

12.5.1 Applications of First Law of Thermodynamics

In this section we briefly explain different thermodynamic processes.

i. Isothermal Process

A thermodynamic process in which temperature remains constant is called **isothermal process**.

For an ideal gas, internal energy U is also constant in isothermal process.

$$\text{Since } \Delta T = 0, \Delta U = 0$$

$$\text{Hence from 1st law, } \Delta Q = \Delta W$$

ie., the heat supplied to a system is completely utilised to do work.

ii. Isochoric Process

In this process, the volume of the system is held constant. Hence the system cannot do any work.

$$\text{ie., } \Delta W = 0, \text{ Hence } \Delta Q = \Delta U$$

This means that if heat is supplied to a system, its internal energy increases, but does not perform any work in an isochoric change.

iii. Isobaric Process

In the process, the pressure of the system remains constant. Hence when heat is supplied to the system, internal energy increases and the system performs work.

$$\text{Also } \therefore \Delta Q = \Delta U + \Delta W$$

iv. Adiabatic Process

A thermodynamic process in which no heat enters or leaves the system is known as **adiabatic process**.

Since no transfer of energy on heat occurs

$$\therefore \Delta Q = 0, \text{ Hence } \Delta U = -\Delta W$$

This shows that when a gas expands adiabatically, ΔW is positive and

hence ΔU is negative. It means that internal energy of the system should decrease. Then the gas gets cooled, so adiabatic expansion produces cooling effect.

v. Melting Process

In the melting process, internal energy increases. Amount of heat required for melting is $\Delta Q = mL$, where 'm' is the mass of the substance and 'L' its latent heat.

When a solid changes into liquid, there is only a very small change in volume, and hence it can be neglected. i.e., $\Delta V = 0$ and hence $\Delta W = 0$. [since $\Delta W = P\Delta V$]

$$\therefore \Delta Q = \Delta U \quad \text{and} \quad \therefore \Delta U = mL$$

vi. Boiling Process

During boiling, a liquid changes to vapour at constant temperature and pressure. There is a change in volume, say from V_1 to V_2 .

The heat supplied, $\Delta Q = mL$, m - mass of liquid and L its latent heat of vapourisation.

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

$\Delta U = \Delta Q - \Delta W$, by 1st law of thermodynamics.

$$\Delta U = mL - P(V_2 - V_1)$$

12.8.2 Indicator Diagram or PV Diagram

The graphical representation of the variation of pressure with the variation of volume of a thermodynamic system is called an indicator diagram. For this V is taken along the X- axis and P along the Y- axis. A typical graph of ideal gas is shown in fig. It is a hyperbola.

12.8.3 Cyclic Process

A **Cyclic process** is one which consists of a series of changes through which system passes and finally returns to its initial state.

For a cyclic process, $\Delta U = 0$ and hence $\Delta Q = \Delta W$.

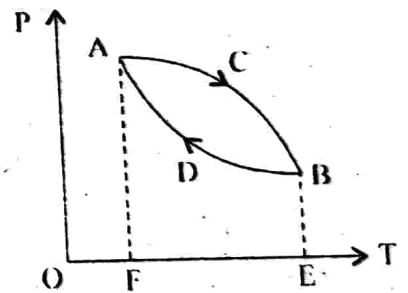


Fig. 7

12.8.3a Work done in a Cyclic Process

Let P_1 and V_1 be the initial pressure and initial volume of a gas (represented by A). When the gas expands, let P_2 and V_2 be its final pressure and volume (represented by B). Thus work done by the gas = Area ACBEF.

Now the system (gas) is taken back to A from B, then work done on the gas = - Area BDAFE. The negative sign shows that the process is in the reverse direction.

Total work done, $W = \text{Area ACBEF} + (- \text{Area BDAFE}) = \text{Area ABCD}$.

In this case the expansion curve ACB lies above the compression curve BDA and so area is positive. Hence work is done by the system. Suppose if the compression curves lies above the expansion curve, then area is negative. In this case work is done on the system.

12.8.4 Condition for Isothermal and Adiabatic Change

i. Isothermal change

- Walls of the container in which the gas is taken must be perfectly conducting, so that there is a free exchange of heat between the gas enclosed in the container and its surroundings.
- The process of expansion or compression should be slow, so that exchange of heat can take place. Equation of an isothermal change for a perfect gas is $PV = a \text{ constant}$ (Boyle's law).

ii. Adiabatic change

- The walls of the container in which a gas is enclosed must be perfectly non conducting, so there is no transfer of heat from the container to its surroundings and vice versa.
- The process of expansion or compression should be so quick, so that there is no time for heat exchange between the gas and the container.

Adiabatic equation for a perfect gas is,

$PV^\gamma = a \text{ constant}$, and γ - is the ratio of specific heat.

Other expressions are $TV^{\gamma-1} = a \text{ constant}$ and $P^{1-\gamma} T^\gamma = a \text{ constant}$

You must know

$$\text{Specific heat, } s = \frac{\Delta Q}{m \Delta T}$$

i. For isothermal change, $\Delta T = 0 \therefore s = \text{infinity}$

ii. For adiabatic change, $\Delta Q = 0 \therefore s = 0$

12.8.6 Expression For Work done in Isothermal and Adiabatic Processes

The work done by a gas for an infinitesimal change of volume dV against the constant external pressure is, $\Delta W = PdV$

If the volume is changing from V_1 to V_2 , then the net work done by the

system is $W = \int_{V_1}^{V_2} PdV$

i. Isothermal change

For an isothermal change, temperature T is constant. We have, $PV = \mu RT$

$$\therefore W = \int_{V_1}^{V_2} \frac{\mu RT}{V} dV = \mu RT \int_{V_1}^{V_2} \frac{1}{V} dV = \mu RT [\log_e V]_{V_1}^{V_2} = 2.303 \mu RT [\log V_2 - \log V_1]$$

$$W = 2.303 \mu RT \log \left(\frac{V_2}{V_1} \right)$$

Since $P_1 V_1 = P_2 V_2$ and $\left(\frac{V_2}{V_1} = \frac{P_1}{P_2} \right)$

$$W = 2.303 \mu RT \log \left(\frac{P_1}{P_2} \right)$$

ii. Adiabatic change

The total work done 'W' is, $W = \int_{V_1}^{V_2} PdV = \int_{V_1}^{V_2} \frac{K}{V^\gamma} dV$ [since $PV^\gamma = K$, a constant]

$$= K \left[\frac{V^{-\gamma+1}}{-\gamma+1} \right]_{V_1}^{V_2} = \frac{K}{1-\gamma} [V_2^{1-\gamma} - V_1^{1-\gamma}] = \frac{K}{\gamma-1} [V_1^{1-\gamma} - V_2^{1-\gamma}] = \frac{K}{\gamma-1} \left[\frac{V_1}{V_1^\gamma} - \frac{V_2}{V_2^\gamma} \right]$$

$$= \frac{1}{\gamma-1} \left[\frac{P_1 V_1^\gamma V_1}{V_1^\gamma} - \frac{P_2 V_2^\gamma V_2}{V_2^\gamma} \right] \quad [\text{since } P_1 V_1^\gamma = P_2 V_2^\gamma = K]$$

$$\therefore W = \frac{1}{\gamma-1} [P_1 V_1 - P_2 V_2]$$

Let T_1 and T_2 be the initial and final temperature. We have the standard gas equation $PV = RT$. Then $P_1 V_1 = RT_1$ and $P_2 V_2 = RT_2$

$$\therefore W = \frac{R}{\gamma-1} (T_1 - T_2)$$

The work done in an isothermal or adiabatic process can be found out from the indicator diagram.

In the figure, the point A represents initial state of a system and B, the final state. The total work done by the gas during the expansion process (curve AB) is

$$W = \int_{V_1}^{V_2} PdV$$

Area under the curve is ABCD. Let us divide this into a number of small strips. Let PQRS be one such small strip.

$$\text{Area of PQRS} = PS \times SR = P \Delta V$$

$$\text{Hence, } W = \int_{V_1}^{V_2} PdV = \int_D^C \text{Area of strip PQRS}$$

$$\text{ie., } W = \text{Area ABCD}$$

Thus the work done by a gas in isothermal and adiabatic process is equal to the area under the curve.

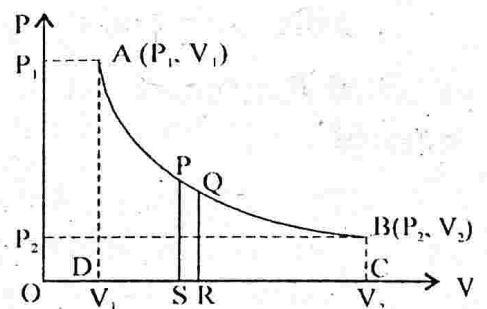


Fig. 9