Padasalai's Telegram Groups!

(தலைப்பிற்கு கீழே உள்ள லிங்கை கிளிக் செய்து குழுவில் இணையவும்!)

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Unit-8 Heat and Thermodynamics.

Heat:

When an object at higher concentration is placed in contact with another object at lower temperature, there will be a spontaneous flow of energy from the object at higher temperature to the one at lower temperature. This energy is called heat.

Temperature:

Temperature is the degree of hotness or coolness of a body.

Conversion of Temperature:

Scale	Kelvin	Kelvin

Boyle's law & Charle's law and Ideal gas law:

Boyle's Law:
At constant temp.
The pressure of gas is inversely proportional to the volume.

$P \propto \frac{1}{V}$. It was discovered by Robert Boyle and is known as Boyle's law.

Charles law: When the gas is kept at constant pressure, the volume of gas is directly proportional to absolute temperature.

$V \propto T$. By combining two equations

$PV \propto T$

$PV = KT$

here $K = C$ [Positive constant]

$PV = CT$ → (1) ideal gas equation.

C - No. of particles in a gas container.

$P(nV) = (nC)T$

$\frac{Pnv}{T} = nC$

k - Boltzmann constant.

from (1).

$PV = NKT$ →

Mole:
One mole is defined as the amount of substance which contain Avogadro number (N_A) of particles.

$$N = \mu N_A \quad \rightarrow (3)$$

where,
μ - moles of particle
N_A - Avogadro number.

sub. eqn (3) in (2).

$$PV = \mu N_A kT \quad \rightarrow (4)$$

where,

$N_A k = R$ [universal gas constant] $\rightarrow (5)$.

$= 8.314$ J/mol.K.

sub. eqn (5) in (4).

$$\boxed{PV = \mu RT}.$$

Eqn 8.8:

Ideal gas equation,

$$P_1 V_1 = NkT_1 \quad \rightarrow (1)$$
$$P_2 V_2 = NkT_2 \quad \rightarrow (2)$$

We know,

$V_1 = V_2 = V$

$$P_1 V = NkT_1 \quad \rightarrow (3)$$
$$P_2 V = NkT_2 \quad \rightarrow (4).$$

$\div$ the eqn (3) & (4)

$$\frac{P_1 V}{P_2 V} = \frac{NkT_1}{NkT_2}$$

$$\frac{P_1}{P_2} = \frac{T_1}{T_2}$$

$$P_2 = \frac{T_2}{T_1} P_1$$

$$P_2 = \frac{81.8}{303} \times 1.4 \times 10^5 \text{ Pa}.$$

$$P_2 = 269.6 \text{ kPa}$$

Example: 8.3

the ideal gas law.

$PV = NkT$

$N = \frac{PV}{KT}$

Here Volume = 5.5 litre

1 litre $= \frac{1}{1000}$

$= \frac{1}{10^3}$

$= 10^{-3} m^3$

1 atm: 101325 Pa

1 atm = 1.01×10^5

1 atm = 101.325 kPascal

101325 Pa = 1.01325 bar.

$$N = \frac{PV}{KT} = \frac{1.01 \times 10^5 Pa \times 5.5 \times 10^{-3} m^3}{1.38 \times 10^{-23} JK^{-1} \times 310K}$$

$= 1.29 \times 10^{23}$ Molecules.

Oxygen = 21%

$= 1.29 \times 10^{23} \times \frac{21}{100}$

$= 27.09 \times 10^{23} \times 10^{-2}$

$= 27.09 \times 10^{21}$

$= 2.7 \times 10^{22}$ molecules.

Specific heat capacity:

Specific heat capacity of a substance is defined as the amount of heat energy required to raise the temp. of 1 kg of a substance by 1 Kelvin or 1°C.

Heat capacity, $S = \frac{\Delta Q}{\Delta T}$.

Here,

ΔQ - Amount of heat energy

ΔT - Change in temperature

$\Delta Q = mS\Delta t$

Where,

m - mass of substance.

Specific heat capacity, $S = \frac{1}{m}\left[\frac{\Delta Q}{\Delta T}\right]$.

Molar heat capacity:

It is defined as heat energy required to increase the temp. of one mole of substance by 1K or 1°C.

$$c = \frac{1}{\mu}\left(\frac{\Delta Q}{\Delta T}\right).$$

Where,

c - molar specific heat

μ - no. of moles.

The SI unit is $Jmol^{-1}K^{-1}$.

Thermal expansion of solid, liquid and gases:

Thermal expansion is the tendency of matter to change in shape, area and volume due to a change in temperature.

Classification of expansion:

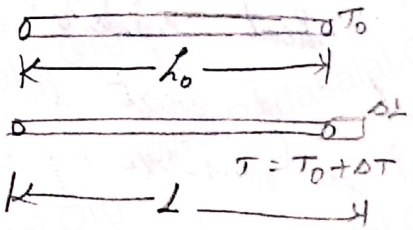
Linear expansion

Area expansion

Volume expansion

Linear expansion:

$$\frac{\Delta L}{L_0} = \alpha \Delta T.$$



Where,

$\frac{\Delta L}{L}$ - Change in length

ΔT - Change in temperature.

$$\frac{\Delta L}{L_0} = \alpha_L \Delta T$$

Therefore,

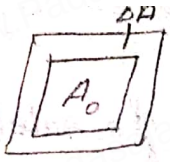
$$\alpha_L = \frac{\Delta L}{L \Delta T}$$

Where,

α_L = Co-efficient in length

L - Original length.

Area Expansion:



$$\frac{\Delta A}{A} = 2\alpha_A \Delta T$$

$$\alpha_A = \frac{\Delta A}{2A\Delta T}$$

$$\frac{\Delta A}{A_0} = 2\alpha\Delta T$$

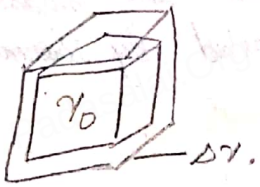
Where,

α_A - co-efficient of area expansion

A - Original Area

ΔT - Change in temperature.

Volume expansion:



$$\frac{\Delta V}{V} = 3\alpha_V \Delta T$$

$$\alpha_V = \frac{\Delta V}{3V\Delta T}$$

Where,

α_V - co-efficient of volume expansion

ΔV - change in volume

V - Original volume

ΔT - change in temperature.

Example 8.6

$\alpha = 10 \times 10^{-6}$ per °C

$L = 300m.$

$\Delta T = 25° - 2°$

$\Delta T = 23°$

To find height, h

$$\frac{\Delta L}{L} = \alpha_L \Delta T$$

$$\Delta L = \alpha_L \Delta T \cdot L$$

$$= 10 \times 10^{-6} \times 300 \times 23.$$

$$= 10 \times 10^{-6} \times 3 \times 10^{2} \times 23$$

$$= 690 \times 10^{-4}$$

$$= 69 \times 10^{-3}$$

$$\Delta L = 0.069m$$

$\frac{2 \times 30}{69}$

Triple point:

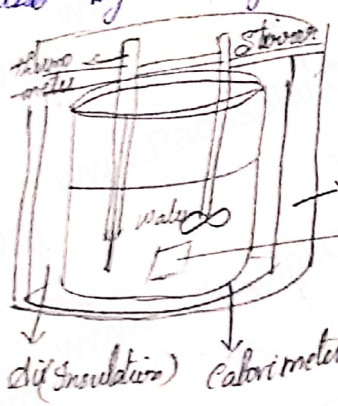
Triple point of a substance is the temp. and ... at which the three phases (gas, liquid and solid) of that substance co-exist in thermodynamic equilibrium.

triple point of water - 273.1 K

Partial vapour pressure - 611.657 Pascal.

Calorimetry:

It means the measurement of the amount of heat released or absorbed by thermodynamic system during the heating process.



At equilibrium:

Heat gain = Heat lost.

$$Q_{gain} = -Q_{lost} \rightarrow (1)$$

Here,

T_1 - High temperature

T_2 - Room temperature.

T_f - final equilibrium temperature.

We know that,

Specific heat capacity, $\Delta Q = ms\Delta t$

Here,

$$Q_{lost} = m_1 s_1 (T_f - T_1) \rightarrow (2).$$

$$Q_{gain} = m_2 s_2 (T_f - T_2) \rightarrow (3).$$

sub. equ (2) & (3) in (1).

$$m_2 s_2 (T_f - T_2) = -m_1 s_1 (T_f - T_1).$$

$$m_2 s_2 T_f - m_2 s_2 T_2 = -m_1 s_1 T_f + m_1 s_1 T_1$$

$$m_2 s_2 T_f + m_1 s_1 T_f = m_2 s_2 T_2 + m_1 s_1 T_1$$

$T_f(m_1 s_2 + m_2 s_1) = m_1 s_1 T_1 + m_2 s_2 T_2$

The final temperature: $T_f = \dfrac{m_1 s_1 T_1 + m_2 s_2 T_2}{m_1 s_1 + m_2 s_2}$

Example 8.7

$$T_f = \frac{m_1 s_1 T_1 + m_2 s_2 T_2}{m_1 s_1 + m_2 s_2}.$$

Given:

$m_1 = 5L = 5kg \quad m_2 = 4L = 4kg.$

$s_1 = s_2.$

$T_1 = 50°C = 323K \quad T_2 = 30°C = 303K$

$$T_f = \frac{m_1 T_1 + m_2 T_2}{m_1 + m_2} = \frac{5 \times 323 + 4 \times 303}{5+4}.$$

$$= \frac{1615 + 1212}{9}$$

$$= \frac{2827}{9} = 314.11K.$$

$$= 314.11K - 273K = 41.11°C.$$

$\approx 41°C$

If we mix equal amount of water, (extra)

$$T_f = \frac{m_1 s_1 T_1 + m_2 s_2 T_2}{m_1 s_1 + m_2 s_2}$$

if $m_1 = m_2 = m$

$$T_f = \frac{m s_1 T_1 + m s_2 T_2}{m s_1 + m s_2} = \frac{m(s_1 T_1 + s_2 T_2)}{m(s_1 + s_2)}$$

$$\boxed{T_f = \frac{s_1 T_1 + s_2 T_2}{s_1 + s_2}}.$$

if $s_1 = s_2 = s.$

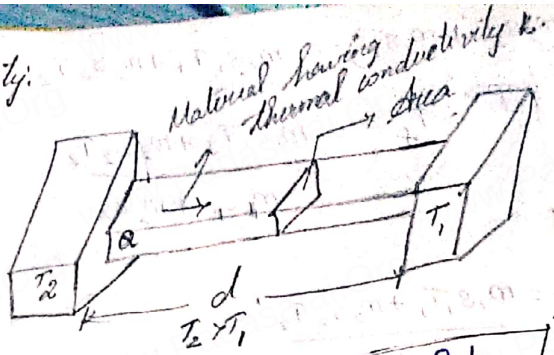
$$T_f = \frac{s T_1 + s T_2}{s+s} = \frac{s(T_1 + T_2)}{2s}$$

$$\boxed{T_f = \frac{T_1 + T_2}{2}}$$

$$T_f = \frac{323 + 303}{2} = \frac{626}{2} = 313K$$

$= 40°C.$

Thermal conductivity:



Double.

$$\frac{Q}{t} = \frac{KA\Delta T}{L} \Rightarrow \boxed{K = \frac{QL}{tKA\Delta T}}$$

Where,

k - co-efficient of thermal conductivity.

SI unit - $m^{-1}K^{-1}$ (or) $Wm^{-1}K^{-1}$.

Steady state:

The state at which temp. attains constant value everywhere and there is no further transfer of heat anywhere is called steady state.

Newton's law of cooling:

Let us consider,

m - Object of mass

S - Specific heat capacity

T - Temperature of object

T_s - Temperature of surrounding

dT - small amount of temp.

dt - time.

We know

$$S = \frac{1}{m} \cdot \frac{\Delta Q}{\Delta t}$$

$$S = \frac{1}{m} \cdot \frac{dQ}{dt}$$

$$dQ = m \cdot S \cdot dt$$

Here,

$t = T$

$$dQ = m.S.dT. \rightarrow (1)$$

÷ by dt on both sides.

$$\frac{dQ}{dt} = \frac{mS\,dT}{dt}$$

From Newton's law of cooling,

$$\frac{dQ}{dt} \propto -(T - T_S)$$

$$\frac{dQ}{dt} = -a(T - T_S) \quad \rightarrow (2)$$

where, a - positive constant.

from equ (1) & (2)

$$-a(T - T_S) = ms\frac{dT}{dt}$$

$$\frac{-a}{ms}dt = \frac{dT}{T - T_S}$$

$$\frac{dT}{T - T_S} = -\frac{a}{ms}dt$$

Taking $\int$ on both sides.

$$\int \frac{dT}{T - T_S} = \int \frac{-a}{ms}dt$$

$$\int \frac{1}{T - T_S}dT = \frac{-a}{ms}\int 1 \cdot dt$$

$$\log(T - T_S) = \frac{-a}{ms}t + b_1$$

where

b_1 - integration constant.

Take exponential on both sides

$$e^{\log(T - T_S)} = e^{\frac{-a}{ms}(t) + b_1}$$

$$T - T_S = e^{b_1} \times e^{\frac{-at}{ms}}$$

$$T - T_S = b_2 e^{\frac{-a}{ms}t}$$

here, $b_2 = e^{b_1}$

$$T = T_S + b_2 e^{\frac{-a}{ms}t}$$

Example: 8.8

The hot water cools = 92°C - 84°C

= 8°C

Time = 3 min.

The average temp $= 98°C - 84°C$

$= 88°C$

At room temp, average temp $= 61°C$

We know that,

$$\frac{dT}{T - T_S} = -\frac{a}{ms} dt$$

$$\frac{dT}{dt} = -\frac{a}{ms}(T - T_S).$$

$$\frac{8°C}{3\,min} = -\frac{a}{ms}(61°C) \rightarrow (1)$$

Wly. average temp of $= 65°C - 60°C$

$= 62.5°C$

Average temp. at room temp. $= 35.5°C$

$$\frac{5°C}{dt} = -\frac{a}{ms}(35.5°C) \rightarrow (2).$$

Divide the eqn. (1) & (2).

$$\frac{\frac{8°C}{3\,min}}{\frac{5°C}{dt}} = \frac{-\frac{a}{ms}(61°C)}{-\frac{a}{ms}(35.5°C)}$$

$$\frac{8 \times dt}{3 \times 5} = \frac{-a}{ms} \times \frac{ms}{-a} \times \frac{61}{35.5}$$

$$\frac{8}{15} dt = \frac{61}{35.5}$$

$$dt = \frac{61 \times 15}{8 \times 35.5} = \frac{915}{284}$$

$$= \frac{915}{284}$$

$$dt = 3.22 \text{ min}$$

Prevost theory:

Only at absolute zero temperature a body will stop emitting. Therefore, Prevost theory states that all bodies emits thermal radiation at all temperature above absolute zero irrespective of the nature of surroundings.

Black Body.

A perfect black body is one which absorbs heat radiation of all wavelength that fall upon it.

Wein's law:

The wavelength of maximum intensity of emission of a black body radiation is inversely proportional to the absolute temperature of the black body.

$$\lambda_m \propto \frac{1}{T} \quad (or) \quad \lambda_m = \frac{b}{T}.$$

Where,

b - Wein's constant The value is 2.898×10^{-3} mk.

Example 8.9:

From Weins displacement law,

$$\lambda_m = \frac{b}{T}.$$

Here for object A,

$$\lambda_A = \frac{b}{T}.$$

$$\lambda_A T_A = b \rightarrow (1) \text{ (constant)}.$$

for object B.

$$\lambda_B = \frac{b}{T}$$

$$\lambda_B \cdot T_B = b \text{ (constant)}$$

from (1) & (2),

$$\lambda_A T_A = \lambda_B T_B.$$

Here, $\lambda_B = \frac{1}{2}\lambda_A$.

$$\frac{T_B}{T_A} = \frac{\lambda_A}{\lambda_B}.$$

$$= \frac{1}{1/2}$$

$$= 2$$

$$T_B = 2T_A$$

from stefan - Boltzmann law,

$$\frac{E_B}{E_A} = \left(\frac{T_B}{T_A}\right)^4$$

$$= (2)^4$$

$$= 16.$$

$$N = 16.$$

Thermodynamics:

It is a branch of physics which describes the laws governing the process of conversion of work into heat and conversion of heat into work.

Equation of state:

The equation which connects the state variables in a specific manner is called equation of state.

Extensive variables:

It is that depends on the mass or size of the system.

Intensive variables:

It is independent on the mass or size of the system:

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

Zeroth law of thermodynamics:

If two system A and B are in thermal equilibrium with a third system C, then A and B are in thermal equilibrium with each other.

Internal Energy:

The internal energy of a thermodynamic system is the sum of kinetic and potential energies of all the molecules of the system with respect to the center of mass of the system.

$$U = E_K + E_P.$$

The energy due to molecular motion including translational, rotational and vibrational motion is called internal kinetic energy. (U_K).

Example 8.11.

Mass of water = 25 l
= 25 kg
Assume (1 l = 1 kg)

$W = mgh = 25 \times 10 \times 25$
$= 625 \times 10$
$W = 6250$ J.

Total energy gained from the food = 800 food cal.
= 800 kcal
1 cal = 4.186
$= 800 \times 10^3 \times 4.186$ J.
$= 8 \times 10^5 \times 4.186$ J
$= 8.37 \times 10^5$ J.

$n = \frac{8.37 \times 10^5 \text{ J}}{6250 \text{ J}} = 134.$

n = 134 times.

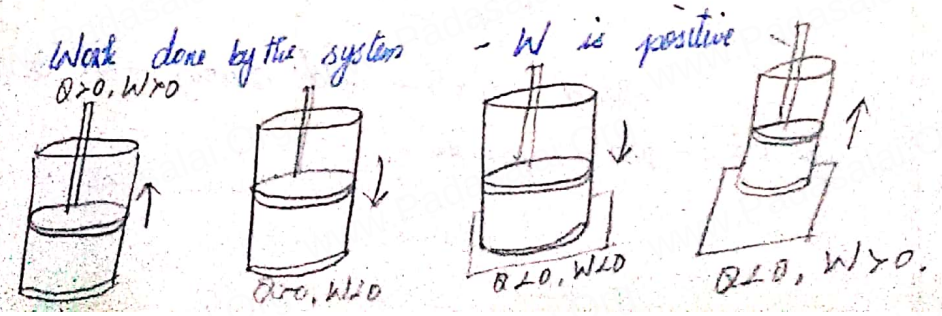
First law of thermodynamics:

Change in internal energy of the system is equal to heat supplied to the system (Q) minus the work done by the system (W) on the surroundings.

$\Delta U = Q - W.$

Sign convention:

System gain heat - Q is positive
System loses heat - Q is negative
Work done on the system - W is negative
Work done by the system - W is positive



Example 8.12.

Work done on the system.

$W = -20kJ$

$= -30,000J$

Heat flowing out of the system.

$Q = -5kcal$

$= -5 \times 10^3 \times 4.184J$

$= -5 \times 10^3 \times 4184 \times 10^{-3}$

$= -20920J.$

First law of Thermodynamics.

$\Delta U = Q - W$

$= -20920 - (-30,000)$

$= -20920 + 30,000$

$\Delta U = 9080J$

30,000
20920
9080

Example 8.13

Work done by the system.

$W = 500kJ$

Heat released from the system.

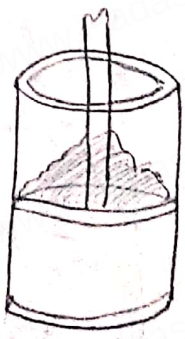
$Q = -230kJ.$

$\Delta U = Q - W$

$= -230 - 500$

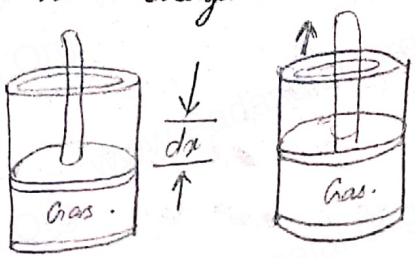
$= -730J.$

Example 8.14.



Sand Particles add slowly - It is a quasi - static process

Work done in a volume changes:



$dW = F dx \rightarrow (1)$.

The force exerted by the gas on the piston,

$F = P \cdot A \rightarrow (2)$.

Where,

P - Pressure exerted by the gas.

A - Area of the piston.

sub (2) in (1).

$dw = P \cdot A \cdot dx \rightarrow (3)$.

Here, $A dx = dv \rightarrow (4)$ [change in volume].

sub eqn (4) in (3).

$dw = P dv \rightarrow (5)$.

Take $\int$ on both sides.

$\int dW = \int P dv$

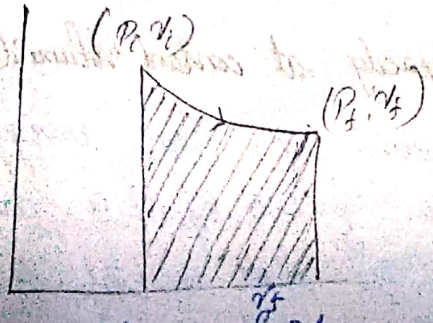
$$W = \int_{v_i}^{v_f} P dv.$$

Suppose, if work is done on the system,

then $v_i > v_f$

Then, W is negative.

Pressure - volume diagram (PV).



$W = \text{Area} = \int_{v_i}^{v_f} P dv.$

Eg: 8-15

a). The Pressure, $P = 1$ atm.
$= 101$ kPa.
$= 1.01 \times 10^5$ Pa → (2)

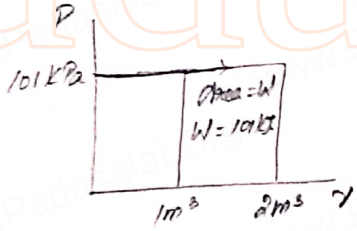
$V_i = 1m^3$ → (3) $V_f = 2m^3$ → (4)

$$W = \int_{V_i}^{V_f} P\,dV$$
$$= P \int_{V_i}^{V_f} dV$$
$$= P [V]_{V_i}^{V_f}$$
$$= P [V_f - V_i]. \rightarrow (5).$$

Sub. eqn (1), (3), (4) in (5).

$$W = 101 \times 10^3 [2 - 1]$$
$$= 101 \times 10^3 [1].$$
$$= 101 \times 10^3$$
$$W = 101 \text{ kJ}$$

b).



Specific heat capacity of constant Pressure: (S_p).

The amount of heat energy required to raise the temperature of one kg of a substance by 1K or 1°C by keeping the pressure constant is called specific heat capacity of a gas at constant Pressure.

Specific heat capacity at constant Volume: (S_v).

The amount of heat energy required to raise the temp. of one kg of a substance by 1K or 1°C by keeping the volume constant is called specific heat capacity at constant volume.

Molar heat capacity at constant pressure (C_p)

The amount of heat required to raise the temperature of one mole of a substance by 1K or 1°C at constant volume is called molar specific heat capacity at constant volume (C_v). If pressure is kept constant, it is called molar specific heat capacity at constant pressure (C_p).

Molar specific heat capacities:

molar specific heat capacity,

$$C = \frac{1}{\mu}\left(\frac{\Delta Q}{\Delta T}\right) \rightarrow (1).$$

If. Q - heat supplied to mole of a gas at constant volume.

ΔT - change in temperature.

from equ (1).

$$Q = \mu C_v \cdot \Delta T \rightarrow (2)$$

from 1st law of thermo dynamics.

$$\Delta U = Q - W \quad (W \to 0)$$

Here, $W = 0$ $\therefore dW = 0$.

$$\Delta U = Q.$$

$$Q = \Delta U \rightarrow (3).$$

Compare equ. (2) & (3).

$$\Delta U = \mu C_v \Delta T$$

$$C_v = \frac{1}{\mu}\frac{\Delta U}{\Delta T}$$

If the limit $\Delta T \to 0$.

$$\boxed{C_v = \frac{1}{\mu}\frac{dU}{dT}}$$

Mayer's Relation:

Consider μ mole of an ideal gas in a container with volume V, pressure P and temperature T.

Here, dT = increase in temperature

dU = change in internal energy.

We know that,

$$C_v = \frac{1}{\mu}\frac{dU}{dT}$$

$$dU = \mu C_v dT \rightarrow (1).$$

We know that,

The molar specific heat capacity at constant pressure

$$C_p = \frac{1}{\mu}\frac{\delta Q}{\delta T}$$

Here,

$\delta Q = Q$ (No change in heat)

$\delta T = dT$ = change in temperature.

Then,

$$Q = \mu C_p dT \rightarrow (2).$$

If W is the work done by the gas,

$$W = P dV \rightarrow (3).$$

from 1st law of thermodynamics.

$$dU = Q - W$$

$$Q = dU + W \rightarrow (4).$$

Sub eqn (1) & (4) & (3) in (5).

$$\mu C_p dT = \mu C_v dT + P dV. \rightarrow (5).$$

from ideal gas equation.

$$PV = \mu RT$$

Differentiate both sides

$$P dV + V dP = \mu R dT$$

Here, pressure is constant.

so, $dP = 0$.

$$P dV + V(0) = \mu R dT$$

$$P dV = \mu R dT \rightarrow (6).$$

Sub eqn (6) in (5).

$$\mu C_p dT = \mu C_v dT + \mu R dT$$

$$\mu C_p dT = \mu dT [C_v + R]$$

$C_p = C_v + R$.

(or)

$\boxed{C_p - C_v = R}$ This is Meyer's relation

where,

R - Universal gas constant

C_p - Molar heat capacity at constant pressure

C_v - Molar heat capacity at constant volume.

Thermodynamic process:

There are four types.

- Isothermal process
- Isobaric process
- Isochoric process
- Adiabatic process.

Isothermal Process:

It is a process in which the temperature remains constant but the pressure and volume of a thermodynamic system will change.

$PV = \mu RT$

Here,

T - constant.

PV = constant.

$P \propto \frac{1}{V}$.

For isothermal process:

1st law of thermodynamics

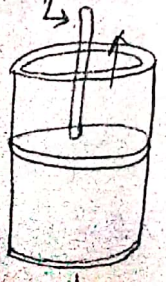
$\Delta U = Q - W$

here, $\Delta U = 0$

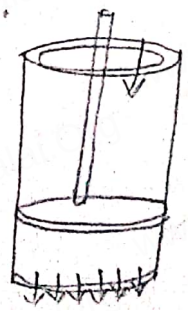
$0 = Q - W$

$\boxed{Q = W}$.

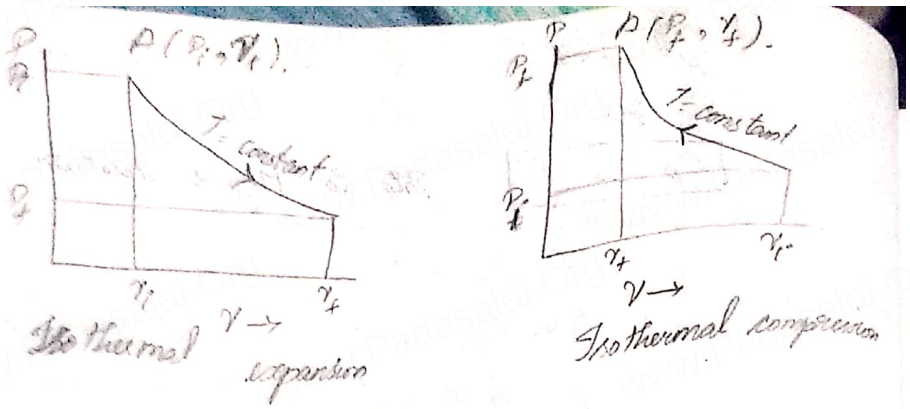
Isothermal expansion and compression:



$Q > 0, W > 0$.



$Q < 0, W < 0$.



Work done in an Isothermal process:

Consider an ideal gas which is allowed to expand ~~egas~~ quasi-statically at constant temperature. from initial state (P_i, V_i) to the final state (P_f, V_f). We can calculate the work done by the gas during this process.

The Work done by the gas, $W = \int_{V_i}^{V_f} P dV \rightarrow (1)$.

from ideal gas equ.

$PV = \mu RT$

$P = \frac{\mu RT}{V} \rightarrow (2)$

sub. equ (2) in (1).

$W = \int_{V_i}^{V_f} \frac{\mu RT}{V} dV$

$= \mu RT \int_{V_i}^{V_f} \frac{1}{V} dV$

$W = \mu RT [\log V]_{V_i}^{V_f}$

$W = \mu RT [\log V_f - \log V_i]$.

$W = \mu RT \log\left(\frac{V_f}{V_i}\right) \rightarrow (3)$

for expansion:

$V_f > V_i$

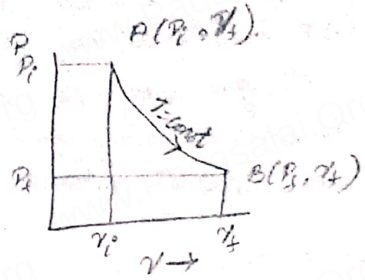
$\frac{V_f}{V_i} > 1$

then

$\log\left(\frac{V_f}{V_i}\right) > 0$.

from eqn ②.

Since, the isothermal process the work done is positive in expansion.



for compression.

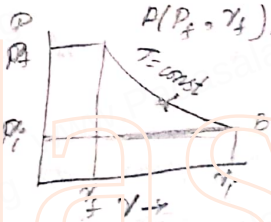
$$V_f < V_i$$

$$\frac{V_f}{V_i} < 1$$

Then

$$\log\left(\frac{V_f}{V_i}\right) < 0$$

Since, from ② the isothermal, the work done is negative in compression.



Example 8.16

a) We know that,

$$W = \mu RT \log\left(\frac{V_f}{V_i}\right).$$

$$= 0.5 \text{ mole} \times \frac{8.31 \text{ J}}{\text{mol.K}} \times 300 \text{ K} \log\left(\frac{4L}{2L}\right).$$

$$W = 5996 \text{ J}$$

$$W = 5.996 \text{ kJ}.$$

b). From 1st law of thermodynamics,

$$\Delta U = Q - W$$

$$Q = W$$

$$Q = 5996 \text{ J}$$

$$Q = 5.996 \text{ kJ}$$

c) $$P_i V_i = P_f V_f = \mu RT$$

$$P_f = \frac{\mu RT}{V_f}$$

$$= 0.5 \times \frac{8.31}{\text{mol.K}} \times \frac{100}{\frac{4 \times 10^{-3}}{2}} = 2077.5 \times 10^{-1}$$

$$= 207.75 \text{ kPa}.$$

Adiabatic process:

This is a process in which no heat flows into or out of the system ($Q = 0$). But the gas can expand by spending its internal energy as gas can be compressed through some external work. So, the pressure, volume and temperature of the system may change in an adiabatic process.

$$\Delta U = -W$$

The equation of state for an adiabatic process, $PV^{\gamma} = \text{constant}$

where,

γ - adiabatic exponent

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

where,

C_p - molar specific heat capacity at constant pressure.

C_v - molar specific heat capacity at constant volume.

If the gas goes from one equilibrium state (P_i, V_i) to another equilibrium state (P_f, V_f).

$P_i V_i^{\gamma} = P_f V_f^{\gamma} = \text{constant}$. $P_i V_i^{\gamma} = \text{constant} \rightarrow (2)$

$P_f V_f^{\gamma} = \text{constant} \rightarrow (3)$.

Compare eqn (2) & (3).

$$P_i V_i^{\gamma} = P_f V_f^{\gamma} \rightarrow (4).$$

from ideal gas eqn.

$PV = \mu RT \rightarrow (5)$.

$P = \frac{\mu RT}{V} \rightarrow (6)$.

Sub (6) in (1)

$$\frac{\mu RT}{V} V^{\gamma} = \text{constant}.$$

$$\frac{T}{V} V^{\gamma} = \frac{\text{constant}}{\mu R}.$$

Thus, $\frac{\text{constant}}{\mu R}$ = Another constant C.

$\frac{T}{V} V^{\gamma} = \text{constant}$

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

$T V^{\gamma-1} = \text{constant} \quad \rightarrow (3)$.

At initial equilibrium state.

from equ (3).

$T_i V_i^{\gamma-1} = \text{constant} \quad \rightarrow (4)$

At final equilibrium state.

from equ (3).

$T_f V_f^{\gamma-1} = \text{constant} \quad \rightarrow (5)$.

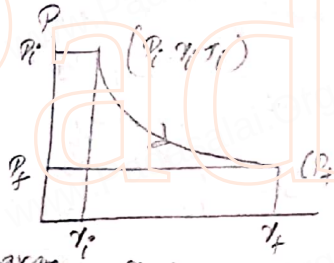
Compare equ (4) & (5).

$T_i V_i^{\gamma-1} = T_f V_f^{\gamma-1}$

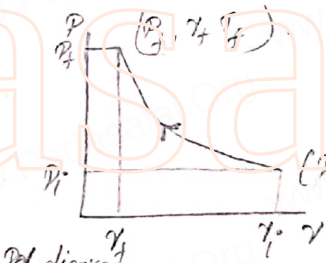
This is the equation of state for adiabatic process in term of T and V.

Similarly, the equation of state for adiabatic process is terms of T and P.

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



PV diagram
Adiabatic expansion



PV diagram.
Adiabatic compression.

Example 8.18:

We know that $T_i V_i^{\gamma-1} = T_f V_f^{\gamma-1}$

$T_f = T_i \left(\frac{V_i}{V_f}\right)^{\gamma-1}$

Here, $T_i = 27°C$
$= 27 + 273$
$T_i = 300K$.

Here, $V_i = V$

$V_f = \frac{V}{4}$.

Assume, $\gamma = \frac{4}{3}$.

$\gamma = 1.4$

$T_f = 300 \left(\frac{V}{V/4}\right)^{1.4-1}$

$T_f = 300 \times (4)^{0.4}$

Here, $x = 4^{0.4}$

Take log on both sides

$\log x = \log(4)^{0.4}$

$\log x = 0.4 \times \log 4$

$\log x = 0.4 \times 0.6020$

$\log x = 0.2408$

Take antilog on both sides

Antilog $\log(x)$ = Antilog (0.2408)

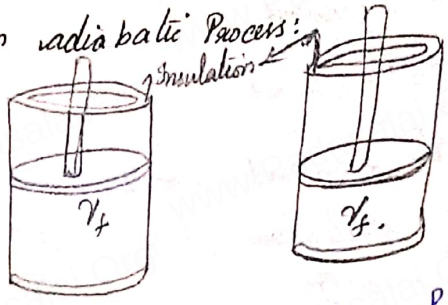
$x = 1.738 + 3$

$x = 1.741$

$T_f = 300 \times 1.741 = 522K$.

$T_f = 522 - 273$ $T_f = 249°C$.

Work done in adiabatic Process:



Let, W be the work done when the system goes from the initial state $(P_i\ V_i\ T_i)$ to the final state $(P_f\ V_f\ T_f)$ adiabatically.

Work done in adiabatic process, $W = \int_{V_i}^{V_f} P \cdot dV \rightarrow (1)$.

We know that,

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

$$P = \frac{\text{constant}}{V^{\gamma}} \rightarrow (2).$$

sub. equ (2) in (1).

$$W = \int_{V_i}^{V_f} \frac{\text{constant}}{V^{\gamma}} dV.$$

$$W = \text{constant} \int_{V_i}^{V_f} \frac{1}{V^{\gamma}} dV$$

$$W = \text{constant} \int_{V_i}^{V_f} V^{-\gamma} dV$$

$$W = \text{constant} \left[\frac{V^{-\gamma+1}}{-\gamma+1}\right]_{V_i}^{V_f}$$

$$W = \text{constant} \left[\frac{V^{-\gamma+1}}{1-\gamma}\right]_{V_i}^{V_f}$$

$$= \frac{\text{constant}}{1-\gamma} \left[V^{-\gamma+1}\right]_{V_i}^{V_f}$$

$$= \frac{\text{constant}}{1-\gamma} \left[V_f^{-\gamma+1} - V_i^{-\gamma+1}\right].$$

$$W = \frac{\text{constant}}{1-\gamma} \left[\frac{1}{V_f^{\gamma-1}} - \frac{1}{V_i^{\gamma-1}}\right]$$

$$W = \frac{1}{1-\gamma} \left[\frac{\text{constant}}{V_f^{\gamma-1}} - \frac{\text{constant}}{V_i^{\gamma-1}}\right] \rightarrow$$

We know that.

$$P_i V_i^{\gamma} = \text{constant} \rightarrow (4).$$

$$P_f V_f^{\gamma} = \text{constant} \rightarrow (5).$$

Sub. equ (4) & (5) in (3).

$$W = \frac{1}{1-\gamma} \left[\frac{P_f V_f^{\gamma}}{V^{\gamma-1}} - \frac{P_i V_i^{\gamma}}{\quad}\right]$$

$$W = \frac{1}{1-\gamma}\left[P_f V_f^{\gamma} V_f^{-\gamma+1} - P_i V_i^{\gamma} V^{-\gamma+1}\right]$$

$$= \frac{1}{1-\gamma}\left[P_f V_f^{\gamma-\gamma+1} - P_i V_i^{\gamma-\gamma+1}\right]$$

$$W = \frac{1}{1-\gamma}\left[P_f V_f - P_i V_i\right].$$

$$W_{adia} = \frac{1}{1-\gamma}\left[P_f V_f - P_i V_i\right] \rightarrow (6)$$

From ideal gas law,

$$P_f V_f = \mu R T_f \rightarrow (7)$$

$$P_i V_i = \mu R T_i \rightarrow (8).$$

Sub. eqn (7) & (8) in (6).

$$W_{adia} = \frac{1}{1-\gamma}\left[\mu R T_f - \mu R T_i\right].$$

$$= \frac{1}{1-\gamma}\left[-\mu R T_i + \mu R T_f\right]$$

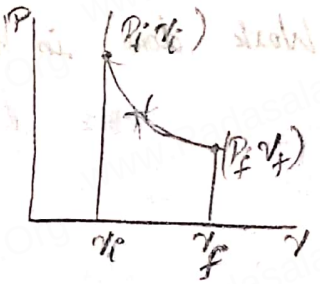
$$= -\frac{\mu R}{1-\gamma}\left[T_i - T_f\right]$$

$$\boxed{W_{adia} = \frac{\mu R}{\gamma-1}\left[T_i - T_f\right]} \rightarrow (9).$$

for expansion:

$T_i > T_f$. the gas cools during adiabatic expansion.

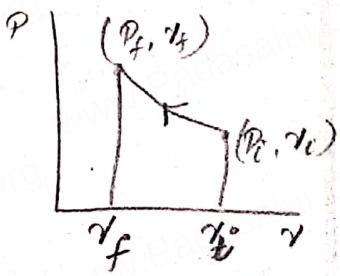
W_{adia} is positive.



for compression:

$T_i < T_f$ the temperature of gas increases during adiabatic compression.

W_{adia} is negative.



Isobaric process:

This is a thermodynamic process that occurs at constant pressure. Even though pressure is constant in this process, temperature, volume and internal energy are not constant.

from ideal gas equation

$PV = \mu RT$

$V = \frac{\mu RT}{P}$

Here, $\frac{\mu R}{P}$ = constant.

V = constant $\times T$ → ①.

$V \propto T$

$T \propto V$

from ①.

$\frac{T}{V}$ = constant

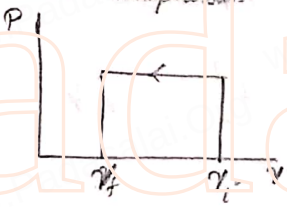
If a gas goes from a state (V_i, T_i) to (V_f, T_f) at constant pressure, then the system satisfies the following equation.

$\frac{T_i}{V_i}$ = constant → ②.

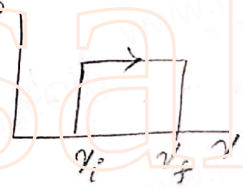
$\frac{T_f}{V_f}$ = constant → ③.

$\frac{T_f}{V_f} = \frac{T_i}{V_i}$ → ④.

Isobaric compression



Isobaric expansion



Work done in the Isobaric process:

Work done in an isobaric process, $W = \int_{V_i}^{V_f} P\,dV$

Here,

Pressure 'P' is constant.

$W = P \int_{V_i}^{V_f} 1 \cdot dV$

$W = P[V]_{V_i}^{V_f}$

$W = P[V_f - V_i]$. → ①

$W = P\Delta V$ → ②.

Where,

ΔV - change in volume.

If ΔV is negative, W is also

ΔV is positive. W is also

from ideal gas equation.

$PV = \mu RT$

$$V = \frac{\mu RT}{P} \rightarrow (3)$$

from eqn (3).

$$V_i = \frac{\mu RT_i}{P} \rightarrow (4).$$

$$V_f = \frac{\mu RT_f}{P} \rightarrow (5).$$

Sub. eqn (4) & (5) in (1)

$$W = P\left[\frac{\mu RT_f}{P} - \frac{\mu RT_i}{P}\right]$$

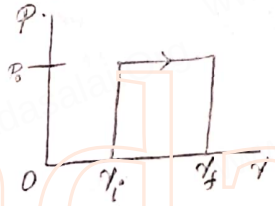
$$W = \frac{P}{P}\left[\mu RT_f - \mu RT_i\right].$$

$$W = \left[\mu RT_f - \mu RT_i\right]$$

$$= \mu R\left[T_f - T_i\right]$$

$$W = \mu RT_f\left[1 - \frac{T_i}{T_f}\right] \rightarrow (6)$$

PV diagram for isobaric process:



from first law of thermodynamics

$$\Delta U = Q - W. \rightarrow (7).$$

Sub. eqn. (6) in (7).

$$\boxed{\Delta U = Q - P\Delta V}$$

Example 8.20: $\gamma = \frac{5}{3}$, $R = 8.3\,J$ $T_i = 27°C$

a). (1) In adiabatic Process.

$$W = \frac{\mu R}{\gamma - 1}\left[T_i - T_f\right].$$

$T_f = ?$

We know that,

$$T_f V_f^{\gamma-1} = T_i V_i^{\gamma-1}$$

$$T_f = T_i\left(\frac{V_i}{V_f}\right)^{\gamma-1}$$

$$= 300\left(\frac{V}{2V}\right)^{\frac{5}{3}-1}$$

$$= 300\left(\frac{1}{2}\right)^{2/3}$$

$$= 300 \times (0.5)^{0.6}.$$

Here, $x = (0.5)^{0.6}$

Take log on both sides

$$\log x = \log(0.5)^{0.6}$$

$$\log x = 0.6 \log(0.5)$$

$$\log x = 0.6 \times -0.30$$

$$\log x = -0.1806$$

$(0.5)^{0.6}$

$0.6 \times -0.30.$

(II) Isobaric process:

$W = P\Delta V$

$= P(V_f - V_i)$.

$V_f = 2V_i$

$W = P(2V_i - V_i)$.

$W = P(V_i)$.

$PV = \mu RT$

Here, $\mu = 1$

$PV = RT$

then,

$P_i V_i = RT_i$

$V_i = \frac{RT_i}{P_i}$

$= \frac{8.3 \times 300}{1 \times 10^5}$

$= 8.3 \times 300 \times 10^{-5}$

$V_i = 24.9 \times 10^{-3} m^3$

$W = P \times V_i$

$= 1 \times 10^5 \times 24.9 \times 10^{-3}$

$W = 2490 J$

$W = 2490 kJ$

(III) Isothermal Process.

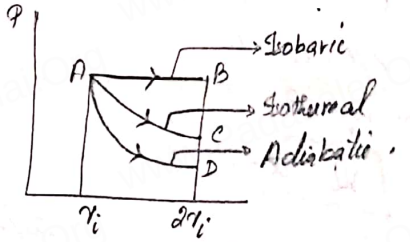
$W = \mu RT \ln\left(\frac{V_f}{V_i}\right)$

$= 1 \times 8.3 \times 300 \ln\left(\frac{2}{1}\right)$

$= 8.3 \times 3 \times 10^2 \times \ln(2)$

$= 24.9 \times 10^2 \times$

C) PV diagram:



Isochoric Process:

This a thermodynamic process in which the volume of the system is kept constant. But Pressure, temperature and internal energy to continue to be variables.

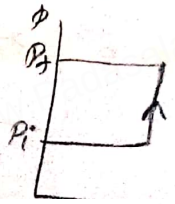
The equation of state for an isochoric process is given by,

$PV = \mu RT$

$P = \frac{\mu RT}{V}$

Here, $\frac{\mu R}{V}$ = constant.

Increased Pressure



If a gas goes from state (P_i, V_i) to (P_f, V_f) at constant volume, then the system satisfies the equation.

$P \propto$ constant $\times T$

$P \propto T$

$\frac{P_i}{T_i}$ = constant $\propto Q$.

$\frac{P_f}{T_f}$ = constant $\propto Q$

from (1) & (2).

$$\frac{P_i}{T_i} = \frac{P_f}{T_f}$$

from first law of thermodynamics.

$$\Delta U = Q - pV$$

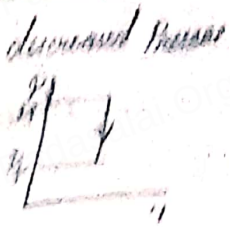
Here,

$\Delta V = 0$

$W = 0$

$$\boxed{\Delta U = Q.}$$

decreased pressure



Example 8.21

We know that,

$$\Delta U = Q - W$$

Here $W = 0$. (Isochoric process).

$$\Delta U = Q.$$

We know that.

$$S_H = \frac{1}{m} \frac{\Delta Q}{\Delta T}$$

$$S_H = \frac{1}{m} \cdot \frac{Q}{\Delta T}$$

$$Q = S_H \cdot m \cdot \Delta T.$$

$$Q = m \cdot S_H \cdot \Delta T$$

mass of water = 500g = 0.5 kg

Change in temperature = ~~293K~~ ~~20°C~~ - 303°C

= 20°C

Then, heat,

$Q = 0.5 \times 4184 \times 20$

$= 4184 \times 2 \times 5 \times 10^1 \times 10^1$

$= 62760$ J

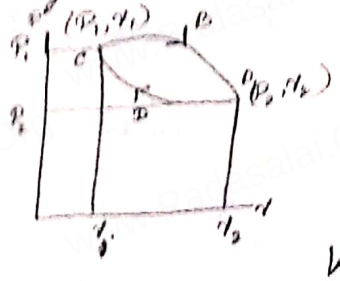
$= 62.760$ kJ

Cyclic Process:

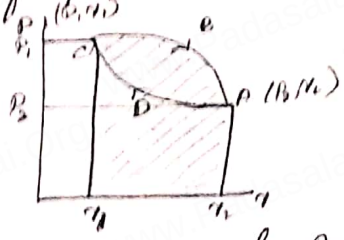
Cyclic Process is a thermodynamic process in which the thermodynamic system return to its initial state after undergoing a series of changes.

$Q_{net} = Q_{in} - Q_{out} = W$

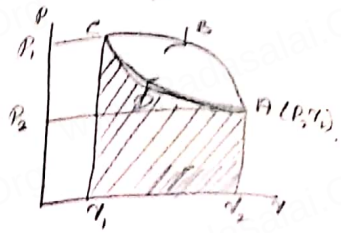
PV diagram for a cyclic Process:



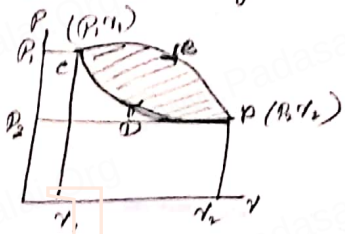
Work done for path CBA.



Work done for path ADC



Net work done in a cyclic Process:



Example 2.22.

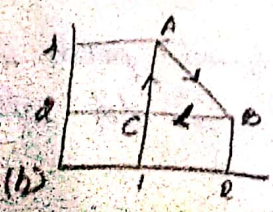
(a)

For case (a) it is a anticlockwise. so, the net work done is negative.

Area under curve BC = 1×4
= -4J

Area under curve DA = 1×2
= 2J

Net work done = -4+2
= -2J.



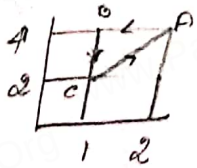
(b)

For case (b) it is in clockwise. so, the net work done is positive.

Area under curve AB,
$= \left(\frac{1}{2} \times 1 \times 2\right) + (1 \times 2)$
$= 1 + 2$
$= 3J$.

Area under curve BC,
$= 1 \times 2 = -2J$.

Net work done = 3-2
= 1J.



It is in anticlockwise so, it is negative.

Area under the curve AB
$= 4 \times 1 = -4J$

Area under curve CA
$= 1 \times 2 + \frac{1}{2} \times 1 \times 2$
$= 2 + 1 = 3J$.

Net work done = -4+3 = -1J

Reversible Process:

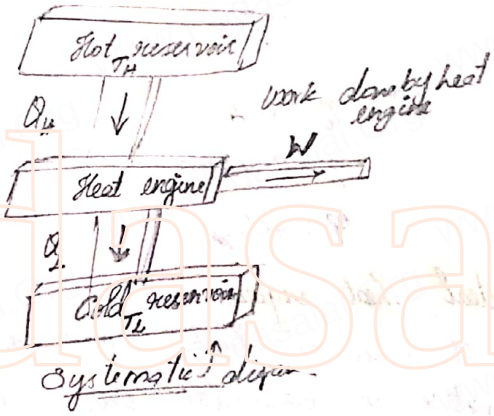
A thermodynamic process can be considered reversibly if it possible to retrace the path in the opposite direction in such a way that the system and surrounding pass through the same states as in the initial, direct process.

Irreversible process:

All natural process are irreversible. Irreversible process cannot be plotted in a PV diagram, because these process cannot have unique values of pressure, temperature at every stage of the process.

Heat engine:

Heat engine is a device which takes heat as input and converts this heat in to work by undergoing a cyclic process.



Systematic diagram

After doing Work.

Hot reservoir at temperature T_H

Q_H

↓

Heat engine ⟶ W

↓

Q_L

Cold reservoir at temperature T_L.

We can write.

Input heat = Work done + ejected heat.

$$Q_H = W + Q_L$$

$$W = Q_H - Q_L$$

Then, the efficiency of heat engine.

$$\eta = \frac{\text{output}}{\text{Input}} = \frac{W}{Q_H} = \frac{Q_H - Q_L}{Q_H}$$

$$= 1 - \frac{Q_L}{Q_H}$$

Kelvin - Planck statement:

It is impossible to construct a heat engine that operates in a cycle whose sole effect is to convert the heat completely into work. This implies that no heat engine in the universe can have 100% efficiency.

Example 8.84.

Efficiency of heat engine,

$$\eta = 1 - \frac{Q_L}{Q_H}$$

$$= 1 - \frac{300}{500}$$

$$= 1 - \frac{3}{5}$$

$$= \frac{5-3}{5}$$

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

$$\eta = 0.4$$

$$\eta = 0.4 \times 100$$

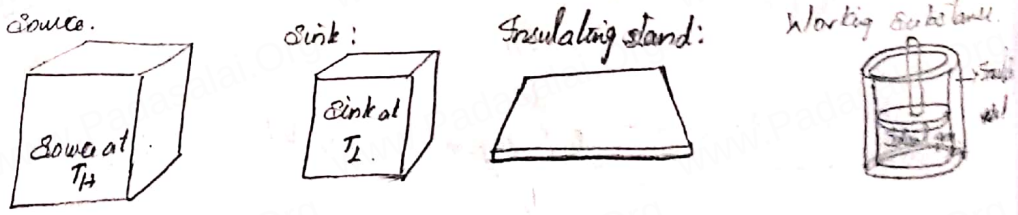
$$\eta = 40\%$$

Carnot's ideal heat engine:

A reversible heat engine, operating in a cycle b/w two temperature in particular way is called a Carnot Engine.

It has four parts:

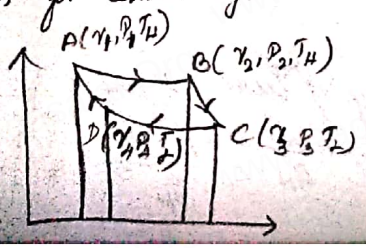
(i) Source (ii) Sink (iii) Insulating stand (iv) Working substance



Carnot's cycle:

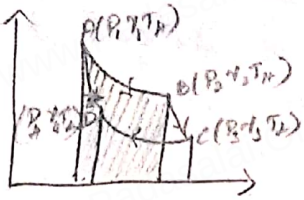
The working substance is subjected to four successive reversible process forming is called Carnot's cycle.

PV diagram for Carnot cycle:



Work done in Carnot cycle:

Isothermal process (Expansion):



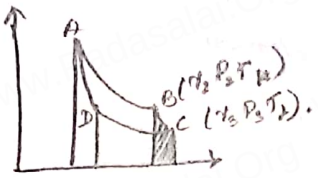
Then, the work done by the gas in isothermal expansion

$$Q_H = W_{A\to B} = \int_{V_1}^{V_2} P\,dV.$$

$$W_{A\to B} = \mu R T_H \ln\left(\frac{V_2}{V_1}\right)$$

= Area under the curve AB.

Adiabatic process: (Expansion).



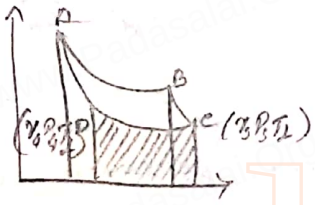
Then, the work done by the gas.

$$W_{B\to C} = \int_{V_2}^{V_3} P\,dV$$

$$W_{B\to C} = \frac{\mu R}{\gamma - 1}\left[T_H - T_L\right]$$

= Area under curve BC.

Isotherm Process: (Compression)



Then, the work done by the gas.

$$W_{C\to D} = \int_{V_3}^{V_4} P\,dV$$

$$= \mu R T_L \ln\left(\frac{V_4}{V_3}\right) = -\mu R T \ln\left(\frac{V_3}{V_4}\right)$$

= Area under the curve CD.

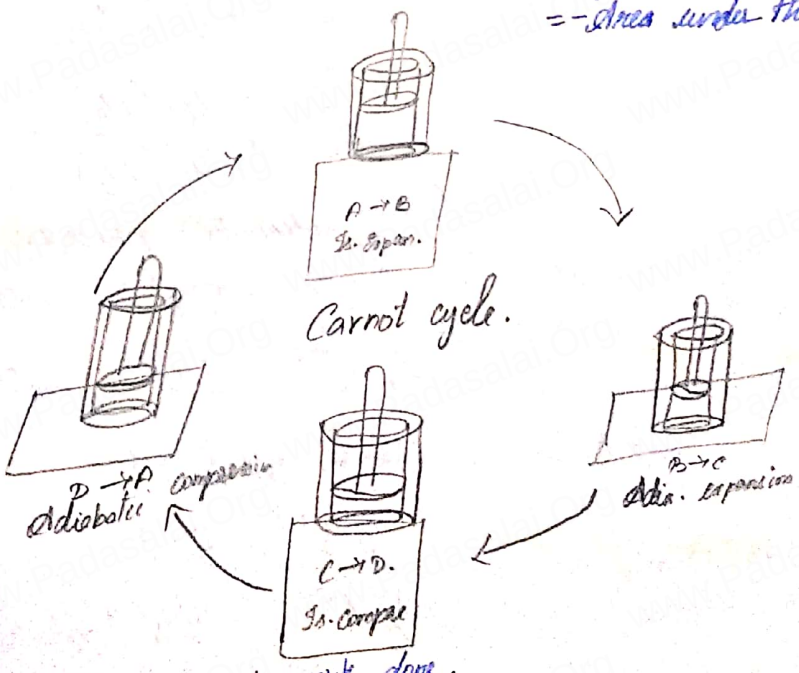
Adiabatic Process: (Compression).

Then, the work done by the gas,

$$W_{D\to A} = \int_{V_4}^{V_1} P\,dV$$

$$= \frac{\mu R}{\gamma - 1}\left(T_L - T_H\right).$$

= - Area under the curve DA.



Let, W - be the net work done.

$$W = W_{A\to B} + W_{B\to C} - W_{C\to D} - W_{D\to A}$$

Since $[W_{B\to C} = W_{D\to A}]$.

$$W = W_{A\to B} + W_{B\to C} - W_{C\to D} - W_{B\to C}$$

$$W = W_{A\to B} - W_{C\to D}.$$

Efficiency of a Carnot engine:

Efficiency is defined as the ratio of work done by the working substance in one cycle to the amount of heat extracted from the source.

$$\eta = \frac{\text{work done}}{\text{heat extracted}}$$

$$\eta = \frac{W}{Q_H} \rightarrow (1)$$

From, first law of thermodynamics.

$$W = Q_H - Q_L$$

We know that,

Efficiency of heat engine, $\eta = \frac{Q_H - Q_L}{Q_H}$

$$\eta = \frac{Q_H}{Q_H} - \frac{Q_L}{Q_H}$$

$$\eta = 1 - \frac{Q_L}{Q_H} \rightarrow (1).$$

During Isothermal process.

$$Q_H = \mu R T_H \log\left(\frac{V_2}{V_1}\right) \rightarrow (2).$$

$$Q_L = \mu R T_L \log\left(\frac{V_3}{V_4}\right) \rightarrow (3).$$

Divide equ (2) by (3).

$$\frac{Q_L}{Q_H} = \frac{\mu R T_L \log(V_3/V_4)}{\mu R T_H \log(V_2/V_1)}$$

$$\frac{Q_L}{Q_H} = \frac{T_L \log(V_3/V_4)}{T_H \log(V_2/V_1)} \rightarrow (4).$$

During adiabatic process.

$$T_H V_2^{\gamma-1} = T_L V_3^{\gamma-1} \rightarrow (5).$$

$$T_H V_1^{\gamma-1} = T_L V_4^{\gamma-1} \rightarrow (6).$$

Divide equ (5) by (6).

$$\frac{T_H V_2^{\gamma-1}}{T_H V_1^{\gamma-1}} = \frac{T_L V_3^{\gamma-1}}{T_L V_4^{\gamma-1}}$$

$$\frac{V_2^{\gamma-1}}{V_1^{\gamma-1}} = \frac{V_3^{\gamma-1}}{V_4^{\gamma-1}}$$

$$\left(\frac{v_2}{v_1}\right)^{\gamma-1} = \left(\frac{v_3}{v_4}\right)^{\gamma-1} \rightarrow (3).$$

from eqn (3),

We can say $\frac{v_2}{v_1} = \frac{v_3}{v_4}$ $\rightarrow (4)$.

Sub eqn (4) in (4).

$$\frac{Q_L}{Q_H} = \frac{T_L \log_e(v_3/v_4)}{T_H \log_e(v_3/v_4)}$$

$$\boxed{\frac{Q_L}{Q_H} = \frac{T_L}{T_H}} \rightarrow (5).$$

Sub. eqn (5) in (1).

$$\boxed{\eta = 1 - \frac{T_L}{T_H}}$$

The temperature value T_L and T_H should be in kelvin scale.

Example 8.25

a). $T_L = 300K$

$T_H = 523K$.

$$\eta = 1 - \frac{T_L}{T_H} = 1 - \frac{300K}{523K}.$$

$$= \frac{523 - 300}{523}$$

$$= \frac{223}{523}$$

$$\eta = 0.43\%$$

Second law of thermodynamics:

It states that "in all energy exchanges, if no energy enters (or) leave the system, the potential energy of the state will always be less than that of the initial state". This is also commonly referred to as entropy.

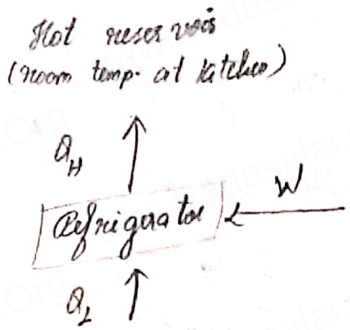
Entropy:

Entropy is also called 'measure of disorder'. All natural process occurs such that the disorder should always increase.

Refrigerator :

Working Principle:

(a)



from 1st law of thermodynamics

$$Q_H - Q_L = W \rightarrow (1).$$

$$Q_H = W + Q_H$$

Co-efficient of performance (COP)(β):

COP is a measure of the efficiency of a refrigerator. It is defined as the ratio of heat extracted from the cold body (sink) to the external work done by the compressor W.

$$\beta = \frac{Q_L}{W} \rightarrow (2)$$

Sub eq(1) in (2).

$$\beta = \frac{Q_L}{Q_H - Q_L}$$

$$= \frac{Q_L}{Q_L\left[\frac{Q_H}{Q_L} - 1\right]}$$

$$\beta = \frac{1}{\left[\frac{Q_H}{Q_L} - 1\right]}$$

We know that,

$$\frac{Q_H}{Q_L} = \frac{T_H}{T_L}$$

$$\beta = \frac{1}{\left[\frac{T_H}{T_L} - 1\right]}$$

$$\beta = \frac{1}{T_H - T_L} \Rightarrow \beta = \frac{T_L}{T - T_L}$$