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→ Macroscopic :- (Classical Thermodynamics)

In macroscopic approach, a certain quantity of matter is considered, without the events occurring at the molecular level being taken into account. Macroscopic thermodynamics is only concerned with the effects of the action of many molecules, and these effects can be perceived/observed by human senses. For example, the macroscopic quantity, pressure, is the avg. rate of change of momentum due to all the molecular collisions made on an unit area. The effects of pressure can be felt. The macroscopic point of view is not concerned with the action of individual molecules, and the force on a given unit area can be measured by using e.g. a pressure gauge. These macroscopic observations are completely independent of the assumptions regarding the nature of matter. All the results of classical or macroscopic thermodynamics can, however, be derived from the ~~and~~ microscopic ~~and~~ or statistical study of matter.

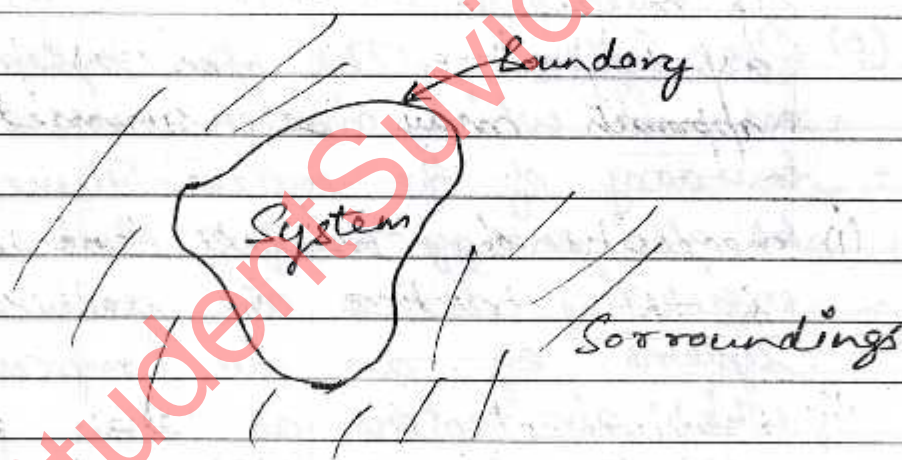
thermodynamics) \Rightarrow Microscopic: (Statistical Thermodynamics)

From the microscopic point of view, matter is composed of myriads of molecules. If it is a gas, each molecule, these change very frequently as a result of collisions. The behaviour of the gas is described by summing up the behaviour of each molecule. Such study is made in microscopic or statistical thermodynamics. The salient aspects/features of microscopic approach may be summed up as;

- (i) necessity of complete knowledge of the structure of matter,
- (ii) Requirement of a large no. of variables for complete specification of the state of matter.
- (iii) Easy and precise measurement is not possible of the variables used to describe the state of matter.

2) Thermodynamic System:-

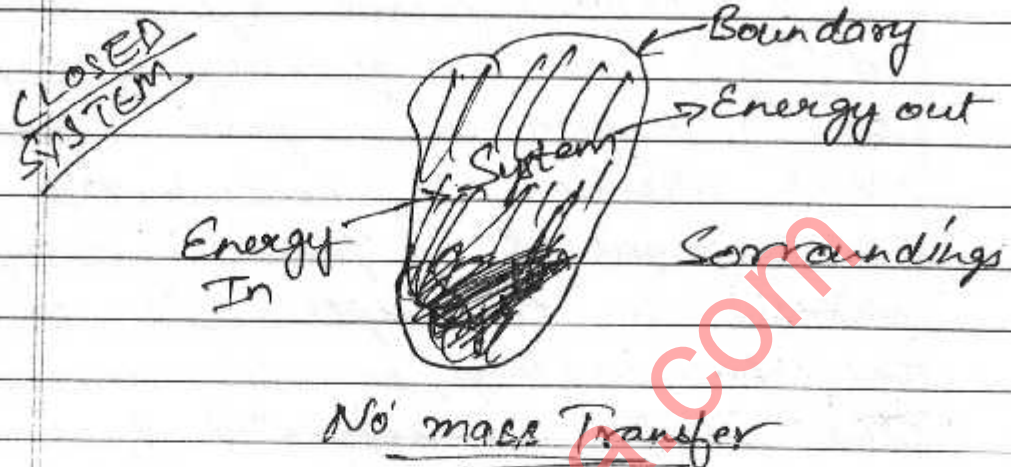
A thermodynamic system is defined as a quantity of matter or a region in space upon which attention is concentrated in the analysis of a problem. Everything external to the system is called the surroundings or the environment. The system is separated from the surroundings by the system boundary.



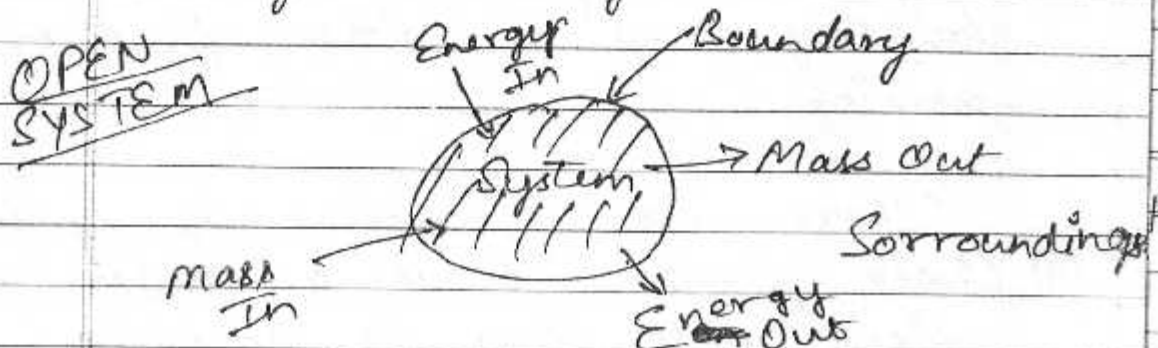
A system and its surroundings together comprise a universe. The boundary may be either fixed or moving.

There are 3 classes of systems;
(a) Closed System:- It is a system of fixed mass. There is no mass transfer across the system boundary.

There maybe energy transfer into or out of the system



(b) Open System:- The open system is one in which matter crosses the boundary of the system. There may be also energy transfer. Most of the engineering devices are generally open systems, eg. an air compressor in which air enters at low pressure and leaves at high pressure and there are energy transfers across the system boundary.



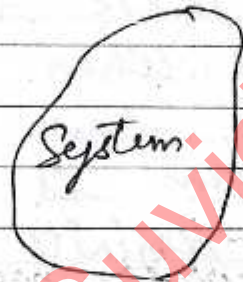
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(c) Isolated System:- It is the one in which there is no interaction between the system and the surroundings. It is of fixed mass and energy, and there is no mass or energy transfer across the system boundary.

ISOLATED
SYSTEM



Surroundings

⇒ Thermodynamic Properties, Process And Cycles:-

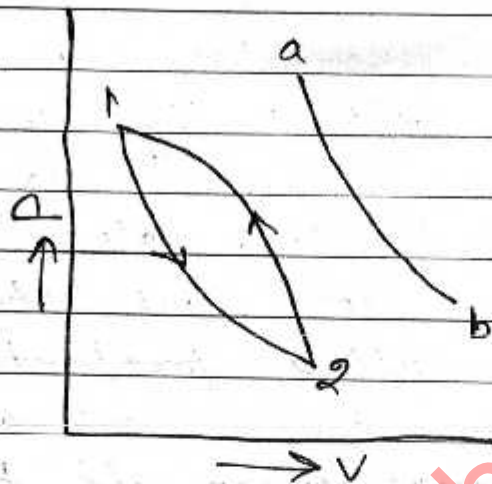
Every system has certain characteristics by which its physical condition may be described, e.g. volume, temperature, pressure, etc. Such characteristics are called properties of a system. These are all macroscopic in nature. When all properties of a system have definite values, the system is said to exist at a definite state.

Properties are the coordinates to describe the state of a system. They are the state variables of the system. Any operation in which one or more of the properties of a system changes is called a change of state.

The succession of states passed through during a change of state is called the path of the change of state. When the path is completely specified, the change of state is called a process.

A thermodynamic ~~process~~ / cycle is defined as a series of state

changes such that the final state is identical with the initial state.



$a-b = A \text{ process}$

$1-2-1 = A \text{ cycle}$

Properties can may be of two types;

- 1) Intensive :- These properties are independent of the mass in system, ~~ex. volume, energy etc.~~
ex. pressure, temperature etc.

Its value remains the same whether one considers the whole system or only a part of it.

- 2) Extensive :- Depends on mass of the system, if the mass is increased, the values of the extensive properties also increase. Ex. Specific volume, specific density etc.

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THERMODYNAMIC

EQUILIBRIUMS

A thermodynamic system is said to be in thermodynamic equilibrium when it is in thermal equilibrium, mechanical equilibrium, radiative equilibrium and chemical equilibrium. The word equilibrium means a state of balance. In an equilibrium state, there are no unbalanced potentials (or driving forces) within the system.

→ Conditions for Equilibrium:-

By considering the differential form of thermodynamic potential, the following relations can be derived:-

- * For a complete isolated system, $\Delta S = 0$ at equilibrium.
- * For a system at constant temperature & volume, $\Delta A = 0$ at equilibrium.
- * For a system at constant temperature and pressure, $\Delta G = 0$ at equilibrium.

Various types of equilibriums are achieved as follows:-

- * Two systems are in thermal equilibrium when their temperatures are the same.
- * Two systems are in mechanical

[Equilibrium = Equil.]

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* The two systems are in diffusive equilibrium when their chemical potentials are the same.

⇒ TYPES OF EQUILIBRIUM:-

→ THERMAL EQUILIBRIUM:-

Thermal equil. is achieved when two systems ^{are} in thermal contact* with each other cease to have a net exchange of energy. It follows that if two systems are in thermal equil., then their temperatures are the same.

Thermal equil. occurs when a system's macroscopic* thermal observables have ceased to change with time. For example, an ideal gas* whose distribution function has stabilised to a specific Maxwell-Boltzmann distribution would be in thermal equilibrium. This outcome allows a single temperature & pressure to be attributed to the whole system. Thermal equil. of a system does not imply absolute uniformity within a system; for example, a river system can be in thermal

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equil. when the macroscopic temp. distribution is stable and not changing in time, even though the spatial temperature distribution exhibits thermal fluctuations.

TERMS :-

→ Thermal Contact: A thermodynamic system is said to be in thermal contact with another system if it can exchange energy with it through the process of heat.

→ Macroscopic: The macroscopic scale is the length scale on which objects or process are of size which is measurable and observable by naked eyes. Ex: A macroscopic view of a ball is just that: a ball.

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QUASISTATIC PROCESS

OR Quasistatic Equilibrium

[QUASI = ALMOST]

In thermodynamics, a quasi-static process is a thermodynamic process that happens infinitely slowly. However, it is important to note that no real process is quasistatic. Therefore in practice, such processes can only be approximated by performing them infinitely slowly.

A quasistatic process often ensures that the system will go through a sequence of states that are infinitely close to equilibrium* (so the system remains in quasistatic equilibrium), in which case the process is typically reversible. An example of a quasistatic process that is not reversible is the slow heat exchange between two bodies at two finitely different temperatures, where the heat exchange rate is controlled

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by an approximately adiabatic* partition between the two bodies. In this case, no matter how slowly the process takes place, the states of the two bodies are never infinitely close to equilibrium, since thermal equilibrium requires that the two bodies be at precisely the same temperature.

← THERMODYNAMIC LAWS →

⇒ Zeroth Law: The zeroth law states that if two systems are in thermal equilibrium with a third system, then they are also in thermal equilibrium with each other.

$$\text{If } T(A) = T(B)$$

$$\text{And } T(B) = T(C)$$

$$\text{Then } T(A) = T(C)$$

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state
state.

⇒ Reversible And Irreversible Process

REVERSIBLE:-

A thermodynamic process is reversible if the system passes through a continuous series of equilibrium states.

The following conditions need to be satisfied for a process to be reversible:-

- * There should be no friction; solid or fluid.
- * The heat exchange to or from the system, if any, should be only through infinitely small temperature difference.
- * The process should be quasi-static, it should proceed at infinitely slow speed.

⇒ FEATURES:-

- * A quasi-static process which can be carried out in the reverse direction along the same path.
- * It can proceed in either direction without violating the second law of thermodynamics.
- * No wastage of energy occurs in a reversible process.
- * A reversible process yields the max work in engines, & requires min work in devices such as refrigerators, pumps & compressors.

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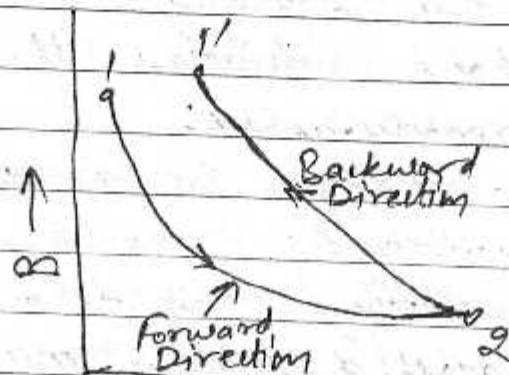
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Some examples of reversible process:-

- Motion without friction
- Restricted and controlled expansion or compression.
- Isothermal and frictionless adiabatic process
- Elastic stretching of a solid.
- Restrained discharge of a battery.
- Electric circuit without zero resistance.

IRREVERSIBLE:-

A process is irreversible if the system passes through a sequence of non-equilibrium states. During such a process, the properties of a system do not have unique value and accordingly cannot be located on any property diagram. When an irreversible process is made to proceed in the backward direction, the original state of the system is not restored.



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A reversible process can be identified by the following characteristics;

- It can be carried out in one direction only.
- It occurs at a finite rate.
- It cannot be reversed without causing permanent changes in the surroundings,
- The system is never in equilibrium state at any instant during an irreversible process.

→ Examples of irreversible process:-

- Spontaneous chemical reaction.
- Viscous Flow; Fluid Flow with friction.
- Plastic deformation
- Diffusion of gases.
- Energy transfer as heat with finite temperature differences.
- Free expansion and throttling process.

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→ First Law:- This law states that energy can be transformed, i.e. changed from one form to another, but cannot be created nor destroyed.

It is usually formulated by stating that the change in the internal energy of a system is equal to the amount of heat supplied to the system, minus the amount of work performed by the system on its surroundings.

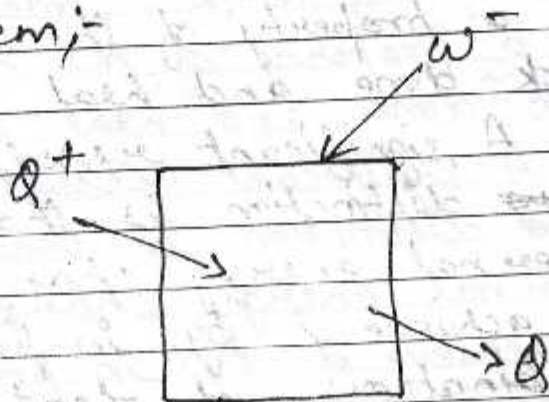
$$\text{i.e., } dU = \delta Q - \delta W$$

where $\delta Q = \text{heat}$

& $\delta W = \text{work}$

Note that the minus sign in front of δW indicates that a true amt. of work done by the system leads to energy being lost from the system.

* Sign Convention of heat & work in a system:-



When a system expands in a quasistatic process, the work done on the environment is the product of pressure (P) and volume (V) change, i.e. PdV , whereas the work done on the system is:-

$$dU = \delta Q - PdV.$$

Work and heat are processes which add or subtract energy, while U is a particular form of energy associated with the system. Thus the term heat for δQ means that amount of energy added as the result of heating, rather than referring to a particular form of energy. Likewise, work energy for δW means "that amount of energy lost as the result of work". Internal energy is a property of the system whereas work done and heat supplied are not. A significant result of this ~~less~~ distinction is that a given internal energy change (dU) can be achieved by, in principle, many combinations of heat and work.

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⇒ I) First Law of thermodynamics applied to a closed system

Consider a closed system where there is no flow into or out of the system, and the fluid mass remains constant. For such system, the first law statement is known as the Non-flow Energy Equation, or NFEE, which can be summarized as follows:

$$\Delta U = Q - W \quad \text{--- (i)}$$

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⇒ II) Internal Energy

Internal energy (I.E.) is defined as the energy associated with the random, disordered motion of molecules. It is separated in scale from the macroscopic ordered energy associated with moving objects; it refers to the invisible microscopic energy on the atomic and molecular scale. For example, at room temperature a glass of water sitting on a table has

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no apparent energy, either potential or kinetic. But on the microscopic scale it is a seething mass of high speed molecules travelling at hundreds of meters per second. If the water were tossed across the room, this microscopic energy would not necessarily be changed when we superimpose an ordered large scale motion on the water as a whole.

During a non-flow process, the change in internal energy is calculated assuming the closed system volume remain constant, the following equation is used;

$$\Delta U = m \cdot C_v \cdot \Delta T \quad \text{---(ii)}$$

where C_v is the specific heat capacity of the fluid, and ΔT is the temperature difference during the process.

→ I (b)] Specific Heat:

Heat maybe defined as energy in transit from a high temperature object to a

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lower temperature object. An object does not possess "heat"; the appropriate term for the microscopic energy in an object is internal energy. The internal energy may be increased or raised by transferring energy to the object from a higher temperature (hotter) object - this is properly called heating.

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Thus, Specific Heat may be defined as the ~~spe~~ amount of heat per unit mass required to raise the temp. by one degree Celsius.

In order to heat or cool a given quantity of a gas in a given time, the following equation is used:-

-(iii)
heat
AT
ence

$$\text{Quantity of Heat} = \text{mass} \times \text{specific heat capacity} \times \text{temperature difference}$$

or

$$Q = m \cdot C \cdot (T_2 - T_1)$$

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a
to a

Since this heat exchange may take place

(21)

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Either at constant pressure :-

$$Q = m C_p (T_2 - T_1) \quad (iii)$$

Or at constant volume :-

$$Q = m C_v (T_2 - T_1) \quad (iv)$$

Where;

$C_p \Rightarrow$ Specific heat at constant pressure
(kJ/kgK)

and

$C_v \Rightarrow$ Specific heat at constant volume
(kJ/kgK)

T(b)(ii) =

Equation for a perfect gas,

$$C_p = C_v + R \quad \text{and} \quad \gamma = \frac{C_p}{C_v} \quad (v)$$

I(b)(i) Specific Heat at Constant Volume (C_v) :-

Consider a closed system of unit mass, the first law of thermodynamics applied to this system is :-

$$dU = Q - W$$

$$\text{or} \quad du = q - w.$$

If the volume is constant, then

$$W=0, \text{ therefore; } q=du$$

(iii)

$$\begin{aligned} \text{But } q &= C_v dT \\ \text{Hence } du &= C_v dT \\ \text{or } C_v &= \frac{du}{dT} \end{aligned}$$

(iv)

pressure

This is known as Joule's Law of internal energy which states that "the internal energy of a perfect gas depends upon its temperature only".

1. (b) (ii) \Rightarrow Specific Heat at Constant Pressure (C_p):

volume.

Consider a constant pressure non-flow process, the first law:-

$$du = q - w$$

(v)

$$\begin{aligned} \text{For a constant pressure process;} \\ W &= p(V_2 - V_1) \\ &= P_2 V_2 - P_1 V_1 \end{aligned}$$

unit (C_v):

Hence;

unit

thermodyn-
is :-

$$\begin{aligned} du &= q - (P_2 V_2 - P_1 V_1) \\ \text{or } u_2 - u_1 &= q - (P_2 V_2 - P_1 V_1) \\ \text{or } q &= (u_2 - u_1) + (P_2 V_2 - P_1 V_1) \end{aligned}$$

$$= u_2 - u_1 + P_2 V_2 - P_1 V_1$$

$$= u_2 + P_2 V_2 - u_1 - P_1 V_1$$

$$= (u_2 + P_2 V_2) - (u_1 + P_1 V_1)$$

but, $q = C_p dT$

hence,

$$h_2 - h_1 = C_p (T_2 - T_1)$$

or $C_p = \frac{dh}{dT}$

→ Relation between specific heats;

Since $H = U + PV$

$$\therefore dH = dU + d(PV)$$

$$d(PV) = d(mRT) = mRdT$$

$$m C_p dT = m C_v dT + m R dT$$

Therefore,

$$C_p = C_v + R \quad (\text{i.e. } C_p > C_v)$$

The ratio;

$\frac{C_p}{C_v} = \gamma$ is called the adiabatic index.

The reason for the difference between C_p & C_v is that in the constant pressure process part of the heat transferred is used in doing work against the moving system boundary and all heat therefore is not available for raising the gas temperature.

→ I.C.]

(a) For

(b) For
here is

→ I(c) } System Work

Work performed on or by the working fluid within a system's boundary is defined as the summation (or integrating) of the product of pressure and volume of the fluid during a process.

$$W = \int P \cdot dV \quad (vi)$$

In calculating the process work, it is important to point out that for each process, the work done will be different, since there are four distinctly different processes. In the following sections, an expression for the work done will be evaluated for each process.

(a) For a constant pressure process, the work during an isobaric process is simply:

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

(b) For a constant volume process, $dV = 0$; hence the work during an isochoric process is simply:-

$$W = 0$$

(c) For a constant temperature, isothermal process, $PV = C$

II)
(a)

$$\text{hence } W = \int \frac{C}{V} dV = C \ln \frac{V_2}{V_1}$$

$$\text{but } P_1 V_1 = P_2 V_2 \therefore \frac{V_2}{V_1} = \frac{P_1}{P_2}$$

Hence the work done can be written in terms of pressure ratio:

$$\therefore W = mRT \ln \frac{P_1}{P_2}$$

(d) For an adiabatic (Polytropic) process,
 $PV^n = C$

$$\therefore W = \frac{P_1 V_1 - P_2 V_2}{n-1}$$

For ideal gas relation, the work done can be written as;

$$W = \frac{mR(T_1 - T_2)}{n-1}$$

II) First Law of Thermodynamics applied to a Close System (cycle);

Since in a cycle the working fluid undergoing changes in its state will retain its initial conditions at any fixed point along the cycle. Hence the energy equation applied to a cycle is:-

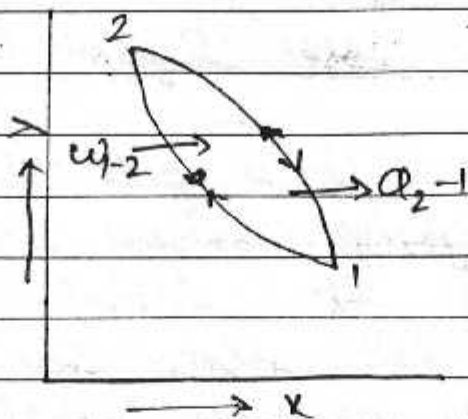
$$\boxed{\sum Q = \sum W}$$

where Σ means the sum of heat or work around the cycle.

The expression $(\sum W) = \int (\sum Q)$ applies only to systems undergoing cycles of the algebraic summation of all energy transfer across system boundaries is zero.

This is also expressed in the form:-
 $\oint dW = \int \delta Q$

(\int = Joule's Equivalent)
 (\oint = cyclic integral for the closed path)



Above system is having a cycle, which consists of a definite amount of work input W_{1-2} to the system followed by the transfer of an amount of heat Q_{2-1} from the system. It has been found that this W_{1-2} is always proportional to heat Q_{2-1} , i.e.,

$$W_{1-2} \propto Q_{2-1}$$

II (b)

and the constant of proportionality is called the Joule's Equivalent.

In the above system, there are only two energy transfer quantities as the system performs a thermodynamic cycle. If these the cycle

involves many more heat & work quantities, the same result will be found which can be expressed algebraically as:-

$$(\sum w)_{\text{cycle}} = \sum (\sum Q)_{\text{cycle}}$$

Practical Application of a closed system (cycle) - Assume compression & expansion to be adiabatic, from first law:-

$$Q_s - Q_R = W_e - W_c$$

Where, Q_s = Heat supplied

Q_R = Heat rejected

W_e = work of expansion

W_c = work of compression

II (b) First law for a closed system undergoing a change of state (Open System):-

If a system undergoes change of state during which both heat transfer and work transfer are involved, the net energy transfer will be stored within the system. If Q is the amt. of heat transferred

to the system and W is the amount of work transferred from the system during the process, the net energy transfer ($Q - W$) will be stored in the system. Energy in storage is neither heat nor work, therefore

(III)

Therefore;

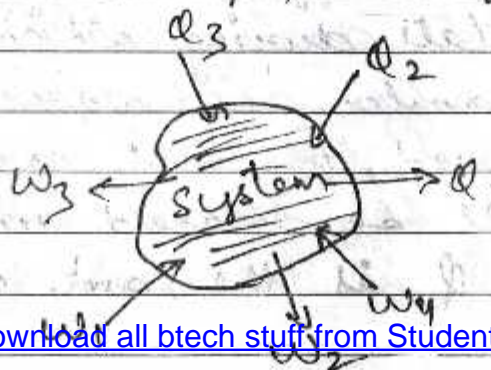
$$\begin{aligned}
 & Q - W = \Delta E \\
 \text{or } & Q - W = \Delta H + \Delta KE + \Delta PE \\
 \text{or } & \rightarrow Q = \Delta E + W \\
 \text{or } & \rightarrow Q - W = m \left[(h_2 - h_1) + \frac{1}{2}(v_2^2 - v_1^2) + g(z_2 - z_1) \right]
 \end{aligned}$$

* Above equation is also known as Steady flow Energy Equation (SFEE)

Here Q , W and ΔE are all expressed in the same units (in Joules). Energy can be stored by a system in different modes.

If there are more energy transfer quantities involved in the process, ~~the first law given is like~~ in below example, the first gives:-

III (a) → Ste



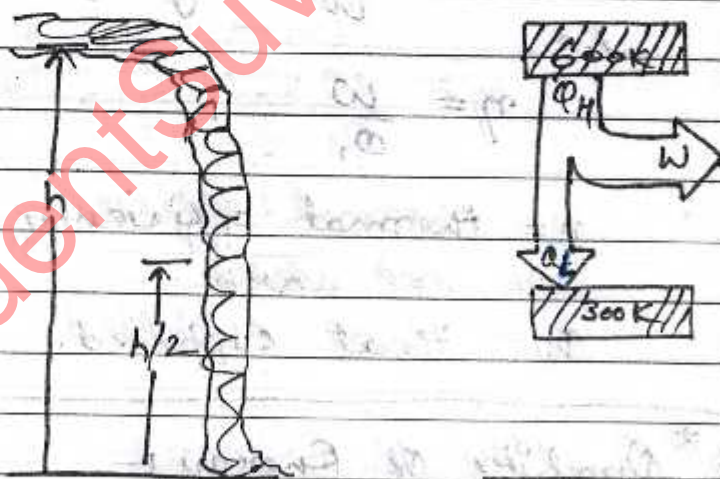
out
system
in
is

$$Q = \Delta E + W$$

$$\text{or } (Q_2 + Q_3 - Q_1) = \Delta E + (W_2 + W_3 - W_1 - W_4)$$

(III) The Second Law of Thermodynamics is

The second law of thermodynamics is a general principle which places constraints upon the direction of heat transfer and the attainable efficiencies of heat engines. Its implications may be visualized in terms of the waterfall analogy.



III (a) Statements of Second Law:-

The second law indicates that, although the net heat supplied, $Q_1 - Q_2$, in a cycle is equal to the ~~net~~ net work done, (W) , the gross heat supplied, Q_1 , must be greater than the ~~net~~ net work

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done; some heat must always be rejected by the system.

$Q_1 > W$, or to be precise:

$$W = Q_1 - Q_2$$

where, Q_1 = Heat Supplied

Q_2 = Heat Rejected.

The ratio of net work output to the heat supplied is known as the thermal efficiency of the system; i.e.,

$$\eta = \frac{W}{Q_1}$$

η = thermal efficiency

W = net work

Q_1 = Heat supplied.

→ * Quality of Energy :

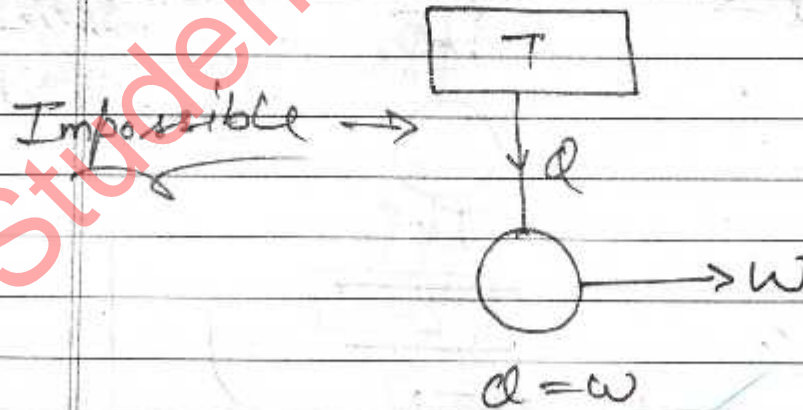
Work/mechanical energy is high grade energy whereas, heat/thermal energy is low grade energy. 100% conversion of high grade energy to low grade energy is possible whereas the converse is not true, due to the thermal imbalance of environment with the system.

There are two ways in which the second law is expressed:-

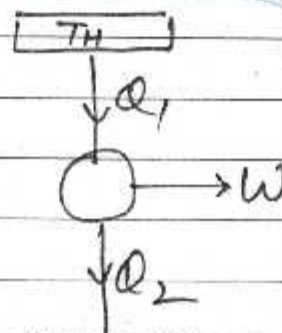
(a) Kelvin - Planck Statement:-

"It is impossible to construct a system which when operating in a cycle will extract heat energy from a single source and do an equal amount of work on the surroundings."

I.e. never possible to achieve 100% thermal efficiency, or, no heat engine can be 100% efficient.



Actual →



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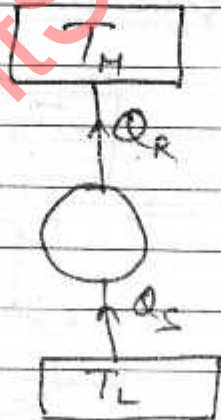
(b) Clausius Statement :-

III (ii) →

"It is impossible to construct a system which when operating in a thermodynamic cycle has no other effect than the transfer of heat energy from a cool (lower temp) to a hotter (higher temp) body without consuming any other form of energy (work)".

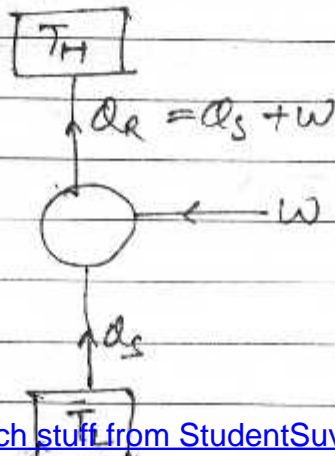
I.E, Some work is done on or by the working ~~substance~~ to the surroundings or vice versa.

Impossible →



Q_S = Heat Supplied
 Q_R = Heat Rejected

Statement :-
 Claim
 (Actual)



III (ii) →

ENTROPY :- (2)

Entropy is related to the randomness of the system which increases when heat is added to the system and decreases when heat is removed from the system.

$$ds = \frac{dQ}{T} = \frac{Q}{T} \quad [xi]$$

or by
roundings

- 1) Entropy can be further defined as; is a state variable whose change is defined for a reversible process at T where Q is the heat absorbed.
- 2) a measure of the amount of energy which is unavailable to do work.
- 3) a measure of the disorder of a system.

III (ii)
(A)

→ Change of Entropy for a perfect gas undergoing a process:-

First Law :-

$$\Delta U = Q - W$$

or $Q = W + \Delta U$

but $Q = T dp$

$$W = p dv$$

(From (xi))

$$\therefore Tds = pdv + CvdT$$

$$\text{or } ds = \frac{pdv}{T} + Cvd\frac{T}{T} \quad (\text{dividing by } T)$$

We know;

$$pv = RT$$

$$p = \frac{RT}{v}$$

$$\text{or } \frac{p}{T} = \frac{R}{v}$$

$$\therefore ds = \frac{Rdv}{v} + Cvd\frac{T}{T}$$

$$\text{or } ds = Cvd\frac{T}{T} + \frac{Rdv}{v}$$

$$S_2 - S_1 = C_v \ln \frac{T_2}{T_1} + R \ln \frac{v_2}{v_1}$$

Since,

$$\frac{P_1 v_1}{T_1} = \frac{P_2 v_2}{T_2}$$

$$\text{then, } \frac{v_2}{v_1} = \frac{P_1 T_2}{P_2 T_1}$$

$$\therefore S_2 - S_1 = C_v \ln \frac{T_2}{T_1} + R \ln \left(\frac{P_1}{P_2} \right) \left(\frac{T_2}{T_1} \right)$$

iding)
yT

Since,

$$R = C_p - C_v$$

or $C_p = R + C_v$

This can also be written in terms of volumes and pressures as:-

$$S_2 - S_1 = C_v \ln \left(\frac{P_2}{P_1} \right) + C_p \ln \left(\frac{V_2}{V_1} \right)$$

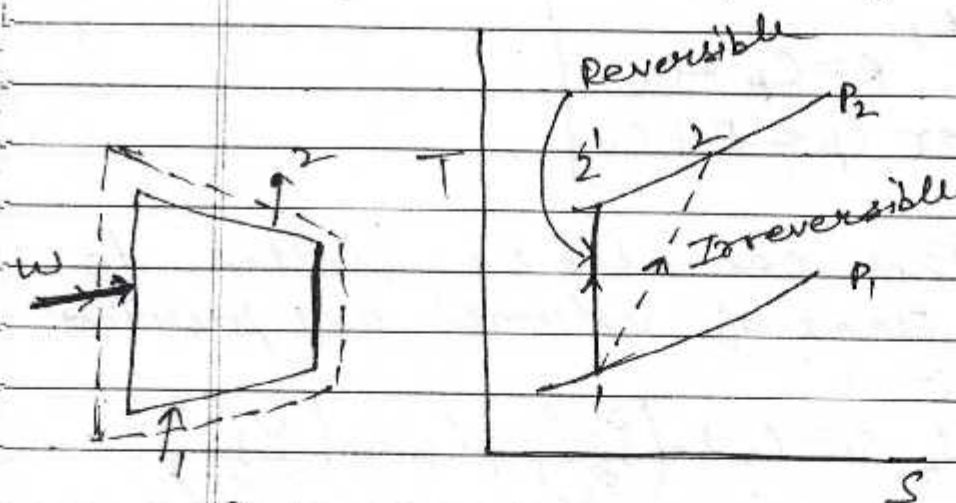
III (ii)
(B)

Implications of the second law of Thermodynamics:-

Since entropy is defined as a property which remains constant during a reversible adiabatic process, it follows that a temperature-entropy diagram would indicate a process by a straight line perpendicular to the entropy axis if the process is purely isentropic. The friction is an irreversible process will cause the temperature of the fluid to be higher than it would have been in a frictionless (reversible) process. The entropy increased during an irreversible process.

$$\left(\frac{T_2}{T_1} \right)$$

Fig:- Isentropic Efficiency Concept.



Isentropic Efficiency:- the entropy change in an irreversible adiabatic ~~process~~ process leads to process efficiency. The ideal constant entropy process is termed isentropic and ratio of the specific work transfer in the ideal process to that in the actual process is called the isentropic efficiency.

Since the majority of adiabatic machines are flow process, isentropic efficiency is usually expressed in terms of the useful work W .

For the Compression process $\Rightarrow \eta_{ic} = \frac{W_{12}'}{W_{12}}$

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And for the expansion process $\Rightarrow \eta_{it} = \frac{w_{12}}{w_{12}'}$

If changes in kinetic & potential energy are negligible, the SFEE may be used to rewrite these expressions in terms of specific enthalpy change and for a perfect gas, enthalpy change may be expressed by temperature change. Thus, for compression processes,

$$\eta_{ic} = \frac{h_2' - h_1}{h_2 - h_1}$$

for perfect gas it becomes

$$\eta_{ic} = \frac{T_2' - T_1}{T_2 - T_1}$$

The physical interpretation of this efficiency is that an irreversible compression process requires more work than an ideal process and in irreversible expansion process gives less work than an ideal process.

IV) THIRD LAW :-

The entropy of a perfect crystal is zero when the temperature of the crystal is equal to absolute zero (0K).

- At 0K, there is no thermal motion, and if the crystal is perfect, there will be no disorder.
- Once the temp. begins to rise above 0, the particles begin to move and entropy gradually increases as the average kinetic energy of the particles increases.
- When temperature reaches the melting point of the substance (T_m), there is an ~~ab~~ abrupt increase in entropy as the substance changes from a solid to a more disordered liquid.

Again the entropy increases gradually as the motion of the particles increases until the temp. reaches the boiling point of the substance (T_b). At this point,

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there is another drastic increase in entropy as the substance changes from a ~~confined~~ confined liquid particles to random motion gas particles.

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→ THERMODYNAMIC PROCESSES :-

* Adiabatic :- No heat transfer from or to the fluid/substance.

* Isothermal :- No change in temperature of the fluid/subs.

* Isobaric :- No change in pressure of the fluid.

* Isochoric :- No change in volume of the fluid.

* Isentropic :- No change of entropy of the fluid.

* Isenthalpic :- No change of enthalpy of the fluid.

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⇒ Enthalpy:-

The enthalpy of a substance, h , is defined as

$$h = u + pv \quad \text{--- (i)}$$

It is an intensive property of a system ~~per~~ (kJ/kg).

Internal Energy change is equal to the heat transferred in a constant volume process involving no work other than pdV work. From equation;

$$dQ = du + pdv$$

it is possible to derive an expression for the heat transfer in a constant pressure process involving no work other than pdV work.

Heat transferred at constant pressure increases the enthalpy of a system. For an ideal gas, the enthalpy becomes

$$h = u + RT$$

Since the internal energy of an ideal gas depends only on the temperature, i.e.,

the enthalpy of an ideal gas also depends only on the temperature, i.e.
 $h = f(T)$

⇒ Energy of an Isolated System:-

An isolated system is one in which there is no interaction of the system with the surroundings. For an isolated system, $dQ = 0$, $dW = 0$.

The first law gives,

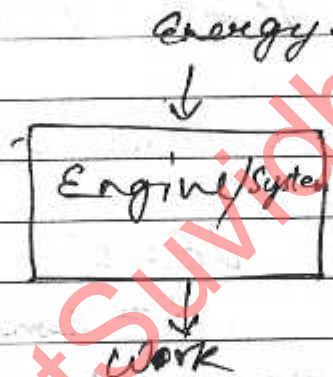
$$dE = 0$$

$$\text{or } E = \text{constant}$$

The energy of an isolated system is always constant.

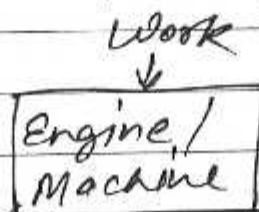
⇒ PMM1 (PMMFK):-

The first law states the general principle of the conservation of energy. There can be no machine which would continuously supply mechanical work without some other form of energy disappearing simultaneously.



Such a fictitious machine is called a perpetual motion machine of first kind (PMMFK). A PMM1/PMMFK is thus impossible.

The converse of the above statement is also true, i.e., there can be no machine which would continuously consume work without some other form of energy appearing simultaneously.



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→ Limitations of First Law:-

The first law deals with the amounts of energy of various forms transferred b/w the systems and its surroundings and with changes in the energy stored in the system. It treats work and heat interactions as equivalent forms of energy in transit and does not indicate the possibility of a spontaneous process proceeding in a certain direction. It is the 2nd Law which assigns a quality to different forms of energy, & also indicates the direction of any spontaneous system.

⇒ Energy Reservoirs (Thermal Energy):-

A thermal energy reservoir (TER) is defined as a large body of infinite heat capacity, which is capable of absorbing or rejecting an unlimited quantity of heat without suffering appreciable changes in its thermodynamic co-ordinates. The changes that do take place in the large body as heat enters or leaves are so very slow and so very so very minute that all process within it are quasi-static.

The thermal energy reservoir TER_H from which heat Q_1 is transferred to the system operating in a heat engine cycle is called the source. The thermal energy reservoir TER_L to which heat Q_2 is rejected from the system during a cycle is the sink. A typical source is a constant temperature furnace where fuel is continuously burnt, and a typical sink is a river or sea or the atmosphere itself.

A mechanical energy reservoir (MER) is a large body enclosed by an adiabatic impermeable wall capable of storing work as potential energy (such as

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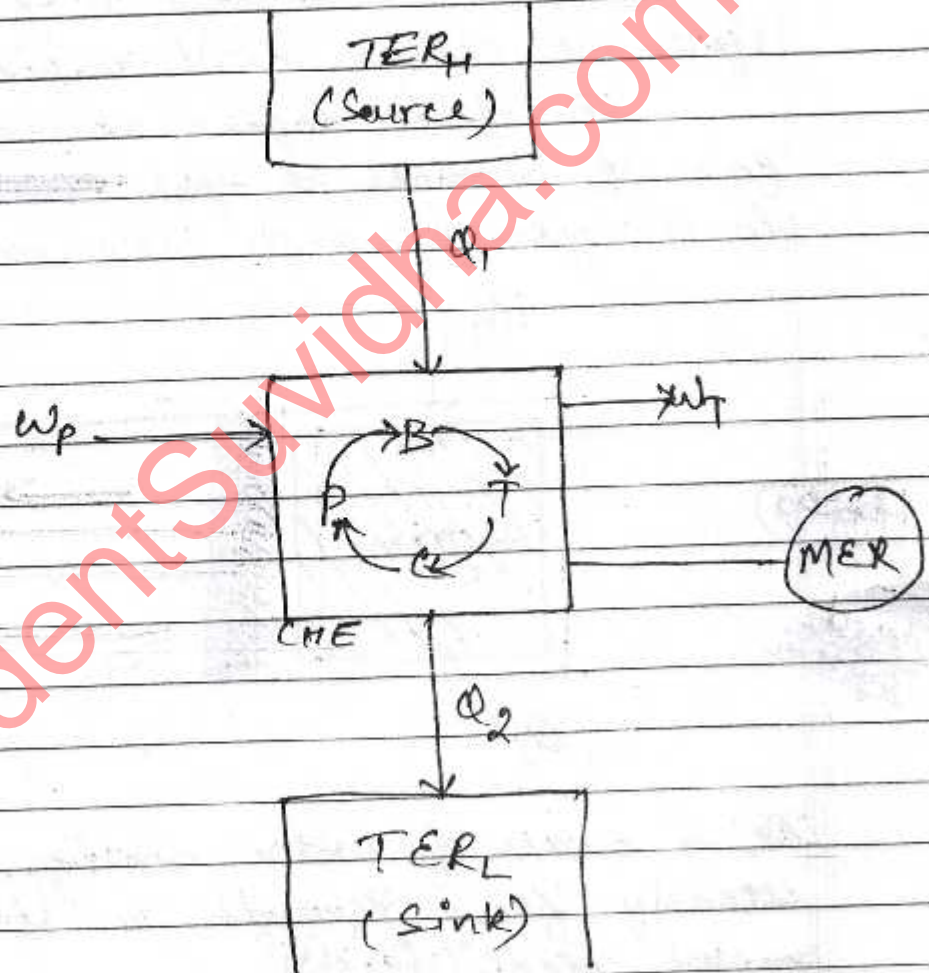
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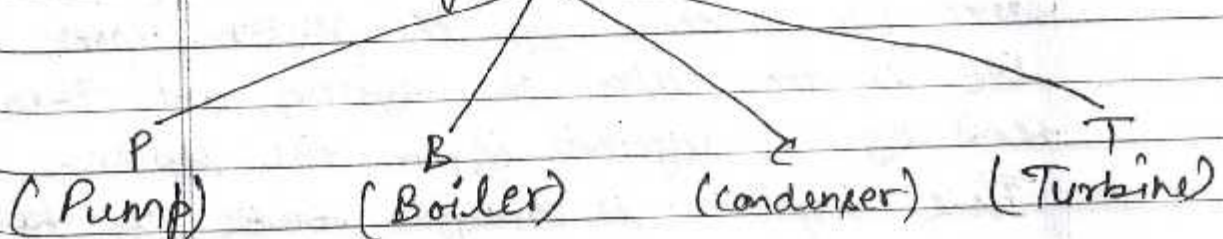
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raised weight or wound spring) or kinetic energy (such as a rotating flywheel). All process of Interest within an MER are essentially quasi-static. An MER receives and delivers mechanical energy quasi-statically.



CHE = Cyclic Heat Engine



⇒ ~~Heat~~ Heat Engine (Cyclic Heat Engine)

A heat engine^{cycle} is a thermodynamic cycle in which there is a net heat transfer to the system and a net work transfer from the system. The system which executes a heat engine cycle is called heat engine.

A heat engine can be in the form of a mass of gas ~~confined~~ confined in a cylinder and piston machine;

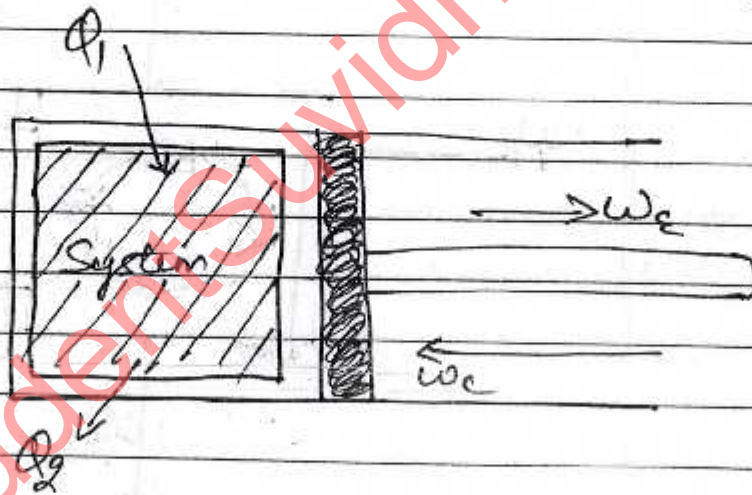


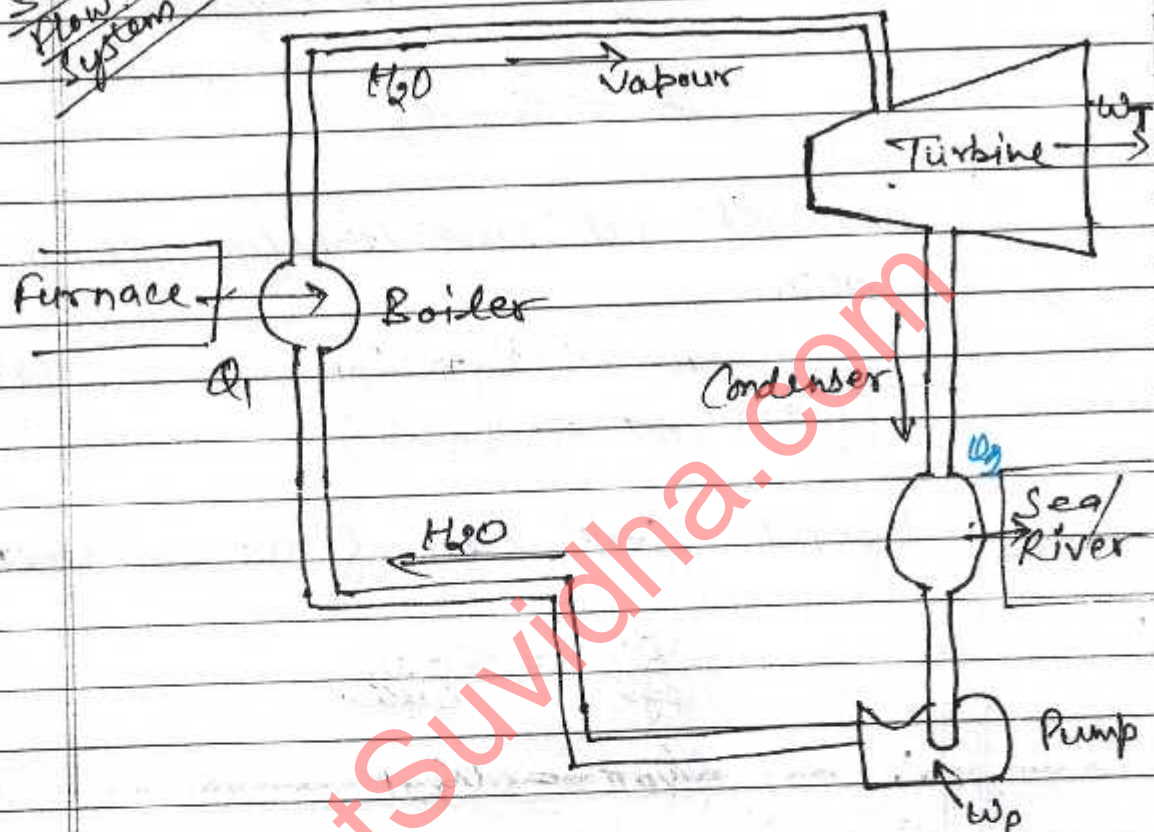
Fig (a)

~~Closed~~
Closed
System

Or a mass of water moving in a steady flow through a steam power plant. (Fig. (b))

In the cyclic heat engine (Fig. (a)) heat Q_1 is transferred to the system, work W_c is done by the system, work W_c is done upon the system, and then heat Q_2 is rejected from the system. The system is brought back to the

Fig (b)

Steady Flow System

initial state through all these four successive processes which constitute a heat engine cycle.

In fig.(b), heat Q_1 is transferred from the furnace to the water in the boiler to form steam which then works on the turbine rotor to produce work W_t , then the steam is condensed to water in the condenser in which an amount Q_2 is rejected from the ~~system~~ system, and finally work W_p is done on the system (water) to pump it to the boiler.

The net heat transfer in a cycle to either of the heat engines.

$$Q_{\text{net}} = Q_1 - Q_2 \quad (1)$$

and the net work transfer in a cycle,

$$W_{\text{net}} = W_T - W_P \quad (2)$$

(or $W_{\text{net}} = W_E - W_C$)

By the first law of TD, we have

$$\sum_{\text{cycle}} Q = \sum_{\text{cycle}} W$$

$$\therefore Q_{\text{net}} = W_{\text{net}}$$

$$\text{or } Q_1 - Q_2 = W_T - W_P \quad (3)$$

The function of a heat engine cycle is to produce work continuously at the expense of heat input to the system. So the net work W_{net} and the heat input Q_1 referred to the cycle are of primary interest. The efficiency of a heat engine or a heat engine cycle is defined as,

$$\eta = \frac{\text{Net work output of the cycle}}{\text{Total heat input to the cycle.}}$$

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cycle

$$= \frac{W_{net}}{Q_1} \quad (4)$$

(1)

From eq. (1), (2), (3) and (4),

a

$$\eta = \frac{W_{net}}{Q_1} = \frac{W_T - W_P}{Q_1} = \frac{Q_1 - Q_2}{Q_1}$$

(2)

$$\eta = 1 - \frac{Q_2}{Q_1} \quad (5)$$

ave

This is also known as the thermal efficiency of a heat engine cycle. A heat engine is very often called upon to extract as much work (net) as possible from a certain heat input, i.e., to maximize the cycle efficiency.

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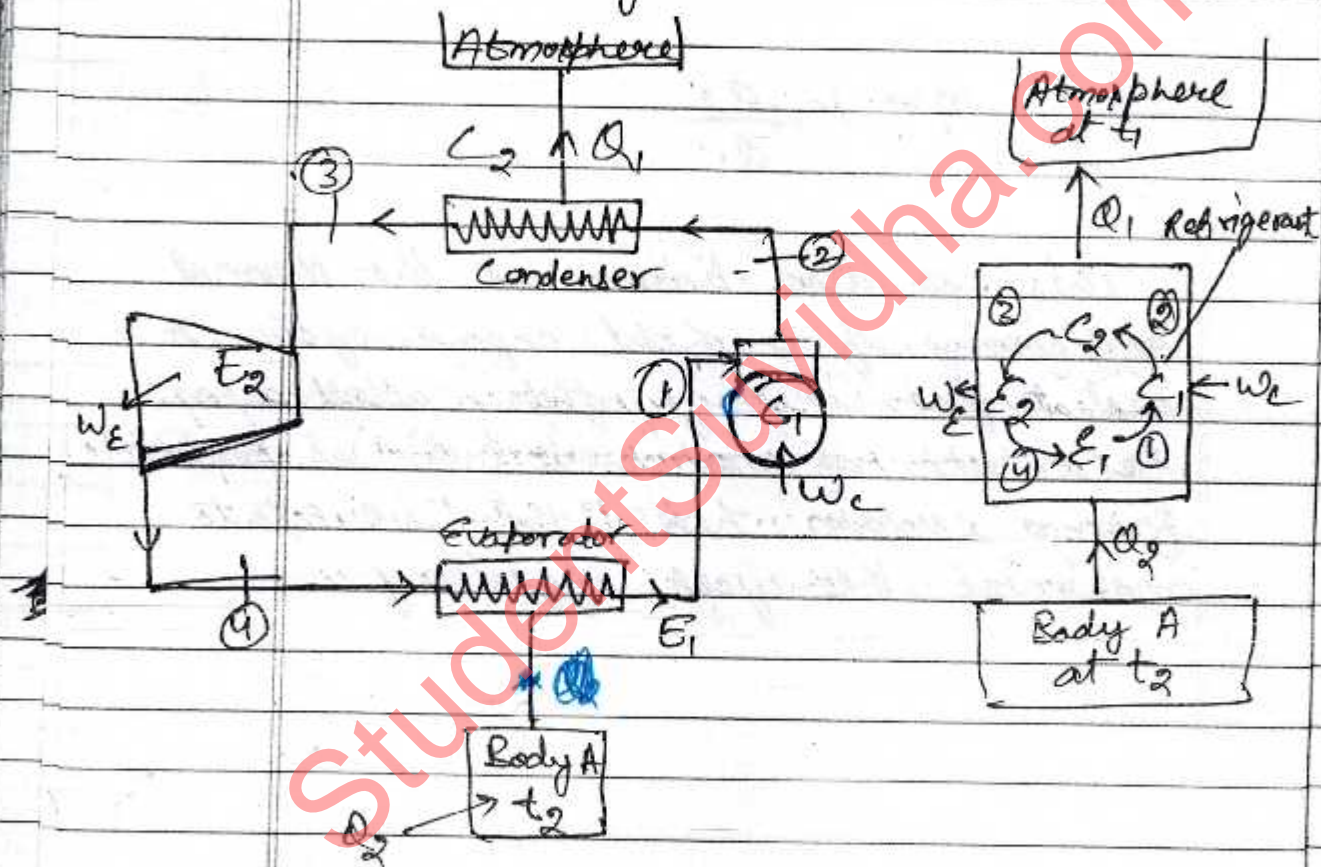
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cycle.

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→ REFRIGERATOR:-

A refrigerator is a device which, operating in a cycle, maintains a body at a temperature lower than the temperature of the surroundings.



Let the body 'A' ~~maintained~~ maintained at t_2 , which is lower than the ambient temperature t_1 . Even though A is insulated, there will always be heat leakage Q_2 into the body from the surroundings by virtue of temperature difference. In order to, keep

body A at the constant temperature t_2 , heat has to be removed from the body at the same rate at which heat is leaking into the body. This heat (Q_2) is absorbed by a working fluid, called the refrigerant, which evaporates in the evaporator E_1 at a temperature lower than t_2 absorbing the latent heat of vaporization from the body A which is cooled or refrigerated (Process 4-1). The vapour is first compressed in the compressor C_1 driven by a motor which absorbs work W_c (Process 1-2), and is then condensed in the condenser C_2 rejecting the latent heat of condensation Q_1 at a temperature higher than that of atmosphere (at t_1) for heat transfer to take place (Process 2-3). The condensate then expands adiabatically through an expander (an engine or turbine) producing work W_p , when a temperature drops to a value lower than t_2 such that heat Q_2 flows from the body A to make the refrigerant evaporate (Process 3-4). Such a cyclic device of flow through E_1 - C_1 - C_2 - E_2 is called a refrigerator. In a refrigerator cycle, attention is concentrated on the body A. Just like efficiency

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in a heat engine cycle, there is a performance parameter in a refrigerator cycle, called the coefficient of performance, abbreviated to COP, which is defined as

$$\text{COP} = \frac{\text{Desired effect}}{\text{Work Input}} = \frac{Q_2}{W}$$

$$\therefore [\text{COP}]_{\text{ref}} = \frac{Q_2}{Q_1 - Q_2}$$

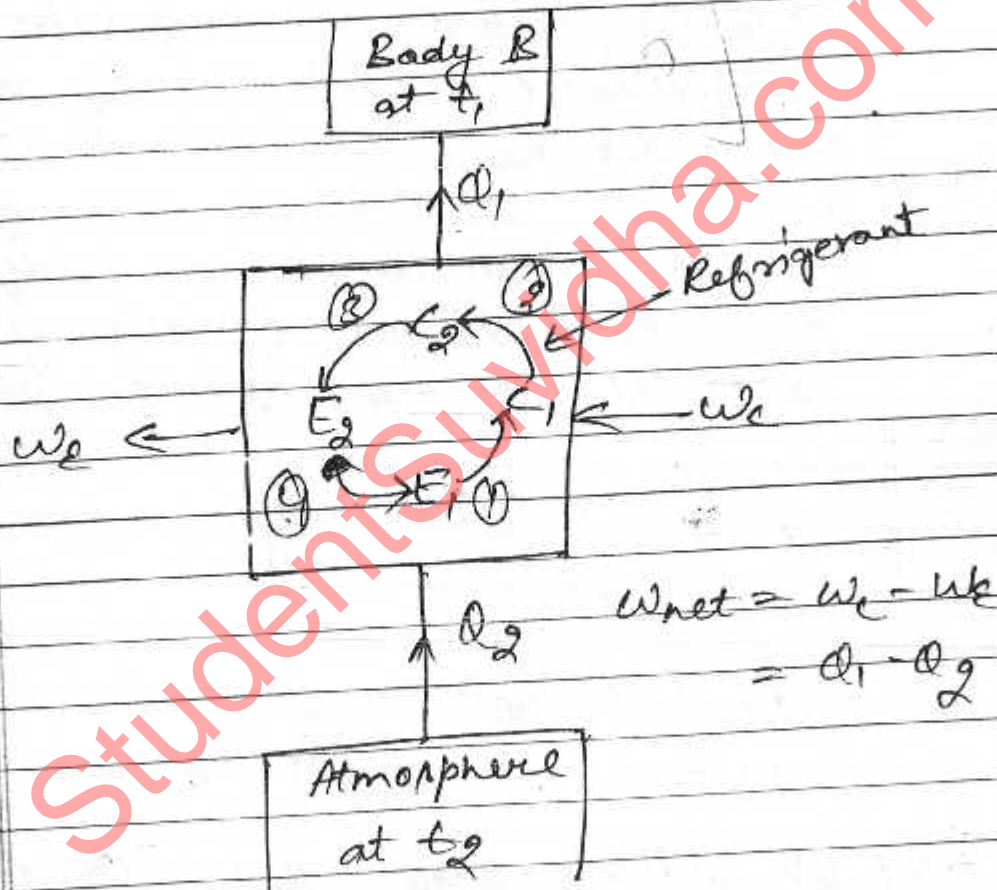
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→ Heat Pump:-

A heat pump is a device which, operating in a cycle, maintains a body, say B, at a temperature higher than the temperature of the surroundings.



By virtue of temperature difference, there will be heat leakage Q from the body to the surroundings. The body will be maintained at the constant temperature t_1 , if heat is discharged into the body at the same rate at which heat leaks out of the body.

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The heat is extracted from the low temperature reservoir, which is nothing but the atmosphere, and discharged into the high temperature body B, with the expenditure of work W in a cyclic device called a heat pump. The working fluid operates in a cycle flowing through the evaporator E_1 , compressor C_1 , condenser C_2 and expander E_2 , similar to a refrigerator but the attention is here focussed on the high temperature body B. Here Q_1 & W are of primary interest, and the COP is defined as;

$Q_1 - W$

Q_2

$$COP = \frac{Q_1}{W}$$

$$\therefore [COP]_{HP} = \frac{Q_1}{Q_1 - Q_2} \quad \text{--- (1)}$$

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We already know that COP of refrigerator is;

$$[COP]_{ref} = \frac{Q_2}{Q_1 - Q_2} \quad \text{--- (2)}$$

from (1) & (2), we found that

$$[COP]_{HP} = [COP]_{ref} + 1 \quad \text{--- (3)}$$