARNOT ENGINE

Efficiency of Carnot engine is defined as the ratio of net mechanical work done per cycle by the gas (W) to the amount of heat energy absorbed per cycle from the source (Q_1)

i.e., $\eta = \frac{W}{Q_1}$ Using, (6)

$$\eta = \frac{Q_1 - Q_2}{Q_1} = 1 - \frac{Q_2}{Q_1} \quad ....(8)$$

Referring to Fig., as A (V_1, P_1) and B (V_2, P_2) lie on the same isothermal,

$$P_1V_1 = P_2V_2 \quad ....(9)$$

As B (V_2, P_2) and C (V_3, P_3) lie on the same adiabatic,

$$\therefore \quad P_2V_2^{\gamma} = P_3V_3^{\gamma} \quad ....(10)$$

Again, C (V_3, P_3) and D (V_4, P_4) lie on the same isothermal,

$$\therefore \quad P_3V_3 = P_4V_4 \quad ....(10)$$

Finally, D (V_4, P_4) and A (V_1,P_1) lie on the same adiabatic,

$$\therefore \quad P_4V_4^{\gamma} = P_1V_1^{\gamma} \quad ....(11)$$

Multiplying (8), (9), (10) and (11), we get

$$(P_1V_1)(P_2V_2^{\gamma})(P_3V_3)(P_4V_4^{\gamma}) = (P_2V_2)(P_3V_3^{\gamma})(P_4V_4)(P_1V_1^{\gamma})$$

$$V_1^{\gamma}V_2^{\gamma}V_3^{\gamma}V_4^{\gamma} = V_1^{\gamma}V_2V_3^{\gamma}V_4$$

$$V_2^{\gamma-1}V_4^{\gamma-1} = V_1^{\gamma-1}V_3^{\gamma-1} \text{ or } (V_2V_4)^{\gamma-1} = (V_1V_3)^{\gamma-1}$$

or $V_2V_4 = V_1V_3$, i.e., $\frac{V_2}{V_1} = \frac{V_3}{V_4}$

os that $\log\frac{V_2}{V_1} = \log\frac{V_3}{V_4}$ ...(12)

Dividing(3) by (1), we get,

$$\frac{Q_2}{Q_1} = \frac{RT_2\log_e V_3/V_4}{RT_1\log_e V_2/V_1} = \frac{T_2}{T_1} \text{ i.e., } \frac{Q_2}{Q_1} = \frac{T_2}{T_1} \quad ...(13)$$

Putting this value in (7), we get,

$$\eta = 1 - \frac{T_2}{T_1} \quad ....(14)$$

1. Eqn(14) reveals that efficiency of Carnot heat engine depends on temperature of source T_1, and temperature of sink T_2. But the efficiency does not depend upon the nature of the working substance.

2. As RHS of eqn.(14) is less than 1, therefore efficiency of Carnot heat engine is less than 100%.

3. For $\eta = 1$ (i.e., 100 %), either $T_1 = \infty$ or $T_2 = 0$ K.

As source at infinite temperature or sink at 0K are not attainable, Carnot heat engine cannot have 100% efficiency.

4. If $T_2 = T_1$ then from (14), $\eta = 1 - \frac{T_1}{T_1} = 1 - 1 = 0$

i.e,, the Carnot engine shall not work. Heme, it is not possible to convert heat energy into mechanical work unless source and sink of heat are at different temperatures.

5. From eqn.(13), we find that

$$\frac{Q_2}{Q_1} = \frac{T_2}{T_1} \quad ...(15)$$

When $T_2 \neq 0$, $Q_2 \neq 0$

i.e., some heat must be rejected to the sink. Heme efficiency of even an ideal heat engine can never be 100 %.

The same thing can be understood as follows :

If heat is not rejected to the sink, the same will be retained by the working substance. Therefore, temperature of working substance will go on increasing and would ultimately become equal to temperature of the source. There will be no transfer of heat energy from the source to the working substance and hence the engine will stop working.

6. Note that Carnot engine is a reversible engine. Therefore, each step of Carnot cycle can be reversed. This will amount to taking heat Q_2 from the cold reservoir (the sink) at lower temp. T_2, doing work W on the system and transferring heat Q_1 to the hot reservoir (the source at higher temp. T_1). This is what is done in a refrigerator/heat pump.

7. An ideal gas has a simple equation of state which allows us to calculate η of Carnot engine readily. That is why we use an ideal gas as the working substance in Carnot engine.

Problem

Carnot engine takes in a thousand kilo calories of heat from a reservoir at 627°C, and exhausts it to a sink at 27°C. How much work does it perform ? What is the efficiency of the engine ?

Sol. Herer,

$Q_1 = 1000$ kcal $= 10^6$ cal. $T_1 = 627°C = 627 + 273 = 900$ K, $T_2 = 27°C = 300$ K ; W= ?

As $\frac{Q_2}{Q_1} = \frac{T_2}{T_1}$ $\therefore$ $\frac{T_2}{T_1} Q_1 = \frac{300}{900} \times 10^6$ cals.

$$= \frac{1}{3} \times 10^6 \text{ cals}$$

Work done/cycle $W = Q_1 - Q_2 = 10^6 - \frac{1}{3} \times 10^6$

$$= \frac{2}{3} \times 10^6 \text{ cals.} = \frac{2 \times 4.2}{3} \times 10^6 \text{ J}$$

$$= 2.8 \times 10^6 \text{ J}$$

Also, $\eta = 1 - \frac{T_2}{T_1} = 1 - \frac{300}{900} = \frac{2}{3} = \frac{2}{3} \times 100 = 66.67\%$.

11.5.9 CARNOT THEOREM

According to this theorem,

(a) Working between two given temperatures, T_1 of hot reservoir (the source) and T_1 of cold reservoir (the sink), no engine can have efficiency more than that of the Carnot engine.

(b) The efficiency of the Carnot engine is independent of the nature of the working substance.

To prove the (a) part of the theorem, imagine a reversible Carnot engine R and an irreversible heat engine I. Working between the same source temp. Tx and same temp, of sink T_2, Fig.

Suppose the two engines are coupled in such a way that irreversible engine *I* acts like a heat engine and reversible carnot engine R acts like a refrigerator.

Let the engine *I*, in one full cycle absorb Q_1 heat from the source, deliver work W' and release the balance $(Q_1 - W')$ to the sink.

We arrange that the engine R returns the same heat Q_1 to the source when work W is done on it in one full cycle. Therefore, it would take heat $(Q_1 - W)$ from the sink as shown in Fig.

Suppose efficiency of reversible engine R is less than the efficiency of irreversible engine I, i.e.,$\eta_R < \eta_I$

For a given Q_1 ; W< W'

$\therefore$ $(Q_1 - W) > (Q_1 - W')$

Thus, on the whole, the coupled I – R system extracts heat $(Q_1 - W) - (Q_1 - W) = (W' - W)$ from the sink (cold reservoir) and delivers the same amount of work $(W' - W)$ in one cycle without any change in the source or anywhere else. This is against Second Law of Thermodynamics.

Hence, the assertion $\eta_R < \eta_I$ (or $(\eta_I > \eta_R)$ is wrong. Hence, working between two given temperatures, no engine can have efficiency greater than the efficiency of revervible Carnot engine. This proves part (a) of the carnot theorem.

We can build a similar argument to show that a reversible heat engine with one particular working substance cannot be more efficient than the engine using another working substance. Hence, the maximum efficiency of Carnot engine $\left(\eta = 1 - \frac{T_2}{T_1}\right)$ is independent of nature of the working substance undergoing Carnot cycle of operations.

Exercises – 1

FIRST LAW OF THERMODYNAMICS
$(\Delta Q = \Delta U + \Delta W)$

1. First law of thermodynamics is given by
[1] dQ = dU + PdV [2] dQ = dU x PdV
[2] dQ = (dU + dV)P [4] dQ = PdU + dV

2. The internal energy of an ideal gas depends upon
[1] Specific volume [2] Pressure
[3] Temperature [4] Density

3. In changing the state of thermodynamics from A to B state, the heat required is Q and the work done by the system is W. The change in its internal energy is
[1] Q + W [2] Q – W
[3] Q [4] Q – W/2

4. If a system undergoes contraction of volume then the work done by the system will be
[1] Zero [2] Negligible
[3] Negative [4] Positive

5. Identify the wrong statement
[1] For isothermal process, $\Delta T = 0$
[2] For isochoric process, $\Delta V = 0$
[3] For isobaric process, $\Delta P = 0$
[4] For adiabatic process, $\Delta Q = 0$
[5] For cyclic process, $\Delta W = 0$

6. The first law of thermodynamics is concerned with the conservation of
[1] Momentum [2] Energy
[3] Mass [4] Temperature

7. 0.100 mol of an ideal gas is heated from 10 to 20° C keeping its (a) volume constant (b) pressure constant. Let ΔU_a and ΔU_b denote the change in the internal energy of the gas due to process (a) and (b) respectively. Then which of the following shall be true
[1] $\Delta U_a = \Delta U_b$ [2] $\Delta U_a > \Delta U_b$
[3] $\Delta U_a = \Delta U_b$ [4] $\Delta U_a = \Delta U_b = 0$

8. The internal energy change in a system that has absorbed 2 Kcals of heat and done 500 J of work is
[1] 8900 J [2] 6400 J
[3] 5400 J [4] 7900 J

9. A system is given 300 calories of heat and it does 600 joules of work. How much does the internal energy of the system change in this process (J = 4.18 joules/cal)
[1] 654 joule [2] 156.5 joule
[3] – 300 joule [4] – 528.2 joule

10. If Q, E and W denote respectively the heat added, change in internal energy and the work done in a closed cycle process, then
[1] E = 0 [2] Q = 0
[3] W = 0 [4] Q = W = 0

11. The specific heat *c* of a solid at low temperature shows temperature dependence according to the relation $c = DT^3$ where D is a constant and T is the temperature in kelvin. A piece of this solid of mass m kg is taken and its temperature is raised from 20 K to 30 K. The amount of the heat required in the process in energy units is
[1] 5×10^4 Dm [2] $(33/4) \times 10^4$ Dm
[3] $(65/4) \times 10^4$ Dm [4] $(5/4) \times 10^4$ Dm

12. An ideal gas is made to go through a cyclic thermodynamical process in four steps. The amount of heat involved are $Q_1 = 600$ J, $Q_2 = -400$ J, $Q_3 = -300$ J and $Q_4 = 200$ J respectively. The corresponding work involved are $W_1 = 300$ J, $W_2 = -200$ J, $W_3 = -150$ J and W4. What is the value of W_4
[1] – 50 J [2] 100 J
[3] 150 J [4] 50J

13. A system performs work ΔW when an amount of heat is ΔQ added to the system, the corresponding change in the internal energy is ΔU. A unique function of the initial and final states (irrespective of the mode of change) is
[1] ΔQ [2] ΔW
[3] ΔU and ΔQ [4] ΔU.

14. A container of volume $1 m^3$ is divided into two equal compartments by a partition. One of these compartments contains an ideal gas at 300 K. The other compartment is vacuum. The whole system is thermally isolated from its surroundings. The partition is removed and the gas expands to occupy the whole volume of the container. Its temperature now would be
[1] 300 K [2] 239 K
[3] 200 K [4] 100 K

15. 110 J of heat is added to a gaseous system, whose internal energy change is 40 J, then the amount of external work done is
[1] 150 J [2] 70 J
[3] 110 J [4] 40J

16. Which of the following is not a thermodynamical function
[1] Enthalpy [2] Work done
[3] Gibb's energy [4] Internal energy

17. When the amount of work done is 333 cal and change in internal energy is 167 cal, then the heat supplied is
[1] 166 cal [2] 333 cal
[3] 500 cal [4] 400 cal

18. First law of thermodynamics states that
[1] System can do work
[2] System has temperature
[3] System has pressure
[4]Heat is a form of energy

19. A thermo-dynamical system is changed from state (P_1,V_1) to (P_2, V_2) by two different process. The quantity which will remain same will be
[1] ΔQ [2] ΔW
[3] $\Delta Q + \Delta W$ [4] $\Delta Q - \Delta W$

20. Which of the following statement is true
[1] Internal energy of a gas depends only on the state of the gas.
[2] In an isothermal process change in internal energy is maximum
[3] Area under pressure, volume graph equals heat supplied in any process
[4] Work done is state dependent but not path dependent

21. A perfect gas contained in a cylinder is kept in vacuum. If the cylinder suddenly bursts, then the temperature of the gas
[1] Remains constant [2] Becomes zero
[3] Increases [4] Decreases

22. When two bodies A and B are in thermal equilibrium
[1] The kinetic energies of all the molecules of A and B will be equal
[2] The potential energies of all the molecules of A and B will be equal
[3] The internal energies of the two bodies will be equal
[4] The average kinetic energy of the molecules of the two bodies will be equal

23. If ΔQ and ΔW represent the heat supplied to the system and the work done on the system respectively, then the first law of thermodynamics can be written as
[1] $\Delta Q = \Delta U + \Delta W$ [2] $\Delta Q = \Delta U - \Delta W$
[3] $\Delta Q = \Delta W - \Delta U$ [4] $\Delta Q = -\Delta W - \Delta U$
where ΔU is the internal energy

24. For free expansion of the gas which of the following is true
[1] $Q = W = 0$ and $\Delta E_{int} = 0$
[2] $Q = 0, W > 0$ and $\Delta E_{int} = -W$
[3] $W = 0, Q > 0$, and $\Delta E_{int} = Q$
[4] $W > 0, Q < 0$ and $\Delta E_{int} = 0$

25. Which of the following can not determine the state of a thermodynamic system
[1] Pressure and volume
[2] Volume and temperature
[3] Temperature and pressure
[4] Any one of pressure, volume or temperature

26. Which of the following is not a thermodynamic co-ordinate
[1] P [2] T [3] V [4] R

27. In a given process for an ideal gas, $dW = 0$ and $dQ < 0$ Then for the gas
[1] The temperature will decrease
[2] The volume will increase
[3] The pressure will remain constant
[4] The temperature will increase

28. The specific heat of hydrogen gas at constant pressure is $C_P = 3.4 \times 10^3$ cal/kg°C and at constant volume is $C_V = 2.4 \times 10^3$ cal/kg°C. If one kilogram hydrogen gas in heated from 10° C to 20° C at constant pressure, the external work done on the gas to maintain it at constant pressure is
[1] 105 cal [2] 104 cal
[3] 103 cal [4] 5×10^3 cal

29. Which of the following parameters does not characterize the thermodynamic state of matter
[1] Volume [2] Temperature*
[3] Pressure [4] Work

30. In a thermodynamic system working substance is ideal gas, its internal energy is in the form of
[1] Kinetic energy only
[2] Kinetic and potential energy
[3] Potential energy [4] None of these

31. Which of the following statements is correct for any thermodynamic system
[1] The internal energy changes in all processes
[2] Internal energy and entropy are state functions
[3] The change in entropy can never be zero
[4] The work done in an adiabatic process is always zero

32. Which of the following is incorrect regarding the first law of thermodynamics
[1] It introduces the concept of the internal energy
[2] It introduces the concept of the entropy
[3] It is a restatement of the principle of conservation of energy
[4] None of the above

33. In a thermodynamic process, pressure of a fixed mass of a gas is changed in such a manner that the gas molecules gives out 20 J of heat and 10 J of work is done on the gas. If the initial internal energy of the gas was 40 J, then the final internal energy will be
[1] 30 J [2] 20 J
[3] 60 J [4] 40 J

34. Heat is not being exchanged in a body. If its internal energy is increased, then
[1] Its temperature will increase
[2] Its temperature will decrease
[3] Its temperature will remain constant
[4] None of these

35. Out of the following which quantity does not depend on path
[1] Temperature [2] Energy
[3] Work [4] None of these

36. First law of thermodynamics is a special case of
[1] Newton's law
[2] Law of conservation of energy
[3] Charle's law
[4] Law of heat exchange

37. One mole of an ideal monoatomic gas is heated at a constant pressure of one atmosphere from 0°C to 100°C . Then the change in the internal energy is
[1] 6.56 joules [2] 8.32×10^2 joules
[3] 12.48×10^2 joules [4] 20.80 joules

38. If the ratio of specific heat of a gas at constant pressure to that at constant volume is γ , the change in internal energy of a mass of gas, when the volume changes from V to 2V constant pressure p, is
[1] $R/(\gamma - 1)$ [2] pV
[3] $pV/(\gamma - 1)$ [3] $\gamma pV/(\gamma - 1)$

39. If $C_V = 4.96$ cal/mole K, then increase in internal energy when temperature of 2 moles of this gas is increased from 340 K to 342 K
[1] 27.80 cal [2] 19.84 cal
[3] 13.90 cal [4] 9.92 cal

40. Temperature is a measurement of coldness or hotness of an object. This definition is based on
[1] Zeroth law of thermodynamics
[2] First law of thermodynamics
[3] Second law of thermodynamics
[4] Newton's law of cooling

41. When heat energy of 1500 joules, is supplied to a gas at constant pressure 2.1 x 10^5N/m^2, there was an increase in its volume equal to 2.5 x 10^{-3}m^3 . The increase in internal energy of the gas in joules is
[1] 450 [2] 525 [3] 975 [4] 2025

42. If heat given to a system is 6 kcal and work done is 6 kJ. Then change in internal energy is
[1] 19.1 kJ [2] 12.5 kJ
[3] 25 kJ [4] Zero

43. In a thermodynamics process, pressure of a fixed mass of a gas is changed in such a manner that the gas releases 20 J of heat and 8 J of work is done on the gas. If the initial internal energy of the gas was 30 J. The final internal energy will be
[1] 18 J [2] 9 J
[3] 4.5 J [4]36 J

44. A monoatomic gas of n-moles is heated from temperature T_1 to T_2 under two different conditions (i) at constant volume and (ii) at constant pressure. The change in internal energy of the gas is
[1] More for (i) [2] More for (ii)
[3] Same in both cases
[4] Independent of number of moles

45. The state of a thermodynamic system is represented by
[1] Pressure only [2] Volume only
[3] Pressure, volume and temperature
[4] Number of moles

46. A perfect gas goes from state A to another state B by absorbing 8 x 10^5 J of heat and doing 6.5 x 10^5J of external work. It is now transferred between the same two states in another process in which it absorbs 10^5 J of heat. Then in the second process
[1] Work done on the gas is 0.5 x 10^5 J
[2] Work done by gas is 0.5 x 10^5 J
[3] Work done on gas is 10^5 J
[4] Work done by gas is 10^5 J

47. During an isothermal expansion, a confined ideal gas does –150 J of work against its surrounding. This implies that
[1] 150 J of heat has been added to the gas
[2] 150 J of heat has been removed from the gas
[3] 300 J of heat has been added to the gas
[4] No heat is transferred because the process is isothermal

48. 100g of water is heated from 30°C to 50°C ignoring the slight expansion of the water, the change in its internal energy is (specific heat of wafer is 4184 J/Kg/K)
[1] 4.2 kJ [2] 8.4 kJ
[3] 84 kJ [4] 2.1 kJ

49. When an ideal monoatomic gas is heated at constant pressure, fraction of heat energy supplied which increases the internal energy of gas, is
[1] 2/5 [2] 3/5 [3] 3/7 [4] 3/4

50. If the amount of heat given to a system is 35 J and the amount of work done on the system is 15 J, then the change in internal energy of the system is
[1] – 50 J [2] 20J [3] 30 J
[4] 50J [5] – 20 J

ISOTHERMAL PROCESS

1. For an ideal gas, in an isothermal process
[1] Heat content remains constant
[2] Heat content and temperature remain constant
[3] Temperature remains constant
[4] None of the above

2. If a quantity of heat 1163.4 joule is supplied to one mole of nitrogen gas, at room temperature at constant pressure, then the rise in temperature is (Given R = 8.31 J $mole^{-1}R^{-1}$)
[1] 54K [2] 28K [3]65K
[4] 8K [4] 40K

3. In an isothermal expansion
[1] Internal energy of the gas increases
[2] Internal energy of the gas decreases
[3] Internal energy remains unchanged
[4] Average kinetic energy of gas molecule decreases

4. In an isothermal reversible expansion, if the volume of 96 g of oxygen at 27°C is increased from 70 litres to 140 litres, then the work done by the gas will be
[1] 300 R $\log_{10}2$ [2] 81 R $\log_e 2$
[3] 900 R $\log_{10}2$ [4] 2.3 x 900 R $\log_{10}2$

5. A vessel containing 5 litres of a gas at 0.8 m pressure is connected to an evacuated vessel of volume 3 litres. The resultant pressure inside will be (assuming whole system to be isolated)
[1] 4/3 m [2] 0.5 m
[3] 2.0 m [4] 3/4 m

6. An ideal gas undergoes a reversible isothermal expansion from a state i to *f*. The change in entropy ΔS ($i \rightarrow f$) is
[1] 0 [2] >0
[3] <0
[4] Not defined for the process

7. The gas law PV/T = constant is true for
[1] Isothermal changes only
[2] Adiabatic changes only
[3] Both isothermal and adiabatic changes
[4] Neither isothermal nor adiabatic changes

8. One mole of O_2 gas having a volume equal to 22.4 litres at 0°C and 1 atmospheric pressure in compressed isothermally so that its volume reduces to 11.2 litres. The work done in this process is
[1] 1672.5 J [2] 1728 J
[3] – 1728 J [4] – 1572.5 J

90 Work done in an isothermal change of a gas depends
[1] Only on temperature
[2] Only on volume expansion ratio
[3] On both the temperature and volume expansion ratio
[4] Only on initial and final pressure

10. Work done per mol in an isothermal change is
[1] $RT \log_{10}\frac{V_2}{V_1}$ [2] $RT \log_{10}\frac{V_1}{V_2}$
[3] $RT \log_e\frac{V_2}{V_1}$ [4] $RT \log_e\frac{V_1}{V_2}$

11. The isothermal Bulk modulus of an ideal gas at pressure P is
[1] P [2] γP [3] P/2 [4] P/γ

12. During an isothermal expansion of an ideal gas
[1] Its internal energy decreases
[2] Its internal energy does not change
[3] The work done by the gas is equal to the quantity of heat absorbed by it
[4] Both (b) and (c) are correct

13. The isothermal bulk modulus of a perfect gas at normal pressure is
[1] 1.013 x 10^5 N/m^2 [2] 1.013 x 10^6 N/m^2
[3] 1.013 x 10^{-11} N/m^2 [4] 1.013 x 10^{11} N/m^2

14. In an isothermal change, an ideal gas obeys
[1] Boyle's law [2] Charle's law
[3] Gaylussac law [4] None of the above

15. In isothermic process, which statement is wrong
[1] Temperature is constant
[2] Internal energy is constant
[3] No exchange of energy
[4] (a) and (b) are correct

16. An ideal gas A and a real gas B have their volumes increased from V to 2V under isothermal conditions. The increase in infernal energy

[1] Will be same in both A and B
[2] Will be zero in both the gases
[3] Of B will be more than that of A
[4] Of A will be more than that of B

17. The specific heat of a gas in an isothermal process is
[1] Infinite [2] Zero
[3] Negative [4] Remains constant

18. A thermally insulated container is divided into two parts by a screen. In one part the pressure and temperature are P and T for an ideal gas filled. In the second part it is vacuum. If now a small hole is created in the screen, then the temperature of the gas will
[1] Decrease [2] Increase
[3] Remain same [4] None of the above

19. A container that suits the occurrence of an isothermal process should be made of
[1] Copper [2] Glass
[3] Wood [4] Cloth

20. In an isothermal process the volume of an ideal gas is halved. One can say that
[1] Internal energy of the system decreases
[2] Work done by the gas is positive
[3] Work done by the gas is negative
[4] Internal energy of the system increases

21. A thermodynamic process in which temperature T of the system remains constant though other variable P and V may change, is called
[1] Isochoric process [2] Isothermal process
[3] Isobaric process [4] None of these

22. If an ideal gas is compressed isothermally then
[1] No work is done against gas
[2] Heat is released by the gas
[3] The internal energy of gas will increase
[4] Pressure does not change

23. When an ideal gas in a cylinder was compressed isothermally by a piston, the work done on the gas was found to be 1.5×10^4 joules . During this process about
[1] 3.6×10^3 cal of heat flowed out from the gas
[2] 3.6×10^3 cal of heat flowed into the gas
[3] 1.5×10^4 cai of heat flowed into the gas
[4] 1.5×10^4 cal of heat flowed out from the gas

24. When heat is given to a gas in an isothermal change, the result will be
[1] External work done
[2] Rise in temperature
[3] Increase in internal energy
[4] External work done and also rise in temp.

25. When 1 g of water at 0°C and 1×10^5 N/m^2 pressure is converted into ice of volume 1.091 cm^3, the external work done will be
[1] 0.0091 joule [2] 0.0182 joule
[3] – 0.0091 joule [4] – 0.0182 joule

26. The latent heat of vaporisation of water is 2240 J/g. If the work done in the process of expansion of 1 g is 168 J, then increase in internal energy is
[1] 2408 J [2] 2240 J
[3] 2072 J [4] 1904 J

27. 540 calories of heat convert 1 cubic centimeter of water at 100° C into 1671 cubic centimeter of steam at 100° C at a pressure of one atmosphere. Then the work done against the atmospheric pressure is nearly
[1] 540 cal [2] 40 cal
[3] Zero cal [4] 500 cal

28. One mole of an ideal gas expands at a constant temperature of 300 K from an initial volume of 10 litres to a final volume of 20 litres. The work done in expanding the gas is (R = 8.31 J/mole–K)
[1] 750 joules [2] 1728 joules
[3] 1500 joules [4] 3456 joules

29. A cylinder fitted with a piston contains 0.2 moles of air at temperature 27°C. The piston is pushed so slowly that the air within the cylinder remains in thermal equilibrium with the surroundings. Find the approximate work done by the system if the final volume is twice the initial volume

[1] 543 J [2] 345 J
[3] 453 J [4] 600 J

30. The volume of an ideal gas is 1 litre and its pressure is equal to 72 cm of mercury column. The volume of gas is made 900 cm^3 by compressing it isothermally. The stress of the gas will be
[1] 8 cm (mercury) [2] 7 cm (mercury)
[3] 6 cm (mercury) [4] 4 cm (mercury)

ADIABATIC PROCESS

1. During an adiabatic process, the pressure of a gas is found to be proportional to the cube of its absolute temperature. The y of the gas is
[1] 1.3 [2] 2 [3] 1.7 [4] 1.5

2. The work done in an adiabatic change in a gas depends only on
[1] Change is pressure [2] Change is volume
[3] Change in temperature
[4] None of the aboe.

3. In adiabatic expansion
[1] $\Delta U = 0$ [2] ΔU = negative
[3] ΔU = positive [4] ΔW = Zero.

4. For which of the following processes is the entropy change zero
[1] Isobaric [2] Isothermal
[3] Adiabatic [4] None of the above

5. 1 mm^3 of a gas is compressed at 1 atmospheric pressure and temperature 27°C to 627°C. What is the final pressure under adiabatic condition (γ for the gas = 1.5)
[1] 27×10^5 N/m^2 [2] 80×10^5 N/m^2
[3] 36×10^5 N/m^2 [4] 56×10^5 N/m^2

6. A monoatomic gas ($\gamma = 5/3$) is suddenly compressed to 1/8 of its original volume adiabatically, then the pressure of the gas will change to
[1] 24/5 [2] 8
[3] 40/3
[4] 32 times its initial pressure

7. The pressure and density of a diatomic gas ($\gamma = 7/5$) change adiabatically from (P, d) to (P', d'). If d'/d = 32, then P'/P should be
[1] 1/128 [2] 32
[3] 128 [4] None of the above

8. An ideal gas at 27° C is compressed adiabatically to 8/27 of its original volume. If $\gamma = 5/3$, then the rise in temperature is
[1] 450 K [2] 375 K
[3] 225 K [4] 405 K

9. Two identical samples of a gas are allowed to expand (i) isothermally (ii) adiabatically. Work done is
[1] More in the isothermal process
[2] More in the adiabatic process
[3] Neither of them
[4] Equal in both processes

10. Which of the two, (i) compressing a gas isothermally until its volume is reduced by half (ii) compressing the same gas adiabatically until its volume is reduced by half, will require more work to be done
[1] (i) [2] (ii)
[3] Both will require the same amount of work
[4] It will depend upon the nature of the gas

11. The slopes of isothermal and adiabatic curves are related as
[1] Isothermal curve slope = adiabatic curve slope
[2] Isothermal curve slope = $\gamma \times$ adiabatic curve slope
[3] Adiabatic curve slope = $\gamma \times$ isothermal curve slope
[4] Adiabatic curve slope = 1/2 x isothermal curve slope

12. Pressure-temperature relationship for an ideal gas undergoing adiabatic change is ($\gamma = C_p/C_v$)
[1] PT^{γ} = constant [2] $PT^{-1+\gamma}$ = constant
[3] $P^{\gamma-1}T^{\gamma}$ = constant [4] $P^{1-\gamma}T^{\gamma}$ = constant

13. Ideal gas undergoes an adiabatic change in its state from (P_1,V_1,T_1) to (P_2,V_2,T_2) . The work done (W) in the process is (μ = number of moles, C_p and C_v are molar specific heats of gas)

[1] $W = \mu(T_1 - T_2)C_p$ [2] $W = \mu(T_1 - T_2)C_v$
[1] $W = \mu(T_1 + T_2)C_p$ [2] $W = \mu(T_1 + T_2)C_v$

14. For a gas of non rigid diatomic molecules, the value of $\gamma = C_p/ C_v$ is

[1] 9/7 [2] 7/5
[3] 5/3 [4] 11/9

15. A thermally insulated vessel contains an ideal gas of molecular mass M and ratio of specific heats γ. It is moving with speed v and is suddenly brought to rest. Assuming no heat is lost to the surroundings, its temperature increases by

[1] $\frac{(\gamma - 1)}{2(\gamma + 1)R} Mv^2$ [2] $\frac{(\gamma - 1)}{2\gamma R} Mv^2$
[3] $\frac{\gamma Mv^2}{2R}$ [4] $\frac{(\gamma - 1)}{2R} Mv^2$

16. If γ denotes the ratio of two specific heats of a gas, the ratio of slopes of adiabatic and isothermal PV curves at their point of intersection is

[1] $1/\gamma$ [2] γ
[3] $\gamma - 1$ [4] $\gamma + 1$

17. Air in a cylinder is suddenly compressed by a piston, which is then maintained at the same position. With the passage of time

[1] The pressure decreases
[2] The pressure increases
[3] The pressure remains the same
[4] The pressure may increase or decrease depending upon the nature of the gas

18. When a gas expands adiabatically

[1] No energy is required for expansion
[2] Energy is required and it comes from the wall of the container of the gas
[3] Internal energy of the gas is used in doing work
[4] Law of conservation of energy does not hold

19. A thermos flask made of stainless steel contains several tiny lead shots. If the flask is quickly shaken, up and down several times, the temperature of lead shots

[1] Increases by adiabatic process
[2] Increases by isothermal process
[3] Decreases by adiabatic process
[4] Remains same

20. In changing the state of a gas adiabatically from an equilibrium state A to another equilibrium state B, an amount of work equal to 22.3 J is done on the system. If the gas is taken from state A to B via a process in which the net heat absorbed by the system is 9.35 calories, the net work done by the system in latter case will be

[1] 5.9 J [2] 16.9 J
[3] 9.3 J [4] 4.6 J

21. The adiabatic Bulk modulus of a perfect gas at pressure is given by

[1] P [2] 2P
[3] P/2 [4] γ P

22. An adiabatic process occurs at constant

[1] Temperature
[2] Pressure
[3] Heat
[4]Temperature and pressure

23. How many times a diatomic gas should be expanded adiabatically so as to reduce the root mean square velocity to half

[1] 64 [2] 32 [3] 16 [4] 8.

24. For adiabatic processes $\left(\gamma = \frac{C_p}{C_v}\right)$

[2] $P^{\gamma}V$ = constant [2] $T^{\gamma}V$ = constant
[3] $TV^{\gamma-1}$= constant [4] TV^{γ}= constant

25. An ideal gas is expanded adiabatically at an initial temperature of 300 K so that its volume is doubled. The final temperature of the hydrogen gas is (γ = 1.40)

[1] 227.36 K [2] 500.30 K
[3] 454.76 K [4] – 47°C

26. A given system undergoes a change in which the work done by the system equals the decrease in its internal energy. The system must have undergone an
[1] Isothermal change
[2] Adiabatic change
[3] Isobaric change
[4] Isochoric change

27. During the adiabatic expansion of 2 moles of a gas, the internal energy was found to have decreased by 100 J. The work done by the gas this process is
[1] Zero [2] 100 J
[3] 200 J [4]100 J

28. In an adiabatic expansion of a gas initial and final temperatures are T_1 and T_2 respectively, then the change in internal energy of the gas is
[1] $\frac{R}{\gamma-1}(T_2-T_1)$ [2] $\frac{R}{\gamma-1}(T_1-T_2)$
[3] $R(T_1-T_2)$ [4] Zero

29. Helium at 27° C has a volume of 8 litres. It is suddenly compressed to a volume of 1 litre. The temperature of the gas will be [$\gamma = 5/3$]
[1] 108° C [2] 9327° C
[3] 1200° C [4] 927° C

30. The work of 146 kJ is performed in order to compress one kilo mole of a gas adiabatically and in this process the temperature of the gas increases by 7°C. The gas is ($R = 8.3\ J\ mol^{-1}\ K^{-1}$)
[1] Triatomic
[2] A mixture of monoatomic and diatomic
[3] Monoatomic [4] Diatomic

31. For an adiabatic expansion of a perfect gas, the value of $\Delta P/P$ is equal to
[1] $-\sqrt{\gamma}\frac{\Delta V}{V}$ [2] $-\frac{\Delta V}{V}$
[3] $-\gamma\frac{\Delta V}{V}$ [4] $-\gamma^2\frac{\Delta V}{V}$

32. At N.T.P. one mole of diatomic gas is compressed adiabatically to half of its volume, $\gamma = 1.41$. The work done on gas will be
[1] 1280 J [2] 1610 J
[3] 1815 J [4] 2025 J

33. For adiabatic process, wrong statement is
[1] $dQ = 0$ [2] $dU = -dW$
[3] $Q =$ constant
[4] Entropy is not constant

34. A diatomic gas initially at 18°C is compressed adiabatically to one-eighth of its original volume. The temperature after compression will be
[1] 10°C [2] 887°C
[3] 668K [4] 144° C

35. 5.6 liter of helium gas at STP is adiabatically compressed to 0.7 liter. Taking the initial temperature to be T_1, the work done in the process is
[1] $\frac{9}{8}RT_1$ [2] $\frac{3}{2}RT_1$
[3] $\frac{15}{8}RT_1$ [4] $\frac{9}{2}RT_1$

36. The process in which no heat enters or leaves the system is termed as
[1] Isochoric [2] Isobaric
[4] Isothermal [4] Adiabatic

37. Two moles of an ideal monoatomic gas at 27°C occupies a volume of V. If the gas is expanded adiabatically to the volume 2V, then the work done by the gas will be [$\gamma = 5/3, R = 8.31\ J/molK$]
[1] – 216123 J [2] 262123 J
[3] 2500 J [4] – 2500 J

38. Adiabatic modulus of elasticity of a gas is $2.1 \times 10^5\ N/m^2$. What will be its isothermal modulus of elasticity $\left(\frac{C_p}{C_v} = 1.4\right)$
[1] $1.8 \times 10^5\ N/m^2$ [2] $1.5 \times 10^5\ N/m^2$
[3] $1.4 \times 10^5\ N/m^2$ [4] $1.2 \times 10^5\ N/m^2$

39. $\Delta U + \Delta W = 0$ is valid for
[1] Adiabatic process [2] Isothermal process
[3] Isobaric process [4] Isochoric process

40. An ideal gas at a pressures of 1 atmosphere and temperature of 27° C is compressed adiabatically until its pressure becomes 8 times the initial pressure, then the final temperature is ($\gamma = 3/2$)

[1] 627° C [2] 527° C
[3] 427° C [4] 327° C

41. Air is filled in a motor tube at 27° C and at a pressure of 8 atmospheres. The tube suddenly bursts, then temperature of air is
[Given γ of air = 1.5]

[1] 27.5° C [2] 75° K
[3] 150° K [4] 150° C

42. If $\gamma = 2.5$ and volume is equal to 1/8 times to the initial volume then pressure P' is equal to (Initial pressure = P)

[1] $P' = P$ [2] $P' = 2P$
[3] $P' = P \times (2)^{15/2}$ [4] $P' = 7P$

43. During an adiabatic expansion, the increase in volume is associated with which of the following possibilities w.r.t. pressure and temperature

	Pressure	**Temperature**
[1]	Increase	Increase
[2]	Decrease	Decrease
[3]	Increase	Decrease
[4]	Decrease	Increase

44. During an adiabatic process, the pressure of a gas is found to be proportional to the cube of its absolute temperature. The ratio C_p / C_v for the gas is

[1] 3/2 [2] 4/3
[3] 2 [4] 5/3.

45. In adiabatic expansion of a gas
[1] Its pressure increases
[2] Its temperature falls
[3] Its density increases
[4] Its thermal energy increases

46. One mole of an ideal gas at an initial temperature of T K does 6 R joules of work adiabatically. If the ratio of specific heats of this gas at constant pressure and at constant volume is 5/3, the final temperature of gas will be

[1] (T + 2.4) K [2] (T – 2.4)K
[3] (T + 4)K [4] (T – 4) K

47. One mole of an ideal gas with $\gamma = 1.4$, is adiabatically compressed so that its temperature rises from 27°C to 35°C. The change in the internal energy of the gas is (R = 8.3 J/mol.K)

[1] – 166 J [2] 166 J
[3] – 168 J [4] 168 J

48. A gas for which $\gamma = 1.5$ is suddenly compressed to 1/4th of the initial volume. Then the ratio of the final to the initial pressure is

[1] 1:16 [2] 1:8
[3] 1:4 [4] 8:1

49. The internal energy of the gas increases in
[1] Adiabatic expansion
[2] Adiabatic compression
[3] Isothermal expansion
[4] Isothermal compression

50. We consider a thermodynamic system. If ΔU represents the increase in its internal energy and W the work done by the system, which of the following statements is true
[1] AD = –W in an adiabatic process
[2] AD = W in an isothermal process
[3] AD = –W in an isothermal process
[4] AD = W in an adiabatic process

51. Which property of the system does not change in an adiabatic process

[1] Volume [2] Temperature
[3] Density
[4] Total amount of heat

ISOBARIC AND ISOCHORIC PROCESSES

1. A gas expands under constant pressure P from volume V_1 to V_2. The work done by the gas is

[1] $P(V_2 - V_1)$ [2] $P(V_2 - V_1)$
[3] $P(V_1^{\gamma} - V_2^{\gamma})$ [4] $P\dfrac{V_1 V_2}{V_2 - V_1}$

2. When heat in given to a gas in an isobaric process, then
[1] The work is done by the gas
[2] Internal energy of the gas increases
[3] Both (1) and (2)
[4] None from (1) and (2)

3. In thermodynamic processes which of the following statements is not true
[1] In an adiabatic process the system is insulated from the surroundings
[2] In an isochoric process pressure remains constant
[3] In an isothermal process the temperature remains constant
[4] In an adiabatic process PV^{γ} = constant

4. A frictionless piston–cylinder based enclosure contains some amount of gas at a pressure of 400 kPa. Then heat is transferred to the gas at constant pressure in a quasi–static process. The piston moves up slowly through a height of 10 cm. If the piston has a cross-section area of 0.3 m^2, the work done by the gas in this process is
[1] 6 kJ [2] 12 kJ
[3] 7.5 kJ [4] 24 kJ

5. In which process the PV indicator diagram is a straight line parallel to volume axis
[1] Isothermal [2] Isobaric
[3] Irreversible [4] Adiabatic

6. In which of processes, does the internal energy of the system remain constant
[1] Isobaric [2] Isothermal
[3] Adiabatic [4] Isochoric

7. Unit mass of a liquid with volume V_1 is completely changed into a gas of volume V_2 at a constant external pressure P and temperature T. If the latent heat of evaporation for the given mass is L, then the increase in the internal energy of the system is
[1] Zero [2] $P(V_2 - V_1)$
[3] $L - P(V_2 - V_1)$ [4] L

8. A gas expands 0.25m^3 at constant pressure 10^3 N/m^2 , the work done is
[1] 2.5 ergs [2] 250 J
[3] 250 W [4] 250 N

9. Two kg of water is converted into steam by boiling at atmospheric pressure. The volume changes from 2 x 10 m^3 to 3.34 m^3. The work done by the system is about
[1] – 340 kJ [2] – 170 kJ
[3] 170 kJ [4] 340 kJ

10. An ideal gas has volume V_0 at 27° C. It is heated at constant pressure so that its volume becomes $2V_0$. The final temperature is
[1] 54° C [2] 32.6° C
[3] 327 °C [4] 150 K

11. If 300 ml of a gas at 27° C is cooled to 7°C at constant pressure, then its final volume will be
[1] 540 m*l* [2] 350 m*l*
[4] 280 m*l* [4] 135 m*l*

12. If the temperature of 1 mole of ideal gas is changed from 0°C to 100°C at constant pressure, then work done in the process is (R = 8.3 J/Mole–Kelvin)
[1] 8.3×10^{-3} J [2] 8.3×10^{-2} J
[2] 8.3×10^{2} J [4] 8.3×10^{3} J

13. A sample of gas expands from volume V_1 to V_2. The amount of work done by the gas is greatest when the expansion is
[1] Isothermal [2] Isobaric
[3] Adiabatic [4] Equal in all cases.

14. Which of the following is a slow process
[1] Isothermal [2]Adiabatic
[3] Isobaric [4] None of these

15. How much work to be done in decreasing the volume of an ideal gas by an amount of 2.4×10^{-4} m^3 at normal temperature and constant normal pressure of 1×10^5 N/m^2
[1] 28 J [2] 27 J
[3] 25 J [4] 24 J

16. A Container having 1 mole of a gas at a temperature 27°C has a movable piston which maintains at constant pressure in container of 1 atm. The gas is compressed until temperature becomes 127°C. The work done is (C for gas is 7.03 cal/mol–K)
[1] 703 J [2] 831 J
[3] 121 J [4] 2035 J

17. In a reversible isochoric change
[1] $\Delta W = 0$ [2] $\Delta Q = 0$
[3] $\Delta T = 0$ [4] $\Delta U = 0$

18. Which is incorrect
[1] In an isobaric process, $\Delta P = 0$
[2] In an isochoric process, $\Delta W = 0$
[3] In an isothermal process, $\Delta T = 0$
[4] In an isothermal process, $\Delta Q = 0$

19. Which relation is correct for isometric process
[1] $\Delta Q = \Delta U$ [2] $\Delta W = \Delta U$
[3] $\Delta Q = \Delta W$ [4] None of these

20. In which thermodynamic process, volume remains same
[1] Isobaric [2] Isothermal
[3] Adiabatic [4] Isochoric

21. In an isochoric process if T_1=27°C and T_2 =127°C, then P_1/P_2 will be equal to
[1] 9/59 [2] 2/3
[3] 3/4 [4] None of these

HEAT ENGINE, REFRIGERATOR AND SECOND LAW OF THERMODYNAMICS

1. A Carnot engine working between 300K and 600K has work output of 800 J per cycle. What is amount of heat energy supplied to the engine from source per cycle
[1] 1800 J/cycle [2] 1000 J/cycle
[3] 2000 J/cycle [4] 1600 J/cycle

2. The coefficient of performance of a Carnot refrigerator working between 30°C and 0°C is
[1] 10 [2] 1 [3] 9 [4] 0

3. If the door of a refrigerator is kept open, then which of the following is true
[1] Room is cooled [2] Room is heated
[3] Room is either cooled or heated
[4] Room is neither cooled nor heated

4. If we consider solar system consisting of the earth and sun only as one of the ideal thermodynamic system. The sun works as source of energy having temperature 6000 K and the earth as sink having temperature 300K, the efficiency of solar system would be on the basis of exchange of radiations
[1] 30% [2] 65% [3] 75% [4] 95%

5. What is the value of sink temperature when efficiency of engine is 100 %
[1] 0 K [2] 300 K
[3] 273 K [4] 400 K

6. In a Carnot engine, the temperature of reservoir is 927°C and that of sink is 27°C. If the work done by the engine when it transfers heat from reservoir to sink is 12.6×10^6c/, the quantity of heat absorbed by the engine from the reservoir is
[1] 16.8×10^6 J [2] 4×10^6 J
[3] 7.6×10^6 J [4] 4.2×10^6 J
[5] 20.8×10^6 J

7. In a cyclic process, work done by the system is
[1] Zero
[2] Equal to heat given to the system
[3] More than the heat given to system
[4] Independent of heat given to the system

8. A carnot engine, having an efficiency of $\eta = 1/10$ as heat engine, is used as a refrigerator. If the work done on the system is 10 J, the amount of energy absorbed from the reservoir at lower temperature is
[1] 99 J [2] 90 J [3] 1 J [4] 100 J

9. The change in the entropy of a 1 mole of an ideal gas which went through an isothermal process from an initial state (P_1, V_1,T) to the final state ((P_2, V_2,T)) is equal to

[1] Zero [2] R In T
[3] R In V_1/V_2 [4] R In V_2/V_1

10. Efficiency of Carnot engine is 100% if
[1] $T_2 = 273$ K [2] $T_2 = 0$ K
[3] $T_1 = 273$ K [4] $T_1 = 0$ K

11. A Carnot's engine used first an ideal monoatomic gas then an ideal diatomic gas. If the source and sink temperature are 411° C and 69 °C respectively and the engine extracts 1000 J of heat in each cycle, then area enclosed by the PV diagram is
[1] 100 J [2] 300 J
[3] 500 J [4] 700 J

12. A Carnot engine absorbs an amount Q of heat from a reservoir at an absolute temperature T and rejects heat to a sink at a temperature of T/3. The amount of heat rejected is
[1] Q /4 [2] Q /3
[3] Q /2 [4] 2Q /3

13. The temperature of sink of Carnot engine is 27°C. Efficiency of engine is 25%. Then temperature of source is
[1] 227°C [2] 327°C
[3] 127°C [4] 27°C

14. The temperature of reservoir of Carnot's engine operating with an efficiency of 70% is 1000K. The temperature of its sink is
[1] 300 K [2] 400 K
[3] 500 K [4] 700 K

15. In a Carnot engine, when $T_2=0°C$ and $T_1=200°C$, its efficiency is η_1 and when $T_1 = 0°C$ and $T_2 = -200°C$, Its efficiency is η_2, then what is η_1 / η_2
[1] 0.577 [2] 0.733
[3] 0.638
[4] Cannot be calculated

16. The efficiency of Carnot's engine operating between reservoirs, maintained at temperatures 27°C and –123°C, is
[1] 50% [2] 24% [3] 0.75% [4] 0.4%

17. A Carnot engine has efficiency 1/5. Efficiency becomes 1/3 when temperature of sink is decreased by 50K. What is the temperature of sink
[1] 325 K [2] 375 K
[3] 300 K [4] 350 K

18. A measure of the degree of disorder of a system is known as
[1] Isobaric [2] Isotropy
[3] Enthalpy [4] Entropy

19. A carnot engine has the same efficiency between 800 K to 500 K and *x* K to 600 K. The value of *x* is
[1] 1000 K [2] 960 K
[3] 846 K [4] 754 K

20. A scientist says that the efficiency of his heat engine which operates at source temperature 127°C and sink temperature 27°C is 26%, then
[1] It is impossible
[2] It is possible but less probable
[3] It is quite probable
[4] Data is incomplete

21. Which of the following processes is reversible
[1] Transfer of heat by radiation
[2] Electrical heating of a nichrome wire
[3] Transfer of heat by conduction
[4] Isothermal compression

22. Efficiency of a Carnot engine is 50% when temperature of outlet is 500 K. In order to increase efficiency up to 60% keeping temperature of intake the same what is temperature of outlet
[1] 200 K [2] 400 K
[3] 600 K [4] 800 K

23. Even Carnot engine cannot give 100% efficiency because we cannot
[1] Prevent radiation [2] Find ideal sources
[3] Reach absolute zero temperature
[4] Eliminate friction

24. “Heat cannot by itself flow from a body at lower temperature to a body at higher temperature” is a statement or consequence of
[1] Second law of thermodynamics
[2] Conservation of momentum
[3] Conservation of mass
[4] First law of thermodynamics

25. A Carnot engine takes 3 x 10^6 cal of heat from a reservoir at 627°C, and gives it to a sink at 27°C. The work done by the engine is
[1] 4.2 x 10^6 J [2] 8.4 x 10^6 J
[3]16.8 x 10^6 J [4] Zero

26. The first operation involved in a Carnot cycle is
[1] Isothermal expansion
[2] Adiabatic expansion
[3] Isothermal compression
[4] Adiabatic compression

27. For which combination of working temperatures the efficiency of Carnot’s engine is highest
[1] 80 K, 60 K [2] 100K, 80K
[3] 60 K, 40 K [4] 40 K, 20 K

28. The efficiency of Carnot engine when source temperature is T_1 and sink temperature is T_2 will be
[1] $T_1 – T_2/ T_1$ [2] $T_2– T_1/ T_2$
[3] $T_1 – T_2/ T_2$ [4] T_1 /T_2

29. An ideal heat engine working between temperature T_1 and T_2 has an efficiency η, the new efficiency if both the source and sink temperature are doubled, will be
[1] η/2 [2] η
[3] 2η [4] 3η.

30. An ideal refrigerator has a freezer at a temperature of –13°C. The coefficient of performance of the engine is 5. The temperature of the air (to which heat is rejected) will be
[1] 325°C [2] 325K
[3] 39°C [4] 320°C

31. In a mechanical refrigerator, the low temperature coils are at a temperature of – 23°C and the compressed gas in the condenser has a temperature of 27°C. The theoretical coefficient of performance is
[1] 5 [2] 8 [3] 6 [4] 6.5

32. An engine is supposed to operate between two reservoirs at temperature 727°C and 227°C. The maximum possible efficiency of such an engine is
[1] 1/2 [2] 1/4
[3] 3/4 [4] 1.

33. An ideal gas heat engine operates in Carnot cycle between 227°C and 127°C. It absorbs 6 x 10^4 cals of heat at higher temperature. Amount of heat converted to work is
[1] 2.4 x 10^4 cal [2] 6 x 10^4 cal
[3] 1.2 x 10^4 cal [4] 4.8 x 10^4 cal

34. When you make ice cubes, the entropy of water
[1] Does not change [2] Increases
[3] Decreases
[4] May either increase or decrease depending on the process used

35. If heat Q is added reversibly to a system at temperature T and heat Q' is taken away from it reversibly at temperature T" then which one of the following is correct
[1] $\frac{Q}{T} - \frac{Q'}{T} = 0$ [2] $\frac{Q}{T} - \frac{Q'}{T} > 0$
[3] $\frac{Q}{T} - \frac{Q'}{T} < 0$
[4] $\frac{Q}{T} - \frac{Q'}{T}$ = change in internal energy of the system

36. The efficiency of Carnot’s heat engine is 0.5 when the temperature of the source is T_1 and that of sink is T_2. The efficiency of another Carnot’s heat engine is also 0.5. The temperature of source and sink of the second engine are respectively
[1] $2T_1$, $2T_2$ [2] $2T_1$, $T_2/2$
[3] $T_1 + 5$, $T_2 – 5$ [4] $T_1 + 10$, $T_2 – 10$

37. A Carnot cycle has the reversible processes in the following order

[1] Isothermal expansion, adiabatic expansion, isothermal compression and adiabatic compression

[2] Isothermal compression, adiabatic expansion, isothermal expansion and adiabatic compression

[3] Isothermal expansion, adiabatic compression, isothermal compression and adiabatic expansion

[4] Adiabatic expansion, isothermal expansion, adiabatic compression and isothermal compression

38. A refrigerator with coefficient of performance 1/3 releases 200 J of heat to a hot reservoir. Then the work done on the working substance is

[1] 100/3 J [2] 100 J

[3] 200/3 J [4] 150 J

[5] 50 J.

39. Choose the incorrect statement from the following

S1 : The efficiency of a heat engine can be 1, but the coefficient of performance of a refrigerator can never be infinity

S2: The first law of thermodynamics is basically the principle of conservation of energy

S3: The second law of thermodynamics does not allow several phenomena consistent with the first law

S4 : A process, whose sole result is the transfer of heat from a colder object to a hotter object is impossible

[1] S1 [2] S3

[3] S2 [4] S4

40. A Carnot engine operating between temperatures T_1 and T_2 has efficiency 1/6. When T_2 is lowered by 62 K, its efficiency increases to 1/3. Then T_1 and T_2 are, respectively

[1] 372 K and 310 K [2] 372 K and 330 K

[3] 330 K and 268 K [4] 310 K and 248 K

41. 200 cal of heat is given to a heat engine so that it rejects 150 cal of heat, if source temperature is 400 K, then the sink temperature is

[1] 300 K [2] 200 K

[3] 100 K [4] 50 K

42. A carnot engine, whose efficiency is 40%, takes in heat from a source maintained at a temperature of 500K. It is desired to have an engine of efficiency 60%. Then, the intake temperature for the same exhaust (sink) temperature must be

[1] Efficiency of carnot engine cannot be made larger than 50%

[2] 1200 K

[3] 750 K

[4]600 K

43. Which of the following is a true statement

[1] The total entropy of thermally interacting systems is conserved

[2] Carnot engine has 100% efficiency

[3] Total entropy does not change in a reversible process

[4] Total entropy in an irreversible process can either increase or decrease

-----x-----x-----x-----

Exercise – 1

First Law of Thermodynamics($\Delta Q = \Delta U + \Delta W$)

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
A	**1**	**3**	**2**	**3**	**5**	**2**	**1**	**4**	**1**	**1**	**3**	**3**	**4**	**1**	**2**
Q	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
A	**2**	**3**	**4**	**4**	**1**	**1**	**3**	**2**	**1**	**4**	**4**	**1**	**2**	**4**	**1**
Q	31	32	33	34	35	36	37	38	39	40	41	42	43	44	45
A	**2**	**1**	**1**	**1**	**1**	**2**	**3**	**3**	**2**	**1**	**3**	**1**	**1**	**3**	**3**
Q	46	47	48	49	50										
A	**1**	**2**	**2**	**2**	**4**										

Isothermal Process

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
A	**3**	**5**	**3**	**4**	**2**	**2**	**3**	**4**	**3**	**3**	**1**	**4**	**1**	**1**	**3**
Q	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
A	**3**	**1**	**3**	**1**	**3**	**2**	**2**	**1**	**1**	**1**	**3**	**2**	**2**	**2**	**1**

Adiabatic Process

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
A	**4**	**3**	**2**	**3**	**1**	**4**	**3**	**2**	**1**	**2**	**3**	**4**	**2**	**2**	**4**
Q	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
A	**2**	**1**	**3**	**1**	**2**	**4**	**3**	**2**	**3**	**1**	**2**	**4**	**1**	**2**	**2**
Q	31	32	33	34	35	36	37	38	39	40	41	42	43	44	45
A	**3**	**3**	**4**	**3**	**1**	**4**	**2**	**2**	**1**	**4**	**3**	**3**	**2**	**1**	**2**
Q	46	47	48	49	50	51									
A	**4**	**2**	**4**	**2**	**1**	**4**									

Isobaric and Isochoric Processes

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
A	**1**	**3**	**2**	**2**	**2**	**2**	**3**	**2**	**4**	**3**	**3**	**3**	**2**	**1**	**4**
Q	16	17	18	19	20	21									
A	**2**	**1**	**4**	**1**	**4**	**3**									

Heat Engine, refrigerator and Second Law of Thermodynamics

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
A	**4**	**3**	**2**	**4**	**1**	**1**	**2**	**2**	**4**	**2**	**3**	**2**	**3**	**1**	**1**
Q	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
A	**1**	**3**	**4**	**2**	**1**	**4**	**2**	**3**	**1**	**2**	**1**	**4**	**1**	**2**	**3**
Q	31	32	33	34	35	36	37	38	39	40	41	42	43		
A	**1**	**1**	**3**	**3**	**1**	**1**	**1**	**4**	**1**	**1**	**1**	**3**	**3**		