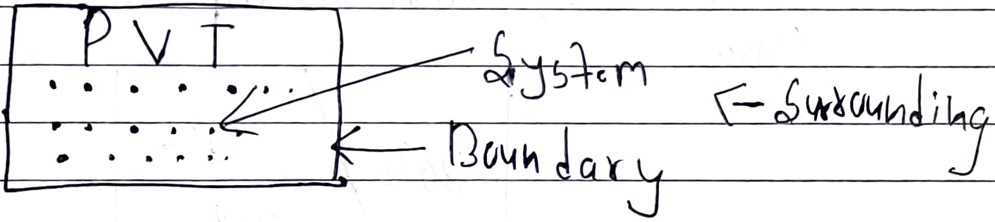
First law Of THERMODYNAMICS

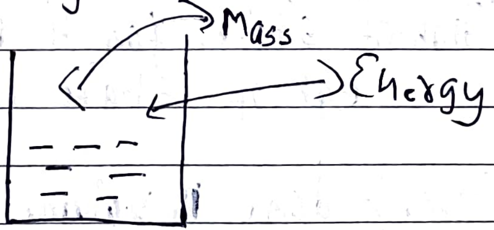
- Thermodynamics: Thermodynamics deals with the study of relationship involving heat, mechanical work & other aspects of energy & energy transfer.

- Thermodynamics System: ~~At~~ Definite quantity of matt bounded by some closed surface & having a certain pressure, volume & tempr is called thermodynamics syste

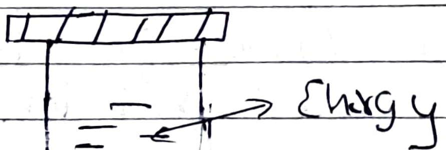


There are three types of thermodynamic system:

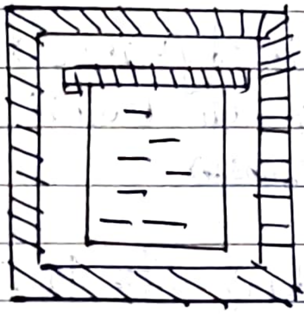
ⓐ Open System: It system can exchange both mass & energy with its surrounding, such system is called open system.



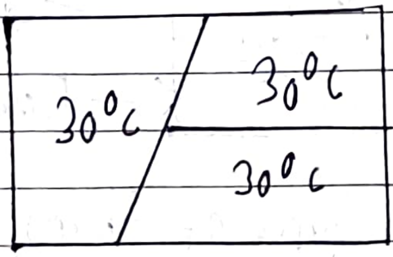
ⓑ Closed System: It a system can exchange only energy but not the mass with its, surrounding such system is called closed system



(c) Isolated System: If a system can neither exchange mass nor energy with its surrounding such system is called isolated system.



(d) Thermal Equilibrium: A thermodynamic system is said to be in thermal equilibrium if temp^{r} of different parts remain same.



② Equation of state: The general relationship bet^{n} pressure, volume & temp^{r} of a sample system is called equation of state. for n mole of ideal gas,
eqn of state is $PV = nRT$

Work done in volume changes

-> If the volume of gas increases, work done is +ve In this case work done by the gas.

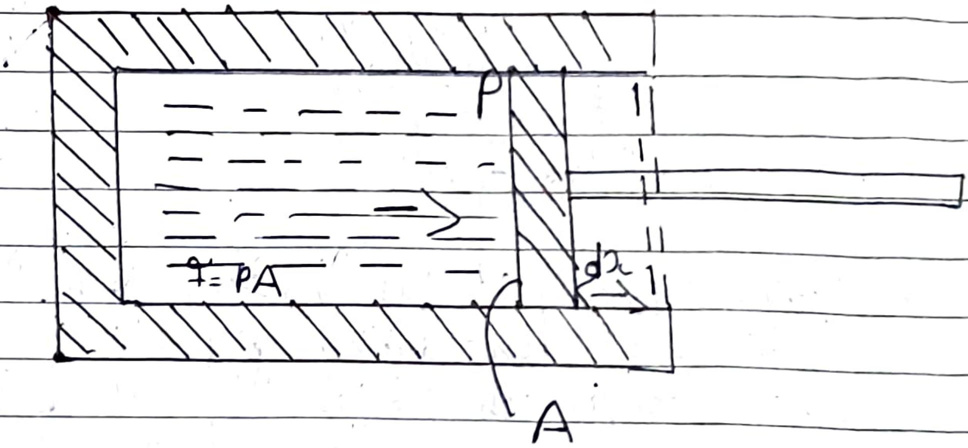
-> If the vol of gas decreases, work done is negative in this case work is done on the gas

-> If the vol of gas is constant, work done is 0.

* Work done by the gas during expansion

→ Let us consider a gas is filled in a cylinder filled with a frictionless & movcable piston. Let P be the pressure exerted by gas & A be the CSA of the piston. The force exerted by the gas on the piston is given by

$f = P.A$ — ① $\left(P = \frac{f}{A}\right)$



During expansion of gas, the piston displaces through dx & a small amount of work ~~done~~ dw is done & given by

$$dw = f.dx$$

$$dw = P.A\,dx$$ (from ①)

Where $A\,dx = dv$ is the small increase in the vol of the gas. Then the work done by the gas is given by

$$dw = P.dv$$ — (ii)

The total work done by the gas during expansion is obtained by integrating eqn (ii) with the limit V_1 to V_2.

i.e,

$$W = \int dw = \int_{V_1}^{V_2} P \cdot dV$$

$$= P \int_{V_2}^{V_2} dV \quad (\because P \text{ is constant})$$

$$= P[V]_{V_1}^{V_2}$$

$$= P(V_2 - V_1)$$

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

-> During expansion of gas $V_2 > V_1$ & so $V_2 - V_1 = +ve$. So the work done during expansion is +ve.

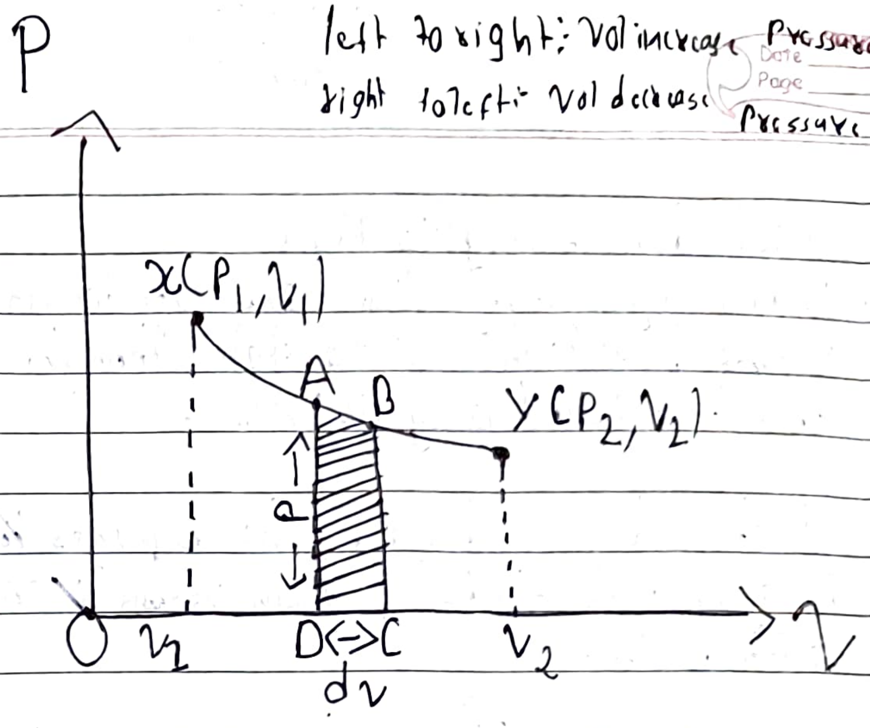
-> During compression of gas $V_2 < V_1$ & so $V_2 - V_1 = -ve$. So the work done during compression is -ve.

-> When the volume of gas is constant, $V_2 = V_1$ & so $V_2 - V_1 = 0$. So the work done is 0 when the vol is constant.

Graphical representation of work done or indicator diagram or P-V diagram.

-> When we plot the pressure & volume of gas in graph, it is called indicator diagram or P-V diagram.

left to right: vol increase, Pressure decrease
right to left: vol decrease, Pressure increase



When an ideal gas having pressure P_1 & Volume V_1 expands to Volume V_2 pressure becomes P_2, a curve XY is obtained which is as shown in ~~way~~ the graph. let us take two points A & B on the curve XY very close to each other in which pressure is almost constant.

i.e P & Volume increases by dV.

Small work done (dw) = P. dV

= Area of strip ABCD —①

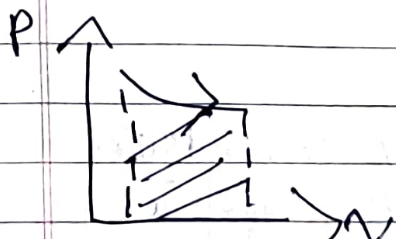
The total work done by the gas during expansion from X to Y is obtained by integrating eqn ① within the limit V_1 to V_2.

i.e, $W = \int dw$

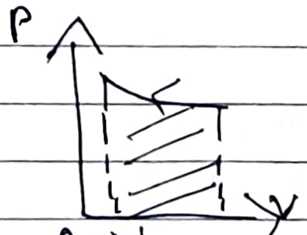
$= \int_{V_1}^{V_2}$ Area of strip ABCD

= Area under the curve XY.

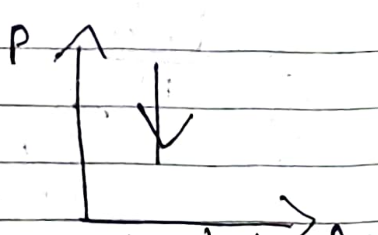
∴ work done = Area under the curve



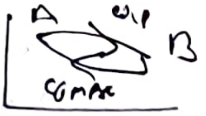
V -> increase
W = +ve



V -> decrease
W = -ve



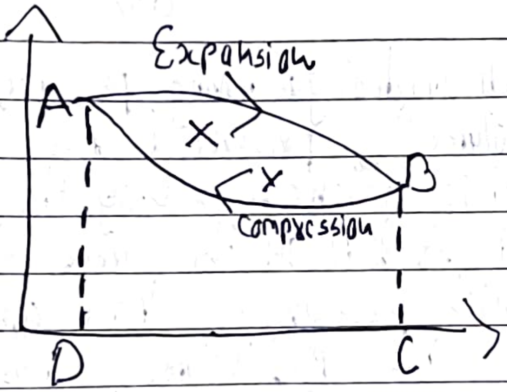
V -> constant
W = 0



Work done in a cyclic process

→ When a gas changes from one state to another & again returns back to the original state through another path, then it is called a cyclic process.

In the indicator diagram, Gas expands from state A to B along path AXB & again returns to original state A along path BYA.



work done during expansion

W_1 = Area AXBCDA

work done during compression

W_2 = Area BYADCB

Net work done in a cyclic process is given by

$W = W_1 - W_2$

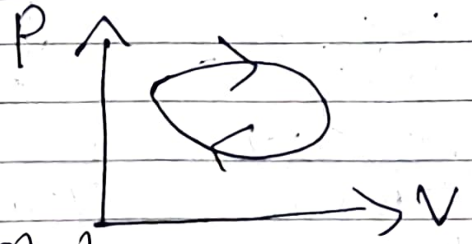
= Area AXBCDA − Area BYADCB

= Area of closed loop AXBYA

∴ work done in a cyclic process = Area of closed loop.

Note

①



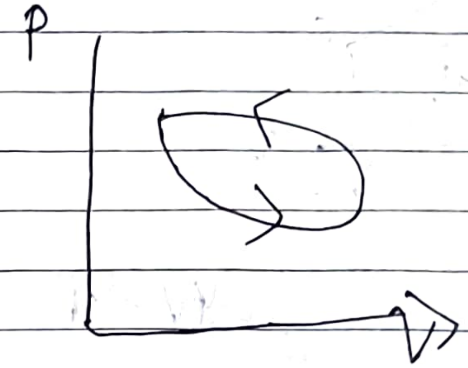
+ve work more

−ve work less

net work done = +ve

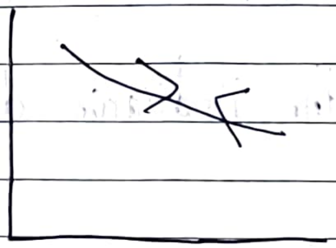
Clockwise

② Anticlock wise



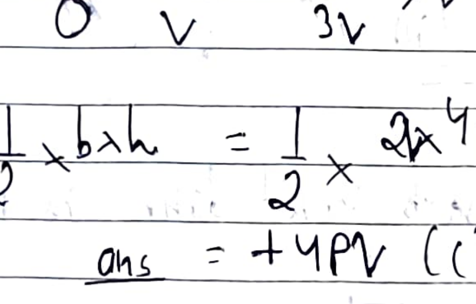
+ve work less
-ve work more
net work done = -ve

③



net work done = 0
Practically not possible
Reversible process

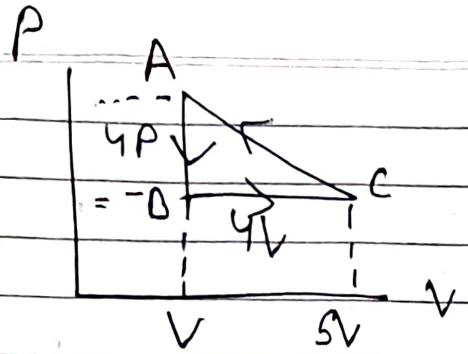
MCQ Calculate the work done from the indicator diagram.



$W = A = \frac{1}{2} \times b \times h = \frac{1}{2} \times 2v \times 4p$

ans $= +4PV$ (clockwise)

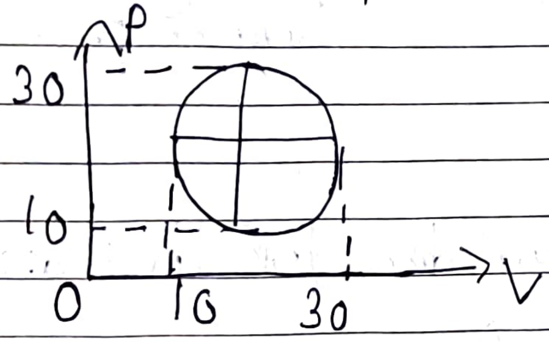
② Calculate the work done from the indicator diagram.



$$W = A = \frac{1}{2} \times b \times h = \frac{1}{2} \times 4V \times 4P$$

$-8PV$ (anticlockwise)

③ Calculate the work done from the indicator diagram.



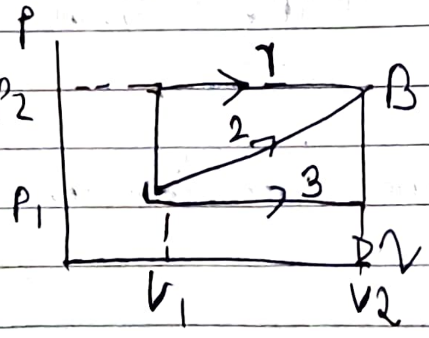
$d = 20, r = 10$

$W = A = \pi r^2$

$\pi \times 10^2$

100π

④ Along which path work done is maximum?



→ Area under the path ① is maximum

- **Internal Energy of a gas (U)**

→ The sum of P.E & K.E of all the molecules of a gas is called its internal Energy. It is denoted by U

i.e

$$\left[U = U_{P.E} + U_{K.E}\right]$$

According to kinetic theory of gas, the molecules of ideal gas donot exert intermolecular force of attraction. Hence the molecules of ideal gas donot posses P.E. As a result the internal energy of ideal gas is wholly kinetic in nature & is a function of tempr only

i.e $U = f(T)$ (for ideal gas)

However, the internal energy of real gas is partly due to P.E & partly due to K.E & is a function of vol & tempr

i.e $U = f(V, T)$ (for real gas)

MCQ 5 The internal energy of ideal gas depends on?

(a) Vol [✓ (b) temp] (c) both (d) none

(6) The internal energy of real gas depends on?

(a) Vol (b) temp [✓ (c) both] (d) none

Imp

First law of thermodynamics

Stmt :- "If a certain quantity of heat is supplied to a system, a part of it is used to do the external work on expansion & the remaining heat is used to increase the internal energy of the system."

If dQ be the amount of heat supplied to the system, du be the increase in internal energy & dw be the external work done, then from 1st law of thermodynamics,

$$dQ = du + dw$$

It is based on the principle of conservation of energy.

Sign notation

⇒ If a system gains heat, $dQ = +ve$,
If a system loses heat, $dQ = -ve$,

If a gas is heated at constant vol, $dV = 0$

From the 1st law of thermodynamics,

$$dQ = dU + dw$$

$$\text{or, } dQ = dU + P.dV$$

$$\text{or, } dQ = dU + 0 \qquad (dV = 0)$$

$$dQ = dU$$

Heat supplied is used to increase the internal energy only.

During a cyclic process, change in internal energy is 0 because the system is brought back to the initial state. i.e $dU = 0$

From 1st law of thermodynamics.

$dQ = dU + dw$

$dQ = 0 + dw \quad (dU = 0)$

$dQ = dw$

Heat supplied is used to do the external work only.

Heat Capacity of gas.

ⓐ Specific heat capacity at constant volume c_v smaller

→ It can be defined as amount of heat required to raise the temp^r of 1kg of gas by 1°C or 1K at constant volume

i.e. $c_v = \frac{dQ}{mdT}$

S.I unit $J\,kg^{-1}K^{-1}$

ⓑ Specific heat capacity at constant pressure c_p smaller

→ It can be defined as amount of heat required to raise the ~~th~~ temp^r of 1kg of gas by 1°C or 1K at constant pressure.

$c_p = \frac{dQ}{mdT}$

S.I = $J\,kg^{-1}K^{-1}$

ⓒ Molar heat capacity at constant Vol C_v capital

→ It can be defn as amount of heat reqd to raise the temp^r of 1mole of gas by 1°C or 1K at constant Volume

i.e. $C_v = \frac{dQ}{ndT}$

S.I unit = $J\,mol^{-1}K^{-1}$

ⓓ Molar heat capacity at constant pressure C_p.

→ It can be defined as amount of heat required to raise the tempr of 1mole of gas by 1°C or 1K at constant pressure.

i.e $C_p = \frac{dQ}{n\,dT}$

$J\,mol^{-1}K^{-1}$

$C_p >> C_v$

Relation between cp & Cp

we have,

$$C_p = \frac{dQ}{m\,dT} \quad ①$$

also,

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

$$m = n \times M$$

Putting in ①

$$C_p = \frac{dQ}{nM\,dT}$$

$$Mc_p = \frac{dQ}{n\,dT}$$

$$Mc_p = C_p$$

Ily, $M.c_v = C_V$

VVI Relation between Cp & Cv or

$$C_p - C_v = R$$

→ Let us consider n mole of an ideal gas having pressure P, Volume V & tempr T.

Suppose gas is heated at const vol, so that tempr increases by dT.

Heat supplied to the gas is given by

$$dQ = nC_v \cdot dT \qquad \left(C_v = \frac{dQ}{ndT}\right)$$

$$\& \; dW = P \cdot dV = 0 \quad (dV = 0)$$

From 1st law,

$$dQ = dU + dW$$

$$nC_v \cdot dT = dU + 0 \quad \text{— (1)}$$

$$\therefore dU = nC_v \cdot dT$$

Again the gas is heated at constant pressure so that vol increases by dV & temp^r increases by dT.

Heat supplied to the gas is given by,

$$dQ = nC_p \cdot dT \qquad \left(C_p = \frac{dQ}{ndT}\right)$$

$$\& \; dW = P \cdot dV$$

From 1st law,

$$dQ = dU + dW$$

$$nC_p \cdot dT = dU + PdV$$

$$nC_p dT = nC_v dT + PdV \quad \text{— (ii)}$$

For n mole of an ideal gas,

$$PV = nRT$$

Differentiating both sides we get,

$$dPV = dnRT$$

$$PdV = nRdT$$

Putting in eqn (1)

$$nC_p dT = nC_v dT + nRdT$$

$$C_p = C_v + R$$

$$C_p - C_v = R \quad \text{Proved}$$

Relation between C_p & C_v

we have,

$$C_p - C_v = R$$

$$\frac{C_p}{M} - \frac{C_v}{M} = \frac{R}{M}$$

$$\boxed{c_p - c_v = r}$$

ⓐ Why a gas has two heat capacities OR $C_p >> C_V$

→ When a gas is heated at constant vol, heat supplied is used to increase the internal energy only. But when the gas is heated at constant pressure heat supplied is used to increase the internal energy as well as to do the external work. So, the heat required to raise a certain temp^r at constant pressure is greater than the heat required to raise the temp^r at constant vol. That is why Cp is greater than Cv.

vol constant C_V	Pressure Constant C_P
$dQ = dU + \cancel{dw} \to 0$	$dQ = dU + dw$
$dQ = dU$	$C_P = \frac{dQ}{n\,dT} = \frac{dU + dw}{n\,dT}$
$C_V = \frac{dQ}{n\,dT} = \frac{dU}{n\,dT}$.	

MCQ When water is heated from 0°C to 4°C then

→ $C_p \leq C_V$

Thermodynamic process.

→ A thermodynamic process is said to take place if some change occurs in the state of thermodynamic system.

The diff thermodynamic process are

① Isothermal process (T → const)
② Adiabatic process (Q → const)
③ Iso choric process (V → const)
④ Isobaric process (P → const)

Isothermal process (T → constant)

-> The thermodynamic process in which pressure & vol changes bu
t temp^r remains constant is called isothermal process.

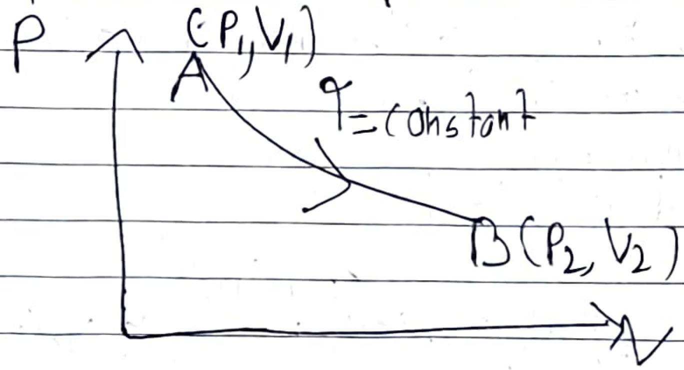
For eg: (i) Melting of ice

(ii) Boiling of water

Cond^n for process to be isothermal

1. The walls of the system should be perfectly conducting
2. The process should be slow & should be performed in infinitesimal small steps.

Eqn of isothermal process



$$P_1V_1 = P_2V_2$$

Let P, V, T be the pressure, vol & temp^r of a ideal gas, then the eqn of state for one mole of gas.

$$PV = RT$$

Where R is universal gas constant. During isothermal process T is constant so the above eqn be comes,

$$PV = \text{constant}$$

$$P_1V_1 = P_2V_2$$

Application of 1st law of thermodynamics to be isothermal process,

-> During an isothermal process temp^r remains constant. As the internal energy of ideal gas depends on temp^r. So there is no change in the internal energy of the gas.

i.e $dU = 0$

From 1st law

$$dQ = dU + dw$$

$$dQ = 0 + dw \qquad (dU = 0)$$

$$dQ = dw$$

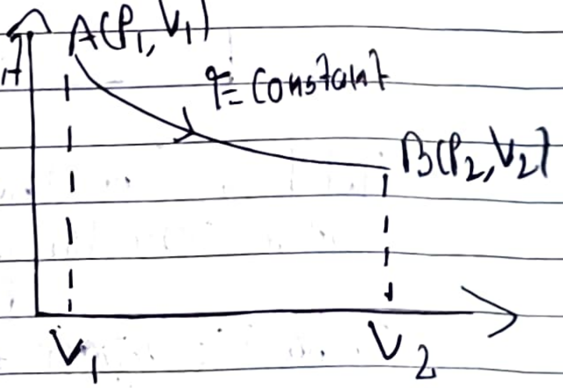
Heat supplied to the system = Work done by the system.

Work done during isothermal process

-> Let us consider n mole of an ideal gas is filled in a cylinder having perfectly conducting walls & fitted with a frictionless & movable piston. A small amount of work done dw is done for the small change in volume dV at constant pressure P is given by,

$$dw = P.dV \quad - ①$$

When the gas expands from initial state A (P_1, V_1) to the final state B (P_2, V_2) work done is given by



$$W = \int dw$$

$$\int_{V_1}^{V_2} P.dV \quad - \text{(ii)}$$

For n mole of an ideal gas

$$PV = nRT$$

$$P = \frac{nRT}{V}$$

Putting this value in eqn (ii)

$$W = \int_{V_1}^{V_2} \frac{nRT}{V} . dV$$

$$nRT \int_{V_1}^{V_2} \frac{1}{V} dV$$

$$nRT \left[\ln V\right]_{V_1}^{V_2}$$

$$nRT \left[\ln V_2 - \ln V_1\right]$$

$$nRT \times \ln\left(\frac{V_2}{V_1}\right)$$

$$W = nRT \log\left(\frac{V_2}{V_1}\right) \quad - ①$$

During isothermal process

$$P_1V_1 = P_2V_2$$

$$\frac{V_2}{V_1} = \frac{P_1}{P_2}$$

Putting this in ①

$$W = nRT\log\left(\frac{P_1}{P_2}\right)$$

Ⓘ 5 mole of an ideal gas are kept at a constant tempr of 53°C while the pressure of gas in increased from 1atm to 3 atm. Calculate the work done.

-> no of mole = 5

Tempr constant = 53°C = 326K

$P_1 = 1$ atm

$P_2 = 3$ atm

$R = 8.314\ J\,mol^{-1}K^{-1}$

$$W = nRT\log\left(\frac{P_1}{P_2}\right)$$

$$5 \times 8.314 \times 326 \times \log\left(\frac{1}{3}\right)$$

$$W = -14888.19 J$$

- **Adiabatic process (Q constant) :** The thermodynamic process in which pressure, volume & tempr changes but there is no exchange of heat between the system & the surrounding is called adiabatic process

For example:

① ~~Bursting~~ Bursting of a bicycle tyre
② Compressor etc

* Condition for a process to be adiabatic

① The walls of the system should be insulating
② The process should be fast so that there is less time for the exchange of heat between the surrounding & system

Application of 1st law of thermodynamics to the adiabatic process (VVI)

→ During adiabatic process, no heat enters or leaves the system so,

$$dQ = 0$$

From 1st law of thermodynamics

$$dQ = dU + dw$$

or, $$0 = dU + dw$$

$$dw = -dU$$

If the gas expands dw is +ve & hence dU is -ve.

The internal energy of the gas decreases & the tempr of the gas drops.

Again,

$$dw = -dU$$

If the gas compresses, dw is -ve & hence dU is +ve.

$-dw = dU$

The internal energy of the gas increases & the temp^r of the gas rises.

Ⓠ Why cooling effect is produced in the tyre of bicycle burst?

Ⓠ Why temp^r of gas drops during adiabatic expansion?

Equation of adiabatic process:

-> Let us consider n mole of an ideal gas as a system. Suppose the gas absorbs dQ amount of heat & performs dw amount of work.

From 1st law of thermodynamics,

$$dQ = dU + dw \quad - ①$$

The change in internal energy is

$$dU = nC_v dT$$

& $dw = P.dV$

During adiabatic process, $dQ = 0$

so eq^n ① becomes

$$0 = nC_v dT + P.dV$$

$$nC_v dT + PdV = 0 \quad - (ii)$$

For n mole of an ideal gas,

$$PV = nRT$$

Differentiating both sides,

$$dPV = dnRT$$

$$P\cdot dv + V\cdot dP = nRdT$$

$$dT = \frac{PdV + V\cdot dP}{nR}$$

Putting this value in eqn II

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

$$\frac{C_vPdV + C_v\cdot VdP + RPdV}{R} = 0$$

$$C_v\underline{PdV} + C_v\cdot VdP + R\underline{PdV} = 0$$

$$(C_v + R)PdV + C_v\cdot V\cdot dP = 0$$

$$C_p\cdot P\cdot dv + C_v\cdot V\cdot dP = 0 \qquad \left(\begin{matrix} C_P - C_v = R \\ C_P = C_v + R \end{matrix}\right)$$

Dividing both sides by $C_v PV$, we get

$$\frac{C_p}{C_V} \times \frac{PdV}{PV} + \frac{C_v\cdot V\cdot dP}{C_V\cdot PV} = 0$$

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

$$\gamma\cdot\frac{dV}{V} + \frac{dP}{P} = 0$$ (where $\frac{C_P}{C_V} = \gamma$ is the ratio of two heat capacities of the gas)

gamma

Integrating both sides,

$$\int \gamma \cdot \frac{dv}{v} + \int \frac{dP}{P} = \int 0$$

$$\gamma \int \frac{dV}{V} + \int \frac{dP}{P} = \text{constant}$$

$$\gamma \log V + \log P = \text{constant}$$

$$\log V^{\gamma} + \log P = \text{constant}$$

$$\log (V^{\gamma} \cdot P) = \text{constant}$$

$$PV^{\gamma} = e^{\text{constant}} \quad \text{or} \quad PV^{\gamma} = \text{constant}$$

$$P_1 V_1^{\gamma} = P_2 V_2^{\gamma}$$

Adiabatic relation in terms of Temp & Vol

we have,

$$PV^{\gamma} = \text{constant} \quad — ①$$

For n mole,

$$PV = nRT$$

$$P = \frac{nRT}{V}$$

Putting this value in eqn - 1

$$\frac{nRT}{V} \cdot V^{\gamma} = \text{constant}$$

$$T \cdot V^{\gamma-1} = \text{constant}$$

$$T_1 V_1^{\gamma-1} = T_2 V_2^{\gamma-1}$$

Adiabatic rlth in terms of pressure & tempr

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

$$P\left(\frac{nRT}{P}\right)^{\gamma} = \text{constant}$$

$$P^{1-\gamma}\, T^{\gamma}\, (nR)^{\gamma} = \text{constant}$$

$$P^{1-\gamma}\, T^{\gamma} = \text{constant}$$

$$P_1^{1-\gamma}\, T_1^{\gamma} = P_2^{1-\gamma}\, T_2^{\gamma}$$

Work Done by the gas ~~id~~ during Adiabatic process

-> Let us consider n mole of gas is filled in a cylinder having insulating walls are fitted with a frictionless & movable piston. When a gas expands from vol V_1 to V_2 work done is given by

$$W = \int_{V_1}^{V_2} P\, dV \quad — ①$$

During adiabatic process,

$$PV^{\gamma} = K \quad (\text{constant})$$

$$P = KV^{-\gamma}$$

Putting in ①

$$W = \int_{V_1}^{V_2} k \frac{dV}{V^\gamma}$$

$$= k \int_{V_1}^{V_2} V^{-\gamma} dV$$

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

$$\frac{1}{\gamma-1} \cdot \left[kV_1^{1-\gamma} - kV_2^{1-\gamma} \right]$$

$$P_1 V_1^\gamma = P_2 V_2^\gamma = k$$

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

$$W = \frac{1}{\gamma-1} (P_1 V_1 - P_2 V_2)$$

let the initial temp is T_1 & final is T_2, then
$P_1 V_1 = nRT_1$ & $P_2 V_2 = nRT_2$

$$W = \frac{1}{\gamma-1} (nRT_1 - nRT_2)$$

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

① PV^{γ} = constant

$TV^{\gamma-1}$ = constant

$P^{1-\gamma} \cdot T^{\gamma}$ = constant

Numerical-2

A gasoline engine takes in air at 25°C & 1 atm & compresses adiabatically to one-length of its original volume. Find final tempr & pressure (γ=1.

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

$$P_2 = P_1 \left(\frac{V_1}{V_2}\right)^{\gamma}$$

$$= 1 \cdot \left(\frac{x}{x/10}\right)^{1.4}$$

$$= 10^{1.4}$$

$$P_2 = 25.11 \text{ atm}$$

$$P_2 = 25.11 \times 1.01 \times 10^5 \text{ Nm}^{-2}$$

$$2537006.29 \text{ Nm}^{-2}$$

ⓑ $$T_1V_1^{\gamma-1} = T_2V_2^{\gamma-1}$$

$$T_2 = T_1 \cdot \left(\frac{V_1}{V_2}\right)^{\gamma-1}$$

$$298 \left(\frac{x}{x/10}\right)^{1.4-1}$$

$$298 \times 10^{0.4}$$

748.5k

* Isochoric process (V -> constant)

If a substance undergoes a process in which volume remains constant, such process is called isochoric process. In this process the gas is kept in a non-expanding chamber. Since vol is constant, work done is zero.

1st law of thermodynamics,

$$dQ = dU + dw$$
$$dQ = dU + 0$$
$$dQ = dU$$
$$dQ = nC_v\,dT$$

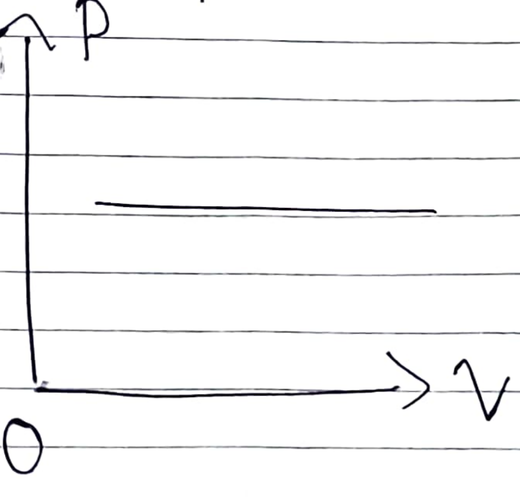
At constant vol,

$$P \propto T$$

$$\left[\frac{P_1}{P_2} = \frac{T_1}{T_2}\right]$$

* Isobaric process (P -> constant)

of a substance is kept in an expanding chamber in which pressure remains constant, such process is called isobaric process.



From 1st law of thermodynamics.

$$dQ = dU + dw$$
$$dQ = nC_v dT + P \cdot dV$$

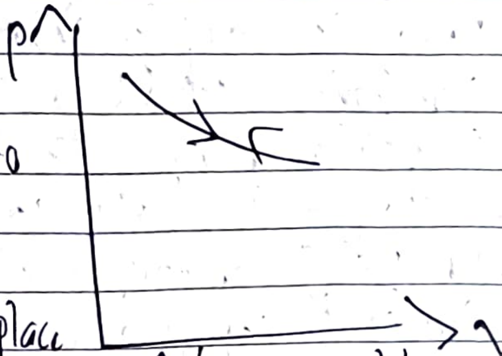
At constant pressure

$$V \propto T$$

$$\left[\frac{V_1}{V_2} = \frac{T_1}{T_2}\right]$$

Reversible Process

-> A reversible process is one which can be retraced in the reverse order;



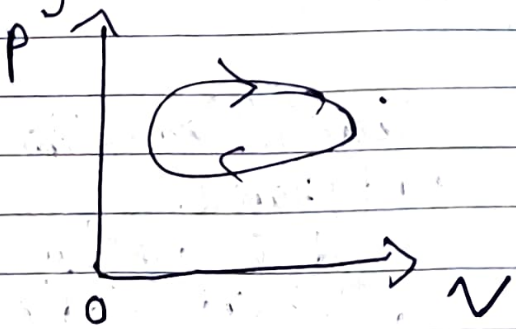
Condition for a process to be reversible:

① All the process taking place in the cycle of operation must be infinitely slow.
② There shouldn't be loss of energy due to conduction, or radiation during the cycle of operation.
③ Work done is 0 as +ve & -ve are opp.

* Irreversible process.

The process which cannot be retraced back in reverse order by reversing the controlling factor is called irreversible process. Almost all natural process are irreversible.

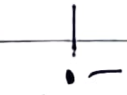
For eg. ① Rusting of iron
② Burning of coal, etc



* Degree for freedom

$\gamma = \frac{C_p}{C_v}$

ⓐ Monoatomic gas (He, Ar, Ne)



$N \rightarrow 1 \leftarrow$ No of atom

$2 \rightarrow 0 \leftarrow$ restriction or bonding

$d.f = 3N - Z$

$3 \times 1 - 0$

$= 3$

$C_v = 3 \times \frac{1}{2} R = \frac{3}{2} R$

Degree of freedom

$C_p = C_v + R = \frac{3}{2} R + R = \frac{5}{2} R$

$\gamma = \frac{C_p}{C_v} = \frac{5/2\ R}{3/2\ R} = 1.67$

$\gamma_{mono} = 1.67$

ⓑ Diatomic gas (O_2, H_2, N_2, air)

O—O $\quad N \rightarrow 2$

$2 - 1$

$D.f = 3N - 2$

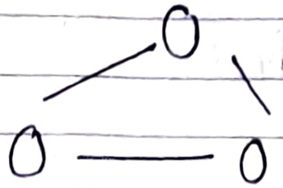
$3 \times 2 - 1$

5

$C_v = \frac{5R}{2}$, $C_p = \frac{5}{2} R + R = \frac{7R}{2}$, $\gamma = \frac{C_p}{C_v} = \frac{7}{5} = 1.4$

$\gamma_{dia} = 1.4$

(c) Triatomic gas (O_3)



$N \rightarrow 3$

$2 \rightarrow 3$

dif = $3N - 2$

$3 \times 3 - 3$

6

$C_V = \frac{6}{2} R$

$C_P = \frac{8R}{2}$

$\gamma = \frac{C_P}{C_V} = \frac{8}{6} = 1.33$

- Adiabatic curve are steeper than isothermal curve. Explain

3 marks

→ For isothermal curve

$$PV = \text{constant}$$

diff both sides

$$d(PV) = d(\text{constant})$$

$$PdV + VdP = 0$$

$$VdP = -PdV$$

$$\left(\frac{dP}{dV}\right)_{iso} = -\frac{P}{V} \quad \text{—— (i)}$$

~~diff~~

For adiabatic curve

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

diff both sides

$$V^{\gamma} dP + P\gamma V^{\gamma-1} dV = 0$$

or, $$V^{\gamma} dP = -P\cdot\gamma V^{\gamma-1} dV$$

$$\frac{dP}{dV} = -\frac{P\cdot\gamma\cdot V^{\gamma-1}}{V^{\gamma}}$$

$$\frac{dP}{dV} = -\frac{P\cdot\gamma}{V}$$

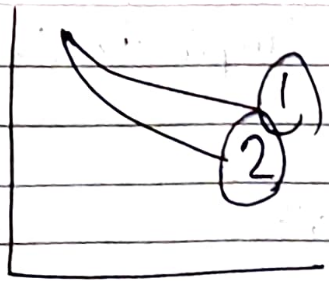
or, $$\left(\frac{dP}{dV}\right)_{adi} = \gamma\left(-\frac{P}{V}\right)$$

or, $$\left(\frac{dP}{dV}\right)_{adi} = \gamma \times \left(\frac{dP}{dV}\right)_{iso} \quad \left[\text{From (i)}\right]$$

Since $\gamma > 1$

$$\left[\therefore \left(\frac{dP}{dV}\right)_{adi} > \left(\frac{dP}{dV}\right)_{iso}\right]$$

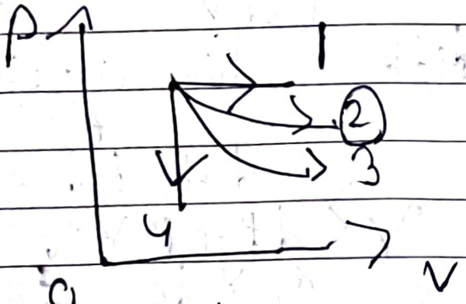
ⓐ The slope of adibatic curve is greater than the slope of isothermal curve by γ times



1 - isothermal
2 - Adibatic

- Identify the process 1, 2, 3 & 4.

1 → isobaric
2 → isothermal
3 ⇒ adiabatic
4 ⇒ isochoric



MCQ Along which process work done is maximum?

ⓐ Isobaric

MCQ: When heat is added to the diatomic gas at isobaric cond, it changes internal energy as well as performs work. The ratio $\Delta Q : \Delta U : \Delta W$ is equal to

ⓐ 7 : 5 : 2 ✓ ⓑ 4 : 3 : 1
c 5 : 3 : 2 ⓓ 3 : 1 : 2

$$dQ = nC_p\,dT$$
$$dU = nC_v\,dT$$
$$\frac{dQ}{dU} = \frac{C_P}{C_V} = \gamma = \frac{7}{5} = \frac{7x}{5x}$$
$$dQ = dU + dW$$
$$7x = 5x + dW$$
$$dW = 2x$$
$$\Delta Q : \Delta U : \Delta W = 7 : 5 : 2$$

MCQ: One mole of monoatomic gas is mixed with two moles of diatomic gas. The molar sp. heat at constant vol of resulting mixture is

(a) 1.83R (b) 2R
(c) 2.17R (d) 3R

monoatomic
$n_1 = 1$
$C_{V_1} = \frac{3}{2}R$

diatomic
$n_2 = 2$
$C_{V_2} = \frac{5R}{2}$

$$C_{V\,mix} = \frac{n_1 C_{V_1} + n_2 C_{V_2}}{n_1 + n_2}$$

$$\frac{1 \times \frac{3}{2}R + 2 \times \frac{5R}{2}}{1+2} = 2.17R$$

MCQ. If 1500 cal of heat is supplied to a gas & 1000J of work is done. What is the change in internal energy?

(a) 3000J (b) 5300J
(c) 4200J (d) 6000J

$dQ = 1500\,cal$
$1500 \times 4.2J$
$6300J$
$dW = 1000J$
$dU = ?$

$dQ = dU + dW$
$6300 = dU + 1000$
$dU = 5300J$

VVI* limitation of 1st law of thermodynamics

(i) The 1st law of thermodynamics doesn't indicate direction of heat flow.

(ii) Heat always flow from body at higher tempr to body at lower tempr but the 1st law doesn't indicate why the heat can't flow from a cold body to a hot one.

(iii) When a bullet strikes a target K.E of a bullet is converted into a heat energy but the 1st law doesn't indicate why heat energy developed in the target cannot converted back into mechanical energy of the bullet.

(iv) The 1st law of thermodynamics doesn't give to what extent the mechanical energy is obtained from the heat energy

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③ The density of gas is 1.775 kg/m³ at 27°C & 10^5 N m^{-2} pressure & sp heat capacity at constant pressure is 846 J kg^{-1} K^{-1}. Calculate the ratio of sp. heat capacity at constant pressure to that at constant volume.

Density (ρ) = 1.775 kg m^{-3}
Tempr (T) = 27°C = 300 K
Pressure (P) = 10^5 N/m²
sp heat capacity at constant pressure (C_p) = 846 J kg^{-1} K^{-1}

$$\frac{C_p}{C_v} = \gamma = ?$$

we have,

$$PV = mrT$$
$$P = \frac{m}{V} rT$$
$$P = \rho r T$$
$$r = \frac{P}{\rho T} = \frac{10^5}{1.77 \times 300} = 187.8 \, J kg^{-1} K^{-1}$$

$$C_p - C_v = r$$
$$C_p = C_p - r$$
$$846 - 187.8$$
$$658.2 \, J kg^{-1} K^{-1}$$
$$\gamma = \frac{C_p}{C_v} = \frac{846}{658.2} = 1.29$$

(19) Air is compressed adiabatically to half of its volume. Calculate the change in its tempr.

-> $V_1 = x$

$V_2 = \frac{x}{2}$

$\gamma = 1.40$

$$T_1 V_1^{\gamma - 1} = T_2 V_2^{\gamma - 1}$$
$$\frac{T_2}{T_1} = \left(\frac{V_1}{V_2}\right)^{\gamma - 1}$$
$$= \left(\frac{x}{x/2}\right)^{1.4 - 1}$$
$$\frac{T_2}{T_1} = 2^{0.4}$$
$$\frac{T_2}{T_1} = 1.32$$
$$\frac{T_2}{T_1} - 1 = 1.32 - 1$$

$$\frac{T_2 - T_1}{T_1} = 0.32$$

$$\frac{\Delta T}{T_1} = 0.32 \times 100\%$$

32%.

⑤ Numerical ⑤

→ For hydrogen the molar heatcapacities at constant volume & pressure are $20.5\,J\,mol^{-1}K^{-1}$ & $28.8\,J\,mol^{-1}K^{-1}$. Calculate

ⓘ The heat needed to raise the tempr of 8g of hydrogen from 10°C to 15°C.

(ii) The increase in the internal energy of the gas. (molar mass of $H_2 = 2g$)

$C_v = 20.5\,J\,mol^{-1}K^{-1}$

$C_p = 28.8\,J\,mol^{-1}K^{-1}$

mass(m) = 8 g

$T_1 = 10°C$

$T_2 = 15°C$

molecular mass (M) = 2 g

no of mole $(n) = \frac{m}{M} = \frac{8}{2} = 4$

$dT = T_2 - T_1 = 15 - 10 = 5°C$

ⓐ $dQ = nC_p dT$

$4 \times 28.8 \times 5$

$= 576\,J$

ⓑ $dU = nC_v dT$

$4 \times 20.5 \times 5$

$\underline{410}$