PRACTICAL CENTRE

CLASS-XII

PHYSICS

UNIT # 16

FIRST LAW OF THERMODYNAMIC

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FIRST LAW OF THERMODYNAMIC 16

16.1- THERMODYNAMICS:

The branch of physics which deals with the transformation of heat energy into mechanical energy is called thermodynamics.

SOME IMPORTANT TERMS FOR THERMODYNAMICS:

1) System and Surrounding:

The part of universe which is under consideration is called system and remaining part is known as surrounding.

Types of Systems:

a) Open System:

In an open system, both heat and matter can be exchanged with the surroundings.

b) Closed System:

In a closed system, only heat can be exchanged with surrounding, but matter cannot.

c) Isolated System:

In an isolated system, neither heat nor matter can be exchanged with the surroundings.

OR

An isolated system is the one which has no surroundings interaction, that is, there is no flow of heat in or out from the system and hence, the system is not capable of doing work, therefore:

$$\Delta Q = 0$$

And $\Delta W = 0$

From the first law of Thermodynamics,

$$\Delta Q = \Delta U + \Delta W$$

We have $0 = \Delta U + 0$

OR $\Delta U = 0$

OR $U_2 = U_1$

The above result proves the law of conservation of energy i.e. the internal energy of an isolated system cannot be changed by any process taking within the system.

2) Internal Energy:

The sum of all types of kinetic and potential energy of all molecules within a system is known as internal energy of the system, U.

Properties:

- The internal energy of a system determine by temperature, random motion of molecule and phase of matter.
- The internal energy of system can increase by doing work and adding heat into the system.
- The internal energy of system can decrease by losing heat to its surrounding.

Internal Energy and Temperature:

The internal energy of an object is intrinsically related to its temperature. The change in internal energy of an ideal gas is equal to:

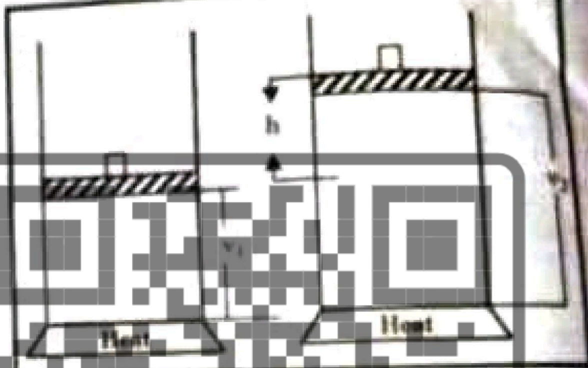
$$\Delta U = \frac{3}{2} K \Delta T$$

OR $\Delta U = (\text{constant})\, \Delta T$

OR $\Delta U \propto \Delta T$

16.2- Work Done in Thermodynamics (Work done by a Gas):

Consider an ideal or perfect gas, enclosed in a cylinder, fitted with frictionless and movable piston of cross-sectional area "A". A constant pressure on the gas is applied by placing a load on the piston. Let the initial volume of the gas is V_1, and the cylinder is placed on a heat reservoir and ΔQ amount of heat flows into the system and the gas expands by pushing the piston upward through a displacement hY and the Volume of the gas becomes V_2.



as $P = \frac{F}{A}$

OR $F = PA$

Work done by the gas is ΔW = Force x Displacement

$\Delta W = P(Ah)$

Where $Ah = \Delta V$, the change in volume of gas

$\therefore$ $\boxed{\Delta W = P\Delta V}$ ------------(1)

Where ΔV is increase in volume of gas at constant pressure.

16.3- First Law of Thermodynamics:

The first law of thermodynamics is merely the statement of the law of conservation of energy when it is stated with reference to heat energy and mechanical energy. It can be stated in the following ways.

"When heat energy is transformed into mechanical energy or when other mechanical energy is converted into heat, the total amount of energy remains constant".

OR

"The change in internal energy of a system equals to the difference between net amount of heat and work done" mathematically it can written as:

$$\Delta U = \Delta Q - \Delta W$$

OR

$$\Delta Q = \Delta U + \Delta W$$

Sign Conventions:

(i) ΔQ is taken as positive when heat enters the system and negative when heat leaves the system.

(ii) ΔW is taken positive when work is done by the system and negative when work is done on the system.

(iii) ΔU is taken positive when internal energy of the system increases and negative when it decreases.

Cyclic Process:

A cyclic process is the one which starts and ends up at the same state, i.e. the system finally attains its initial state:

$$U_2 = U_1$$

$$U_2 - U_1 = 0$$

OR $$\Delta U = 0$$

From the first law of Thermodynamics,

$$\Delta Q = \Delta U + \Delta W$$

OR $$\Delta Q = 0 + \Delta W$$

OR $$\Delta Q = \Delta W$$

It means that the work obtained from a cyclic system can be at maximum equal to the energy supplied to it and no machine in any number of cycles can perform more work than the energy gained by the machine. A perpetual motion machine of the first kind was the concept of an imaginary machine which could do more work than the energy gained by it. The above result of the first law proves that it is impossible to construct such a machine.

16.4- Applications of the First Law of Thermodynamics:

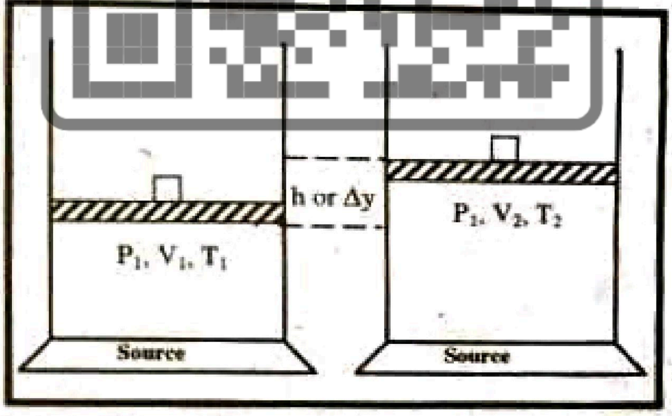
(1) ISOBARIC PROCESS:

Definition:

The thermodynamic process during which the pressure is kept constant is called an isobaric process.

Explanation:

Consider an ideal gas system enclosed in a cylinder provided with a frictionless piston. Let the system be placed on a heat reservoir and let ΔQ be the amount of heat supplied to the system, due to which the kinetic energy of gas molecules increases which increases the internal energy of the system from U_1 to U_2.



$$\Delta U = U_2 - U_1$$

also, the piston of the cylinder moves upward which means change in the volume of the system from V_1 to V_2 but pressure of system remains constant.

$$\Delta V = V_2 - V_1$$

Hence, some work is said to be done by the system against constant pressure

$$\Delta W = \text{Force} \times \text{Displacement}$$

$$\Delta W = F \times \Delta y$$

$$\because \; F = PA$$

$$\Delta W = PA\Delta y$$

$$\because \; A\Delta y = \Delta V = V_2 - V_1$$

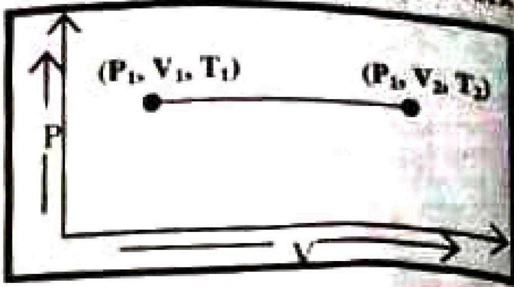
$$\Delta W = P\Delta V$$

$$\Delta W = P(V_2 - V_1)$$

According to the first Law of Thermodynamics,

$$\Delta Q = \Delta U + \Delta W \qquad \textbf{OR} \qquad \Delta Q = \Delta U + P(V_2 - V_1)$$

The above result shows that in an isobaric process, all heat energy given to the system is utilized in two ways one in increasing the internal energy of the system and the second in doing some work against external pressure.



Graphical Representation:

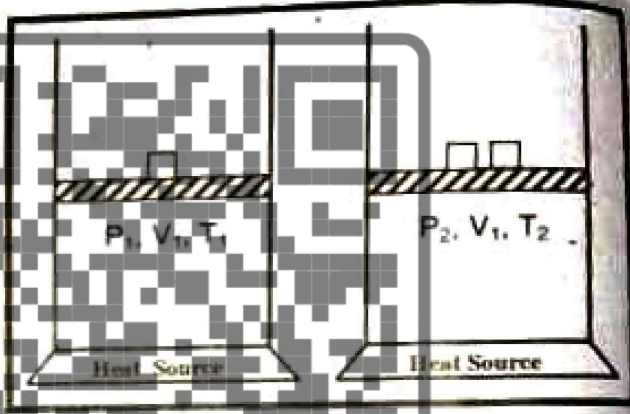
On PV–diagram, the graph of an isobaric process is a straight horizontal line called an ***Isobar***.

(2) ISOCHORIC PROCESS:

Definition:

Thermodynamic process is which volume of the system remains constant is called isochoric process.

Explanation:



Consider a system of an ideal gas in a cylinder provided with a piston, which is fixed. Let the system be placed on a heat reservoir and ΔQ be the amount of heat supplied to the system which increases the kinetic energy of gas molecules and hence, the internal energy of the system changes from U_1 to U_2

$$\Delta U = U_2 - U_1$$

As the piston cannot move, there will be no change in the volume of the system

$$\Delta V = 0$$

Hence, there will be no work done by the system

$$\Delta W = 0$$

According to the first law of thermodynamics,

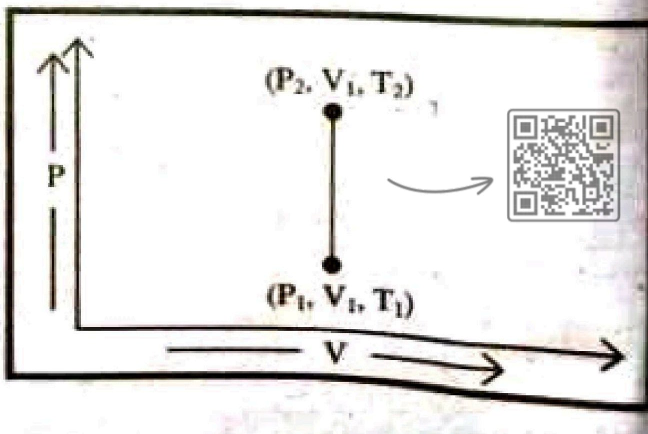
$$\Delta Q = \Delta U + \Delta W$$

OR

$$\Delta Q = \Delta U + O$$

$$\Delta Q = \Delta U$$

The above result shows that in an isochoric process, the heat energy given to the system does nothing, but only changes the internal energy of the system.



Graphical Representation:

On P–V diagram, the graph of an Isochoric Process is a vertical straight line called an ***Isochor***.

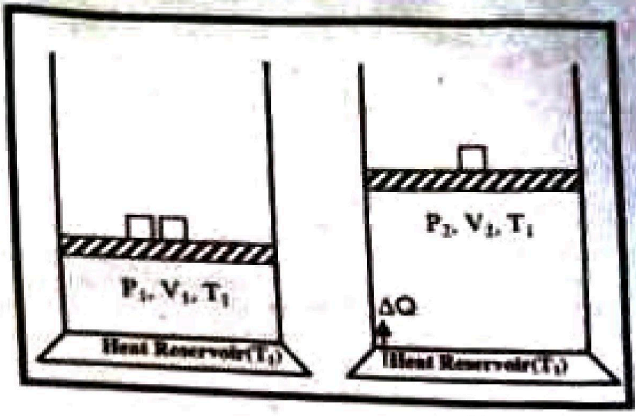
(3) ISOTHERMAL PROCESS:

Definition:

The thermodynamic process which is carried out in such a way that system undergoes but its temperature remains constant is called isothermal process.

Explanation:

Consider a system of a gas in a cylinder, which is provided with a frictionless movable piston. Walls of the cylinder and piston are ideally heat-insulating and its base is ideally heat-conducting. The cylinder is placed on a heat reservoir at a temperature T_1. The gas is allowed to expand by decreasing the load on the piston and the temperature of system is maintained by supplying some heat energy to the system from heat reservoir. Such an expansion is called **Isothermal Expansion.**



As $U_2 = U_1$

=> $\Delta U = 0$

From the first law of thermodynamics

$$\Delta Q = \Delta U + \Delta W$$

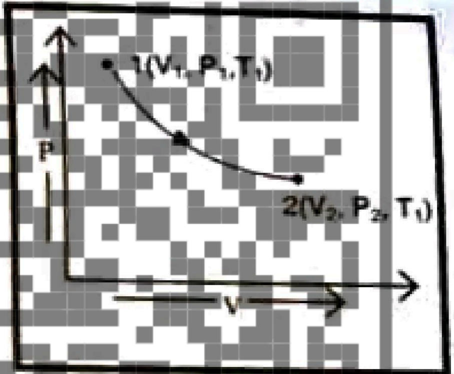
$$\Delta Q = 0 + \Delta W$$

$$\Delta Q = \Delta W$$

The above result shows that, in an isothermal process, the heat energy given to the system is converted all into doing some work.

Graphical Representation:

On P-V diagram, the graph of an isothermal process is a smooth curve called an isotherm.



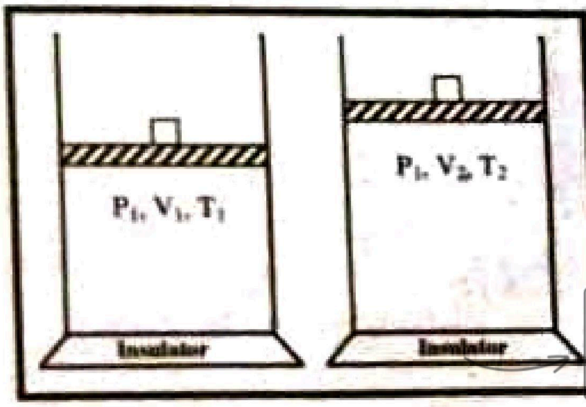
(4) ADIABATIC PROCESS:

Definition:

The thermodynamic process during which no heat enters or leaves the systems is called an adiabatic process.

Explanation:

Consider a system of a gas, in a cylinder, provided with a movable frictionless piston. Let the system be initially at temperature T_1. Now, let the system be placed on insulator while the gas expands and cools off adiabatically and its temperature falls to T_2. Thus, some internal energy of the system is converted in work done.



Since $\Delta Q = 0$

From the first law of Thermodynamics,

$$\Delta Q = \Delta U + \Delta W$$

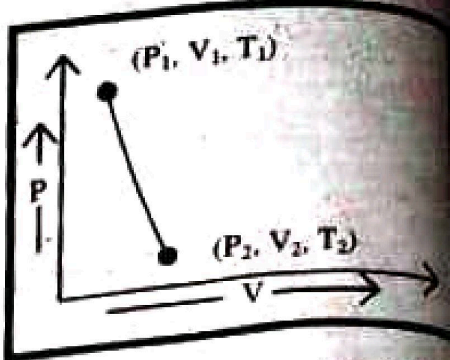
$$0 = \Delta U + \Delta W$$

OR $$-\Delta U = \Delta W$$

The above equation shows that in an adiabatic process, work is done at the cost of the internal energy of the system. This process is called adiabatic expansion.

Graphical Representation:

On P-V diagram, the graph of an adiabatic process is a smooth curve, called **adiabatic curve or adiabate**.



16.5- Heat Capacity:

All substances naturally possess the ability of absorbing some heat. This natural ability of substances is called **Heat Capacity**. Heat capacity of a substance depends upon many factors, such as nature, mass, temperature etc. Upon the absorption of heat, either the temperature of the substance changes or state of the substance changes. The quantity of heat which changes the temperature of substance, is called **Specific Heat** and the quantity of the heat which changes the state of substance, is called **Latent Heat**. When the temperature of the substance changes, its state remains the same and when the state of the substance changes, its temperature remains the same.

Specific Heat Capacity:

All substances absorb heat when heated through some range of temperature. This ability of the objects to absorb heat depends upon their masses their nature, called **Specific Heat Capacity.** So specific heat capacity of a substance is defined as ***"The amount of heat required to raise the temperature of unit mass of a substance through one Kelvin or one degree Celsius"***.

If a substance of mass "m" is heated through a small range of temperature ΔT, such that ΔQ is the amount of heat absorbed, then specific heat capacity of the substance is given by:

$$C = \frac{\Delta Q}{m\Delta T}$$

S.I. Unit of Specific Heat Capacity is $J.Kg^{-1}K^{-1}$.

Molar Specific Heat Capacity:

As we know, Specific Heat capacity of a substance depends upon the mass of the substance in addition to the nature of the substance, which is given by:

$$C = \frac{\Delta Q}{m\Delta T} \quad \text{------------------------ (1)}$$

But, for gases, their heat capacity is independent of the mass of the gas, therefore, if "n" is the no. of moles of a gas and "M" is the molecular mass of the gas, then

$$\textbf{No. of moles of Gas} = \frac{\text{mass of gas}}{\text{Molecular mass of gas}}$$

$$n = \frac{m}{M}$$

OR $\quad m = nM$

$$\therefore \text{Eq.(1)} \Rightarrow \quad C = \frac{\Delta Q}{nM\Delta T}$$

$$MC = \frac{\Delta Q}{n\Delta T} \quad \text{.......................} (2)$$

Where MC is the product of molecular mass of the gas and the specific heat capacity of the gas, which is called ***the molar specific heat capacity*** of a gas, which is given by:

$$MC = C'$$

$$\text{Eq.(2)} \Rightarrow \quad C' = \frac{\Delta Q}{n\Delta T}$$

Thus, molar specific heat capacity of a gas can be defined as ***"the amount of heat required to raise the temperature of one mole of a gas through one Kelvin or one degree Celsius"***

The S.I. unit of molar specific heat capacity is $J.mol^{-1}.K^{-1}$.

Molar specific heat capacity is of two types which are:

(i) Molar specific heat capacity at constant pressure (C_p)

It is defined as *"the amount of heat required to rise the temperature of one mole of a gas through one Kelvin at constant pressure."*

$$C_p = \frac{\Delta Q_p}{n\Delta T}$$

(ii) Molar specific heat capacity at constant volume (C_v).

It is defined as *"the amount of heat required to raise the temperature of one mole of a gas through one Kelvin at constant volume."*

$$C_v = \frac{\Delta Q_V}{n\Delta T}$$

Latent Heat:

Latent heat of a substance is defined as *the amount of heat required to change the state of unit mass of a substance*. Mathematically,

$$H = \frac{\Delta Q}{m}$$

$$OR \quad \boxed{\Delta Q = mH}$$

S.I. unit of Latent Heat is J/Kg.

Latent heat is of two types:

1) Latent Heat of Fusion:

Latent heat of fusion is ***the amount of heat required to melt unit mass of a solid***. $H_f = \frac{\Delta Q}{m}$

$$OR \quad \boxed{\Delta Q = mH_f}$$

2) Latent Heat of Vaporization:

Latent heat of Vaporization is ***the amount of heat required to vopourize unit mass of a liquid.***

$$H_v = \frac{\Delta Q}{m}$$

$$OR \quad \boxed{\Delta Q = mHv}$$

16.6- Q. Prove that the Difference of Molar Specific Heat Capacity at Constant Pressure and that at Constant Volume is Equal to the Universal Gas Constant:

OR

Prove that $C_p - C_v = R$:

Proof:

Let us consider "n" moles of an ideal or perfect gas, enclosed in a cylinder fitted with a frictionless movable piston. The cylinder is placed on a heat reservoir, the temperature of which can be raised or lowered as needed. From kinetic molecular theory, molecules of ideal gas always remain in random motion and collide elastically with the walls of cylinder and with its piston, thereby exerting pressure on the piston which is balanced by the weight of the piston and the load placed on the piston.

At Constant Pressure:

If an ideal gas is heated at constant pressure so that its temperature rises by ΔT, then heat transferred ΔQ_P is given by:

$$\Delta Q_P = nC_P\Delta T$$

Since the gas expands to keep the pressure constant so the work done by the gas is:

$$\Delta W = P\Delta V$$

As we know that in isobaric process the amount of heat energy supplied to the system is used for the increase the enthalpy of the system that is:

$$\Delta Q_P = \Delta H$$

But from 1st Law of Thermodynamics

$$\Delta H = \Delta U + \Delta W$$

$$\therefore \quad \Delta Q_P = \Delta U + \Delta W$$

$$OR \quad nC_P\Delta T = \Delta U + P\Delta V \quad \text{----------- (1)}$$

According to general gas law:

$$PV = nRT$$

At Constant pressure P, amount of the work done due to expansion ΔV caused by the rise in temperature ΔT is given by

$$P\Delta V = nR\Delta T$$

Now eq (1) can be written as:

$$nC_P\Delta T = \Delta U + nR\Delta T \quad \text{------------ (2)}$$

At Constant Volume:

If an ideal gas is heated at constant volume so that its temperature rises by ΔT then heat transferred ΔQ_V is given by:

$$\Delta Q_V = nC_V\Delta T$$

According to first law of thermodynamics

$$\Delta Q_V = \Delta U + \Delta W$$

$$nC_V\Delta T = \Delta U + \Delta W$$

since volume remains constant therefore work done by the system is zero ($\Delta W = 0$)

hence we can write

$$nC_V\Delta T = \Delta U + 0$$

$$\Delta U = nC_V\Delta T$$

Put it in eq (2)

$$nC_P\Delta T = nC_V\Delta T + nR\Delta T$$

for 1 mol of gas put n = 1 therefore,

$$C_P\Delta T = C_V\Delta T + R\Delta T$$

$$C_P\Delta T = \Delta T(C_V + R)$$

$$C_P = C_V + R$$

OR

$$\boxed{C_P - C_V = R}$$

Proved